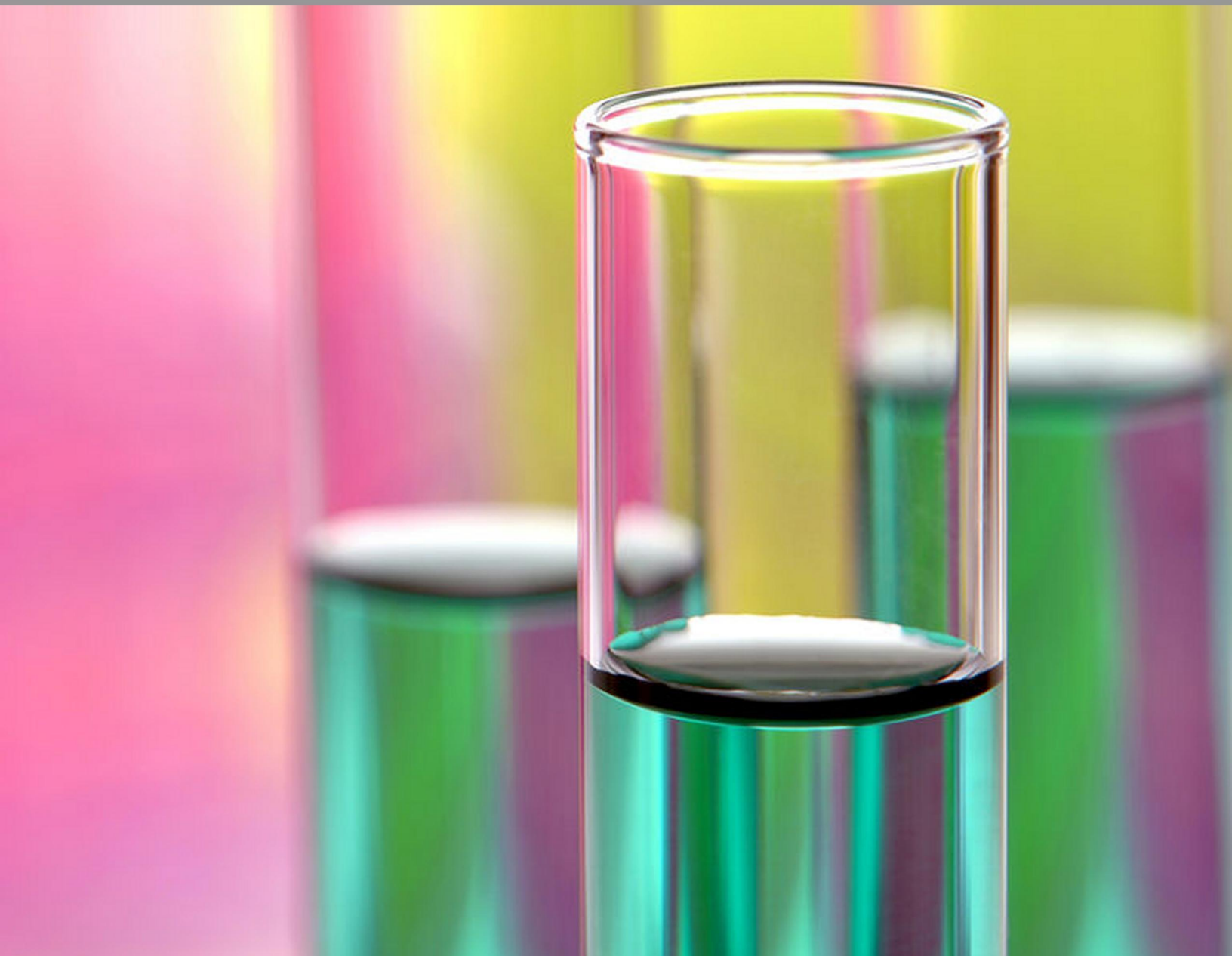


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CK-12 Chemistry - Intermediate Workbook



CK-12 Chemistry - Intermediate Workbook

Donald Calbreath, Ph.D.

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Contents

1	Introduction to Chemistry Worksheets	1
1.1	What is Chemistry?	2
1.2	The Scientific Method	5
2	Matter and Change Worksheets	8
2.1	Properties of Matter	9
2.2	Classification Of Matter	13
2.3	Changes in Matter	16
3	Introduction to Chemistry Worksheets - 1	19
3.1	The International System of Units	20
3.2	Unit Conversions	23
3.3	Uncertainty in Measurements	25
4	Atomic Structure Worksheets	28
4.1	Atoms	29
4.2	The Nuclear Model of the Atom	32
4.3	Isotopes and Atomic Mass	35
5	Electrons in Atoms Worksheets	38
5.1	Light	39
5.2	The Quantum Mechanical Model	42
5.3	Electron Arrangements in Atoms	45
6	The Periodic Table Worksheets	48
6.1	History of the Periodic Table	49
6.2	Electron Configuration and the Periodic Table	52
6.3	Periodic Trends	55
7	Chemical Nomenclature Worksheets	58
7.1	Ionic Compounds	59
7.2	Molecular Compounds	63
7.3	Acids and Bases	65
8	Ionic and Metallic Bonding Worksheets	68
8.1	Ions	69
8.2	Ionic Bonds and Ionic Compounds	72
8.3	Metallic Bonds	75
9	Covalent Bonding Worksheets	78
9.1	Lewis Electron Dot Structures	79
9.2	Molecular Geometry	82
9.3	Polarity and Intermolecular Forces	86

9.4	Hybridization of Atomic Orbitals	89
10	The Mole Worksheets	92
10.1	The Mole Concept	93
10.2	Mass, Volume and the Mole	95
10.3	Chemical Formulas	98
11	Chemical Reactions Worksheets	100
11.1	Chemical Equations	101
11.2	Types of Chemical Reactions	104
12	Stoichiometry Worksheets	107
12.1	Mole Ratios	108
12.2	Stoichiometric Calculations	110
12.3	Limiting Reactant and Percent Yield	111
13	States of Matter Worksheets	114
13.1	Kinetic - Molecular Theory and Gases	115
13.2	Liquids	118
13.3	Solids	121
13.4	Changes of State	124
14	The Behavior of Gases Worksheets	127
14.1	Gas Properties	128
14.2	Gas Laws	131
14.3	Ideal Gases	134
14.4	Gas Mixtures and Molecular Speeds	137
15	Water Worksheets	140
15.1	Properties of Water	141
15.2	Aqueous Solutions	143
15.3	Colloids and Suspensions	146
16	Solutions Worksheets	149
16.1	Solubility	150
16.2	Solution Concentration	153
16.3	Colligative Properties	156
16.4	Net Ionic Equations	159
17	Thermochemistry Worksheets	162
17.1	Heat Flow	163
17.2	Thermochemical Equations	166
17.3	Heat and Changes of State	169
17.4	Hess's Law	172
18	Kinetics Worksheets	175
18.1	Rates of Reactions	176
18.2	Rate Laws	179
18.3	Reaction Mechanisms	182
19	Equilibrium Worksheets	185
19.1	The Nature of Equilibrium	186
19.2	Le Châtelier's Principle	189

19.3	Solubility Equilibrium	193
20	Entropy and Free Energy Worksheets	195
20.1	Entropy	196
20.2	Spontaneous Reactions and Free Energy	199
20.3	Free Energy and Equilibrium	202
21	Acids and Bases Worksheets	205
21.1	Acid-Base Definitions	206
21.2	The pH Concept	211
21.3	Acid and Base Strength	215
21.4	Acid-Base Neutralization	218
21.5	Salt Solutions	221
22	Oxidation Reduction Reactions Worksheets	224
22.1	The Nature of Oxidation and Reduction	225
22.2	Oxidation Numbers	228
22.3	Balancing Redox Reactions	231
23	Electrochemistry Worksheets	233
23.1	Electrochemical Cells	234
23.2	Cell Potentials	237
23.3	Electrolysis	240

CHAPTER

1

Introduction to Chemistry Worksheets

Chapter Outline

1.1 WHAT IS CHEMISTRY?

1.2 THE SCIENTIFIC METHOD

1.1 What is Chemistry?

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Define chemistry.

1. Chemistry is the study of
 - a. living systems
 - b. the stars and planets
 - c. all matter
 - d. reactions in a test tube
2. All of the following are characteristics of matter except
 - a. matter can disappear and reappear
 - b. matter has mass
 - c. matter occupies space
 - d. all things are composed of matter
3. Which of the following is not a chemistry topic?
 - a. the composition of ocean water
 - b. what ocean fish eat
 - c. the height of waves in surf
 - d. what a surf board is made of
4. Which of the following was not a goal of the alchemists?
 - a. elixir of life
 - b. changing lead into rubies
 - c. philosophers stone
 - d. making gold from base metals

Lesson Objective: Differentiate between the macroscopic and the microscopic as it relates to chemistry.

5. The term used for materials we can see is
 - a. macrolitic
 - b. macrosteric
 - c. macrochemical
 - d. macroscopic
6. True/False: "Microscopic" refers only to those materials we can see under a microscope.

7. A brick is a form of _____ matter.
8. True/False: Individual acetylsalicylic acid (aspirin) molecules are microscopic.
9. True/False: Studying how the molecules change in the reaction between oxygen and methane is an example of a macroscopic process.

Lesson Objective: Know the relationship between pure chemistry and applied chemistry.

10. True/False: Pure chemistry is focused on understanding basic chemical processes.
11. True/False: Applied chemistry does not have a specific goal or application.
12. Research that develops a new type of fireproof fabric is (pure, applied) chemistry.
13. Studying the chemical processes involved in the formation of elements in the sun is an example of (pure, applied) chemistry.
14. The following is an example of pure chemistry
 - a. developing a new fertilizer
 - b. studying how pesticides kill harmful insects
 - c. measuring the effect of temperature on how fast a reaction occurs
 - d. finding new ways to lower cholesterol levels in the blood
15. Read the following passage and answer the questions:
16. The line between pure chemistry and applied chemistry is not always distinct. For example, in the early 1960s, chemists at DuPont were searching for a new lightweight and strong fiber that could be used in tires. Chemist Stephanie Kwolek discovered, somewhat by accident, that a certain solution she had made displayed unique characteristics that were unlike those of other previously developed substances. The rapid expansion in the field of polymer chemistry that has occurred over the past 50 years was partially due to the results of her research. Polymers are very, very large molecules comprised of smaller subunits that are repeated over and over again in extremely long chains. The polymer that was discovered at DuPont was eventually given the name Kevlar. It is used not only in tires but also in bulletproof body armor because of its high strength and light weight. Polymer chemistry continues to be an active and vibrant field of chemistry, as both a pure and an applied discipline.
 - a. Was the initial research that Stephanie Kwolek carried out an example of pure or applied chemistry?
 - b. When did the project become an example of applied chemistry?

Lesson Objective: Identify and describe the five primary disciplines of chemistry.

16. An analytical chemist is involved with
 - a. studies of what penguins eat
 - b. research to develop new rocket fuels
 - c. the synthesis of new carbon compounds
 - d. measurement of the amount of minerals in cereals
17. Making new compounds for high-speed tires is best done by the
 - a. physical chemist
 - b. organic chemist
 - c. inorganic chemist
 - d. biochemist
18. True/False: A biochemist studies chemical processes in cells.
19. True/False: A physical chemist would be the best person to determine the amount of gold in an ore sample.
20. Reactions involving carbon compounds are studied by an _____ chemist.

21. The _____ chemist would study how fast nitrogen and hydrogen form ammonia.
22. How the body uses food is best studied by a _____.
23. Chemical processes in rocks might best be studied by an _____ chemist.

Lesson Objective: Describe some of the concerns of the modern world in which chemistry has played and will continue to play a role.

24. One advantage of using nuclear energy instead of fossil fuels is
 - a. nuclear energy is safe
 - b. nuclear wastes are easy to store
 - c. it is easy to work with nuclear materials
 - d. more energy can be obtained from a given amount of nuclear material

25. Why was lead removed from gasoline?
 - a. it made cars run too fast
 - b. lead caused brain damage
 - c. lead became too expensive to put in gasoline
 - d. mercury is a better additive than lead

26. _____ is a field that involves the manipulation of DNA.
27. Diseased arteries can be replaced by _____ to help patients.
28. True/False: Drugs work because of their effects on chemical processes in the body.
29. True/False: Superconductors are only 55% efficient.
30. Liquid crystals are found in displays in _____ and _____.
31. True/False: Excessive use of fertilizers can harm the water supply.
32. Why is it important for chemists to carry out research looking at materials that harm the environment?

1.2 The Scientific Method

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Describe how the Renaissance period in history changed the approach to science.

1. The Renaissance
 - a. occurred between the twelfth and thirteenth centuries
 - b. took place during the Middle Ages
 - c. was characterized by a dogmatic approach to learning
 - d. brought about new ways of thinking

2. One new approach brought by the Renaissance was
 - a. new methods in painting
 - b. experimentation in science
 - c. the sun was believed to orbit around the earth
 - d. the use of leeches to treat disease

3. True/False: Leonardo da Vinci developed the idea of astronomy.
4. True/False: Copernicus studied movement of the earth.

Lesson Objective: Identify the steps of the scientific method.

5. The scientific method
 - a. is a recipe for doing science
 - b. will always give the right answer
 - c. is a systematic approach to the study of phenomena
 - d. involves preconceived ideas

6. The first step in the scientific method is
 - a. to ask a question
 - b. to set up lab equipment
 - c. to perform an experiment
 - d. to calculate your data

7. A scientific experiment must be
 - a. inexpensive
 - b. repeatable
 - c. complicated

- d. hard to explain
8. True/False: A hypothesis is a “first-guess” explanation.
 9. True/False: A reproducible experiment means something is wrong with your hypothesis.
 10. True/False: A law is always true.
 11. A hypothesis is a _____ that proposes an answer to a question.
 12. To show that a hypothesis is valid, we must be able to _____ it.
 13. We want to test the effect of adding salt to ice to study melting. The _____ group would have no salt added to the ice.
 14. A hypothesis that cannot be proven by experiments must be either _____ or _____.
 15. Read the following passage and answer the questions at the end:
 16. If many experiments are performed and each and every one supports the hypothesis, it may be that the hypothesis can be accepted. A theory is an explanation that has been repeatedly tested and confirmed by multiple researchers and has gained wide acceptance. The term theory is one that is often misunderstood by people in the general population who are unfamiliar with the scientific method. Though a scientist may state that a theory cannot be proven correct, he or she is certainly not saying that it is in doubt. A scientific theory has generally been confirmed over and over and over again, although small details within a theory may be adjusted from time to time as new discoveries are made. In a later chapter, you will learn about atomic theory and how it has changed over the last few centuries. While the existence of atoms has not been in question for some time, new details about how the atom functions have resulted in many updates and alterations of the essential theory.
 - a. Why do many people misunderstand the definition of a theory?
 - b. Does a theory ever change?
 - c. Does a law ever change?

Lesson Objective: Differentiate between the independent variable and the dependent variable in an experiment.

16. The dependent variable in an experiment
 - a. never changes
 - b. will vary according to the temperature
 - c. shows changes during the experiment.
 - d. is always the larger value.
17. The independent variable in an experiment
 - a. is changed to see its effect on the experiment
 - b. never changes
 - c. changes as the dependent variable is altered
 - d. is not affected by the experiment
18. A control group
 - a. changes with the independent variable
 - b. changes with the dependent variable
 - c. is not exposed to the independent variable
 - d. is not exposed to the dependent variable
19. True/False: To study the effect of sunlight on different plants, I expose the plants to the same amount of sunlight. The independent variable is the sunlight.

20. True/False: A good control group for the plant/sunlight study described in the previous questions would be the same plants kept in the dark.
21. True/False: The independent variable is the one that is changed during the experiment.
22. I want to study the effect of different chemicals on how fast ice melts. I sprinkle five different chemicals in ice and measure the speed of melting. Which is the independent variable in the experiment?
23. Many drugs are used to lower blood cholesterol in order to prevent heart attacks. I want to study the effect of drug A on blood cholesterol concentrations. Patients who are not treated with either drug make up the _____ group. To study the effect of drug B on cholesterol levels, I give five groups of patients five different doses of drug B and measure their blood cholesterol concentrations. Drug B is the _____ variable and the cholesterol concentration I measure is the _____ variable.

Lesson Objective: Describe how scientists work in research groups and communicate their results.

24. Scientists usually work
 - a. in teams
 - b. alone
 - c. in labs
 - d. by committee
25. One of the following is usually not paid for by a research grant
 - a. lab equipment
 - b. vacations
 - c. salaries
 - d. lab supplies
26. True/False: Research groups often have regular meetings to share results and ideas.
27. True/False: It is a good idea to keep your research a secret so that nobody else can steal your ideas.
28. Research progress usually occurs in _____ steps.
29. Many scientists work for a _____ .
30. The National Science _____ funds many grants.
31. Science journal articles are usually _____ and _____.
32. Why are journal articles important in the spread of scientific knowledge?

CHAPTER **2**

Matter and Change Worksheets

Chapter Outline

- 2.1 PROPERTIES OF MATTER
 - 2.2 CLASSIFICATION OF MATTER
 - 2.3 CHANGES IN MATTER
-

2.1 Properties of Matter

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe a substance according to its physical properties.

- One of the following is not a physical property of a substance.
 - conductivity
 - reactivity with acids
 - degree of magnetism
 - melting point
- Which of the following is a pure substance?
 - ocean water
 - vanilla ice cream
 - a diamond
 - quartz crystals
- True or False: The identity of a pure substance will change when it is melted or frozen.
- True or False: A physical change does not alter the identity of the substance.
- The ability of copper to be made into sheets is called _____ .
- Not all metals are _____ .
- Sodium metal melts at 97°Celsius. The melting point is a _____ property.

Lesson Objective: Distinguish between extensive and intensive properties.

- An example of an intensive property would be
 - mass of the material
 - volume of the material
 - color of the material
 - length of the material
- How copper conducts electricity is an
 - intensive property
 - intangible property
 - extensive property
 - expansive property
- True or False: The length of a copper cable is an extensive property.
- True or False: Intensive properties depend on the color of the material.

Lesson Objective: Describe the three states of matter.

- The state of matter that has a definite shape and volume is

- a. a crystal of sodium chloride
 - b. neon gas in a tube
 - c. a quart of milk
 - d. carbon dioxide in air
13. Only one of the following items can be compressed.
- a. sand
 - b. ice
 - c. steam
 - d. water
14. Which of the following is true about gases?
- a. gas molecules are arranged in an orderly fashion
 - b. gas molecules are in constant motion
 - c. hydrogen and nitrogen are liquids at room temperature
 - d. a vapor is a solid at room temperature
15. Which of the following is not true of liquids?
- a. liquids have a definite volume
 - b. liquids can form vapors at higher temperatures
 - c. the molecules of a liquid are not rigidly ordered
 - d. liquids are very compressible.
16. True or False: Solids are very compressible.
17. True or False: The particles in a solid are arranged in an orderly fashion.
18. True or False: There are large spaces between the molecules of a gas.

Match the following answers to the definitions in Questions 19-22:

- a. solid
 - b. gas
 - c. vapor
 - d. liquid
19. _____ easily compressible
20. _____ gaseous state of a substance that is a solid or liquid at room temperature.
21. _____ a form of matter with a definite volume, but an indefinite shape.
22. _____ the form of matter illustrated by a gold nugget.

Lesson Objective: Identify physical changes to matter.

23. Read the following passage and answer the questions.
24. As an ice cube melts and it acquires the ability to flow, its shape changes. However, its composition does not change. Melting is an example of a physical change. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include vaporization (liquid to gas), freezing (liquid to solid), and condensation (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved in water, the salt is said to have entered the aqueous state. Boiling off the water will cause the salt to reform in its solid state. An example of an irreversible physical change would be grinding a piece of wood into sawdust. Such a change is irreversible because the sawdust cannot be reconstituted into the same piece of wood that it was before. Cutting the grass or pulverizing a rock would be other irreversible physical changes.

- a. Would boiling water be reversible if the water vapor goes off into the atmosphere?
 - b. Is burning gasoline in a car engine a physical change?
25. One of the following is not a physical change
- a. burning wood
 - b. chopping wood up
 - c. making furniture from wood
 - d. stacking wood
26. All of the following except one are reversible physical changes
- a. boiling water
 - b. making ice cubes
 - c. bagging corn kernels
 - d. making popcorn
27. True or False: A physical change involves a change in the identity of the sample.
28. True or False: Making earrings from a gold nugget is an irreversible physical change.
29. True or False: Bending aluminum sheets is an irreversible physical change.
30. You cook an egg on the stove and heat your coffee in the microwave. Which process is a physical change? Explain your answer.
-

Answer Key

1. b
 2. c
 3. false
 4. true
 5. malleability
 6. conductive
 7. physical
 8. c
 9. a
 10. true
 11. false
 12. a
 13. c
 14. b
 15. d
 16. false
 17. true
 18. true
 19. b
 20. c
 21. d
 22. a
1. Yes, because the water vapor could condense and form rain somewhere else.
 2. No, the gasoline molecules are converted to other substances.
23. a
 24. d
 25. false

26. true
27. false
28. Heating the coffee is a physical change since the material is not altered. Cooking the eggs changes the chemical composition and is not a physical change.

}}

2.2 Classification Of Matter

Worksheet

{{Teachers lbody=

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Define a mixture and understand why mixtures are different than pure substances.

- Which of the following is a pure substance
 - oxygen
 - air
 - salt water
 - maple syrup
- One of the following is not a property of a mixture
 - always has the same composition
 - is a physical blend of two or more substances
 - each component retains its own identity
 - can be separated into components
- True or False: Sodium chloride is a mixture of two elements.
- True or False: A mixture can have more than two components.
- Classify each of the following as either a mixture or a pure substance:
 - orange juice _____
 - paper _____
 - nitrogen _____
 - blood _____

Lesson Objective: Classify mixtures as homogeneous or heterogeneous.

- Which of the following is a homogeneous mixture?
 - crunchy peanut butter
 - pepperoni and mushroom pizza
 - a glass of root beer
 - macaroni and cheese
- A phase has all of the following characteristics except one
 - uniform composition
 - properties are the same throughout
 - is always a pure substance
 - can be a gas or a liquid
- Only one of the following is a characteristic of a heterogeneous mixture
 - composition is uniform throughout the mixture

- b. contains two or more phases
 - c. is a pure substance
 - d. the mixture composition is always the same
9. True or False: Oil and water always mix evenly.
10. True or False: A heterogeneous mixture consists of two or more phases
11. Identify each of the following as either a homogeneous or a heterogeneous mixture.
- a. motor oil _____
 - b. sand _____
 - c. sawdust _____
 - d. milk _____

Lesson Objective: Describe several ways to separate mixtures.

12. Distillation could not be used to separate one of the following:
- a. rubbing alcohol and water
 - b. motor oil and brake fluid
 - c. sand and gold
 - d. gasoline fractions
13. The best way to separate lead pellets from ginger ale is
- a. hand-pick all the pellets
 - b. distillation
 - c. filtration
 - d. melt the lead
14. True or False: Distillation separates materials based on melting point.
15. True or False: Hot water is used to cool a distillation column.
16. Read the passage below and answer the following questions
17. Sand does not dissolve in water. In order to separate sand from water, the mixture can be poured through a filter. The tiny holes of a filter, called pores, allow the water to pass through while trapping the sand grains behind. Filtration is a technique that separates solids from liquids in a heterogeneous mixture. Distillation is a method of separating a mixture based on differences in boiling points. A small-scale laboratory distillation apparatus is used to separate substances based on differences in their boiling points. For example, the impure liquid could be a mixture of alcohol and water. As the mixture is heated, the alcohol begins to boil before the water does because of its lower boiling point. The alcohol vapors rise up in the flask and into the long cylindrical tube, which is called a condenser. Cold water is continually pumped into the outside of the condenser, which causes the alcohol vapors to condense back into the liquid state. The purified alcohol then drips into the flask labeled “distilled liquid,” while the water is left behind in the original flask. Distillation is commonly used to purify water by separating it from the various dissolved minerals that are found in natural water sources.
- a. Distillation is based on what physical property? _____
 - b. In a mixture of acetone (boiling point 132°C) and ethanol (78°C), which component will distill off first?

 - c. What will happen to the components if the mixture is placed in an oven at 200°C?

Lesson Objective: Distinguish between elements and compounds.

17. Only one of the following is not a characteristic of a compound
- a. composed of two or more elements
 - b. cannot be broken down into simpler substances
 - c. an undergo chemical changes

- d. elements are combined in fixed proportions
18. The symbol Cu designates the element
- copper
 - lead
 - mercury
 - cobalt
19. Antimony has the following chemical symbol:
- Au
 - Ag
 - Sb
 - Sn
20. The Latin name for tin is
- natrium
 - kalium
 - stibium
 - stannum
21. True or False: The chemical symbol for potassium is K.
22. True or False: The symbol Fe is used to represent the element lead.
23. True or False: A chemical symbol for an element can have one or two letters.
24. True or False: A specific chemical symbol can be used to designate more than one element.
25. Would you rather have jewelry made of argentum or natrium? _____
26. The two elements found in a mixture of aurum and plumbum are _____ and _____.
27. You have a mixture of potassium chloride, water, and olive oil. How would you separate these materials from one another?

2.3 Changes in Matter

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Identify the chemical properties of a substance.

1. Chemical properties can only be observed
 - a. in the dark
 - b. as the substance is being changed
 - c. before materials are mixed together
 - d. when materials are frozen
2. A chemical property describes
 - a. the behavior of material when heated
 - b. the color of the product
 - c. the ability of a substance to undergo chemical change
 - d. the amount of material used
3. True or False: A chemical property is the same as a physical property.
4. True or False: Change is necessary in order to identify a chemical property.
5. True or False: Rust formation is not a chemical property of iron.

Lesson Objective: Describe chemical changes and differentiate them from physical changes.

6. In a physical change
 - a. a new material is created
 - b. the color will change
 - c. the composition of the material is altered
 - d. the identity of the substance does not change
7. One of the following terms is not used to describe a chemical change
 - a. transmutation
 - b. fermentation
 - c. decomposition
 - d. combination
8. Which of the following terms is used to describe a material on the left-hand side of the arrow in a chemical reaction?
 - a. product
 - b. reagent
 - c. resultant
 - d. reactant
9. Rusting is an example of what kind of change?

- a. chemical
 - b. physical
 - c. decomposition
 - d. biological
10. True or False: Chemical reactions are usually written left to right.
 11. True or False: The product is a starting material in a chemical reaction.
 12. True or False: The rusting of iron is a chemical reaction between tin and iron.
 13. True or False: Zinc does not react with sulfur at room temperature.
 14. One or more _____ must be formed if a chemical reaction has taken place.
 15. What is the product of the reaction between zinc and sulfur? _____
 16. Write the chemical formula for the above product _____
 17. Read the passage below and answer the following questions:
 18. The reaction between zinc and sulfur can be depicted in something called a chemical equation. In words, we could write the reaction as: zinc + sulfur \rightarrow zinc sulfide A more convenient way to express a chemical reaction is to use the symbols and formulas of the substances involved: $\text{Zn} + \text{S} \rightarrow \text{ZnS}$ The substance(s) to the left of the arrow in a chemical equation are called reactants. A reactant is a substance that is present at the start of a chemical reaction. The substance(s) to the right of the arrow are called products. A product is a substance that is present at the end of a chemical reaction. In the equation above, zinc and sulfur are the reactants that chemically combine to form zinc sulfide as a product.
 - a. How do we recognize a reactant in a chemical equation?
 - b. How do we recognize a product in a chemical reaction?
 - c. What does the arrow stand for?
 - d. What is an advantage of using chemical symbols?

Lesson Objective: Use various visual clues to identify whether a chemical reaction is taking place.

18. When heated, mercuric oxide produces
 - a. mercuric nitride
 - b. oxynitride
 - c. oxygen
 - d. nitrogen oxides
19. The reaction of zinc with hydrochloric acid produces
 - a. hydrogen
 - b. chlorine
 - c. oxygen
 - d. nitrogen
20. The reaction of zinc and hydrochloric acid produces
 - a. a cold test tube
 - b. a yellow color
 - c. a warm test tube
 - d. a bad smell
21. The best way to know that a chemical reaction has taken place is to
 - a. look at the shape of the materials
 - b. see if the color has changed
 - c. test the composition of the material after change has taken place
 - d. look at the equation that was written before the reaction was run.
22. True or False: All chemical changes involve a transfer of energy.
23. True or False: A precipitate is not a sign that a chemical reaction has occurred.

24. True or False: When a gas condenses, it releases energy.
25. The metal formed in the decomposition of mercuric oxide is _____.
26. Lead iodide produces a _____ when it is formed from solutions of lead nitrate and potassium iodide.
27. The process of heated magnesium reacting with oxygen is called _____.
28. Many reactions display a change in _____ to show that a reaction has occurred.
29. Lead iodide produces a _____ color when lead nitrate and potassium iodide are mixed.
30. _____ gas would be expected to be released during the reaction of hydrochloric acid with iron since this metal behaves the same way zinc does.
31. One visible form of energy transfer in a chemical reaction is _____.
32. What is the product of the reaction between magnesium and oxygen ? _____
33. You have two bottles of colorless solutions that could be potassium iodide (bottle one), and hydrochloric acid (bottle two). Unfortunately, the labels cannot be read. How would you clearly demonstrate the identity of both bottles? You can make use of any other chemicals you need for the testing.

CHAPTER

3**Introduction to Chemistry
Worksheets - 1****Chapter Outline**

- 3.1 THE INTERNATIONAL SYSTEM OF UNITS**
 - 3.2 UNIT CONVERSIONS**
 - 3.3 UNCERTAINTY IN MEASUREMENTS**
-

3.1 The International System of Units

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Identify the seven base units of the International System of Units.

1. The abbreviation for the base unit for length is the
 - a. Km
 - b. m
 - c. M
 - d. K
2. The abbreviation kg stands for
 - a. kilometer
 - b. kilomole
 - c. kilogram
 - d. kiloliter
3. True or False: SI is the abbreviation for the Spanish term Le Syst eme International d'Unit es.
4. True or False: All metric units are based on multiples of ten.

Lesson Objective: Know the commonly used metric prefixes.

5. An 8 gigabyte computer memory chip can be purchased for \$73.99. This memory capacity is equivalent to
 - a. 800 megabytes
 - b. 8000 megabytes
 - c. 80 megabytes
 - d. 80,000 megabytes
6. The SI prefix kilo has an exponential factor of
 - a. 1×10^3
 - b. 1×10^4
 - c. 1×10^5
 - d. 1×10^2
7. It takes _____ mL to make 2.5 L.
 - a. 250
 - b. 0.25
 - c. 2500
 - d. 25
8. True or False: One decaliter = ten liters.
9. True or False: It takes one hundred kilometers to make one megameter.
10. True or False: A convenient unit measuring very small volumes is the microliter.

11. Give the decimal or scientific notation for the following:

- 10240 _____
- 4.98×10^2 _____

Lesson Objective: Convert between the Celsius and Kelvin temperature scales.

12. Which temperature scale is typically not used for scientific measurements?

- Celsius
- Fahrenheit
- Centigrade
- Kelvin

13. The Celsius scale has ____ degrees between the freezing point and the boiling point of water.

- 75
- 271
- 200
- 100

14. True or False: The Celsius scale was developed by the Scottish physician and mathematician Anders Celsius.

15. True or False: There is no molecular motion at 0 K.

16. An object's _____ energy is the energy due to motion.

17. Temperature is a measure of the average _____ of the _____ in motion.

18. The boiling point of water on the Kelvin scale is _____.

19. Fill in the following table:

TABLE 3.1: Table

Condition	Temperature (K)	Temperature (°C)
freezing point of water		
convert	496 K	
convert		318°C
convert	212 K	
convert		-43°C

Lesson Objective: Understand volume and energy as combinations of SI Units.

20. The units of volume are derived from units of

- sound
- length
- mass
- weight

21. The SI unit of volume is the cubic

- meter
- liter
- gram
- candela

22. True or False: The particles in a hot object move faster than those in a cold object.

23. True or False: Temperature determines the direction of heat transfer.

24. _____ energy is the type of energy stored in matter.

25. The English physicist _____ developed the idea of the SI unit of energy, which is called the _____.

Lesson Objective: Distinguish between mass and weight.

26. Weight measures the effect of _____ on an object
- temperature
 - pressure
 - sunlight
 - gravity
27. True or False: Mass changes with location.
28. Read the passage below and then answer the questions:
29. Energy is defined as the capacity to do work or to produce heat. As discussed previously, kinetic energy is one type of energy and is associated with motion. Another frequently encountered form of energy is potential energy, which is a type of energy that is stored in matter. The joule (J) is the SI unit of energy and is named after English physicist James Prescott Joule (1818-1889). In terms of SI base units, a joule is equal to a kilogram times a meter squared divided by a second squared
30. $(kg \cdot m^2 / s^2)$
31. . How do the units for the joule incorporate the idea of motion? Why is mass a part of the units for the joule?
32. A non-SI unit of energy is the calorie –another way of indicating the amount of chemical energy in a material. One medium potato has 110,000 calories as determined by laboratory measurement (the potato is completely burned with oxygen and the energy released is measured). However, if an internet search is carried out, one potato producer’s web site shows this potato to contain 110 Calories (equivalent to the 110,000 “regular” calories). Why would they list the smaller number instead of the scientifically correct one?

TABLE 3.2: Table

Condition	Temperature (K)	Temperature (°C)
freezing point of water	273.15	0
convert	496	223
convert	591	318
convert	212	-61
convert	230	-43

3.2 Unit Conversions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Identify and use conversion factors.

- In order to convert mL to liters, the following conversion factor should be used:
 - 100 mL/liter
 - 1 L/1000 mL
 - 1 L/100 mL
 - 10 mL/L
- The following conversion factor will allow conversion of m to cm.
 - 1 m/1000 cm
 - 10 cm/1 m
 - 100 cm/m
 - 10 m/1 cm
- Which conversion factor will convert grams to kilograms
 - g/1000 kg
 - kg/1000 g
 - kg/100 g
 - g/10 kg
- True or False: 800 grams is equivalent to 0.8 kilograms.
- True or False: 40 meters is equivalent to 400 centimeters.

Lesson Objective: Use the method of dimensional analysis to convert between units.

- True or False: When using dimensional analysis, the actual size of the quantity changes as well as the numbers involved.
- True or False: Dimensional analysis uses units to solve problems correctly.
- You are driving on a road in Mazatlan, Mexico that has a speed limit of 60 km/hour. Your speedometer reads 52 miles/hour. One km = $\frac{5}{8}$ mile. Will you get a speeding ticket?
- A procedure for the analysis of gold ore requires 35 grams of ore. There are 22 students in the class that need to carry out the analysis. How much gold ore is needed for this class?
- An experiment requires 50 mL of solution A for each student and you have 34 students in the class. The student worker prepares 1.5 liters of solution A for the class. Will there be enough solution A for this experiment?
- There are 25 veeblfletzers (vb) to a flexer(fl). Forty fl have a mass of 56 grams. What is the mass in grams of 335 vb?

Lesson Objective: Understand density as a physical property of matter.

- The SI units for density are

- a. gm/L
 - b. kg/dl
 - c. kg/m³
 - d. kg/cm³
13. The density of carbon dioxide is
 - a. greater than that of water
 - b. less than that of air
 - c. greater than that of radon
 - d. less than that of helium
 14. Diiodomethane has a density of 3.325 g/cm³. Which of the following will float on this liquid?
 - a. copper
 - b. lead
 - c. aluminum
 - d. gold
 15. True or False: Density is an intensive property.
 16. True or False: Ethanol is more dense than corn syrup.
 17. True or False: Helium is lighter than air.
 18. Density is defined as _____ divided by _____ .

Lesson Objective: Perform calculations with derived units, including density.

19. The units for area are _____.
20. Moles/L is a way to express _____.
21. Volume is measured in _____.
22. The surface of a table measures 360 cm × 2.4 m. What is the area in m²?
23. A metal sample has a mass of 24.5 grams and occupies a volume of 2.16 mL. What is the metal?
24. A sample of radon gas occupies a volume of 358 mL at 20°C. What is the mass of the gas?
25. A container holds 24.3 L of a solution containing 0.5 moles/L sodium chloride. How many moles of sodium chloride are there in the solution?
26. The SI units of density are kilograms per cubic meter (kg/m³), because the kg and the m are the SI units for mass and length, respectively. Unfortunately, this unit is awkwardly large for everyday usage in the laboratory. Most solids and liquids have densities that are conveniently expressed in grams per cubic centimeter (g/cm³). Since a cubic centimeter is equal to a milliliter, density units can also be expressed as g/mL. Gases are much less dense than solids and liquids, so their densities are often reported in g/L. Densities of some common substances at 20°C are listed below (Table 1.4). Since most materials expand as temperature increases, the density of a substance is temperature-dependent and usually decreases as temperature increases.
27. Why would the density of a substance decrease as the temperature increases?
28. A student wants to convert cm/sec to miles/hour. He sets up the following conversions:
- 29.

$$\left(\frac{cm}{sec}\right) \left(\frac{60\ seconds}{hour}\right) \left(\frac{m}{100\ cm}\right) \left(\frac{km}{100\ m}\right) \left(\frac{mile}{1.6\ km}\right) = \left(\frac{miles}{hour}\right)$$

30. Will these conversion factors give him the correct answer? Explain your reasoning.

3.3 Uncertainty in Measurements

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Distinguish between accuracy and precision in measurements.

1. A basketball player shoots ten free throws and makes every one of them. Her shooting is
 - a. accurate but not precise
 - b. precise but not accurate
 - c. both accurate and precise
 - d. neither accurate nor precise
2. Another basketball player shoots ten free throws. Some go off to the left of the basket, some go off to the right, a couple don't reach the basket and three go over the top of the backboard. This player's shooting is
 - a. accurate but not precise
 - b. precise but not accurate
 - c. both accurate and precise
 - d. neither accurate nor precise
3. A third basketball player shoots ten free throws. All the shots hit the edge of the backboard on the left side of the basket. This player's shooting is
 - a. accurate but not precise
 - b. precise but not accurate
 - c. both accurate and precise
 - d. neither accurate nor precise
4. True or False: Precision is a measure of how close the series of measurements are to one another.
5. True or False: A systematic error is never due to sloppy lab technique.
6. True or False: The best way to find the accepted value for the density of aluminum is to look it up in a table of densities.
7. _____ precision with low accuracy often results from a _____ error.
8. Scientists try to maximize both _____ and _____ in their measurements.

Lesson Objective: Calculate the percent error of a measured quantity.

9. The percent error is _____ if the accepted value and the experimental value are the same.
 - a. 100%
 - b. 0%
 - c. cannot be calculated
 - d. increased
10. The accepted value for the density of aluminum is _____ g/cm³.
 - a. 2.45

- b. 2.42
 - c. 2.87
 - d. 2.70
11. When calculating percent error, the denominator in the expression is the
- a. accepted value
 - b. experimental value
 - c. 100
 - d. experimental value – accepted value
12. True or False: The error of an experiment is the difference between the experimental and accepted values.
13. True or False: The percent error of a measurement decreases as the measurement accuracy increases.
14. Fill in the following table:

TABLE 3.3: Table

Experiment	Accepted Value	Experimental Value	Percent Error
a.	7.98	6.87	
b.	7.98	8.05	
c.	7.98		+0.13

Lesson Objective: Report measured values to the correct number of significant figures based on the measuring tool.

15. The significant figures in a measurement consist of all the certain digits in that measurement plus ___ uncertain or estimated digit.
- a. no
 - b. two
 - c. one
 - d. three
16. Numbers used in unit conversions always have _____ significant digits.
- a. one
 - b. infinite
 - c. negligible
 - d. two
17. Insignificant digits
- a. are not important
 - b. can be ignored
 - c. are factors of ten
 - d. are placeholders for decimals
18. True or False: Insignificant digits are never reported.
19. True or False: There are no firm rules about reporting significant digits.
20. True or False: All nonzero digits in a number are significant.
21. Fill in the following table:

TABLE 3.4: Table

	Number	How Many Significant Figures?	Rule Number
a.	32.01		
b.	1.98×10^3		
c.	84720		
d.	0.00356		

Lesson Objective: Perform calculations with measured quantities, rounding the answers to the correct number of significant figures.

22. In rounding numbers in addition and subtraction problems, use _____ in deciding how to round.
- the number of decimal places.
 - the rule of 5
 - the average of significant figures
 - the number of significant figures
23. True or False: When 5.67 is multiplied by 24, the answer will be rounded to two significant figures.
24. True or False: The rounding rules are the same for all four mathematical operations.
25. Give the answer (with the correct number of significant figures) for the following calculations:
- $24 \times 932 =$
 - $\frac{65012}{335.0} =$
 - $892 + 7.4 =$
 - $612.24 - 23.7 =$

TABLE 3.5: Table

Experiment	Accepted Value	Experimental Value	Percent Error
a.	7.98	6.87	13.9
b.	7.98	8.05	0.88
c.	7.98	8.11	+0.13

TABLE 3.6: Table

	Number	How Many Significant Figures?	Rule Number
a.	32.01	four	two
b.	1.98×10^3	three	one
c.	84720	four	four
d.	0.00356	three	three

CHAPTER **4**

Atomic Structure Worksheets

Chapter Outline

- 4.1 ATOMS
 - 4.2 THE NUCLEAR MODEL OF THE ATOM
 - 4.3 ISOTOPES AND ATOMIC MASS
-

4.1 Atoms

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Explain the law of conservation of mass, the law of definite proportions, and the law of multiple proportions.

- One of the following is not true about Democritus
 - he was a Greek philosopher
 - he suggested that atoms were the basic units of matter
 - his ideas were quickly accepted
 - his approach was philosophical, not experimental
- By the 1700s chemists had accepted the idea of
 - an element
 - atomic structure
 - existence of electrons
 - element sub-particles
- Carbon and oxygen can combine to form
 - only one compound
 - many different compounds
 - two different compounds
 - three different compounds
- Elements can combine to form
 - other elements
 - gold
 - reactants
 - compounds
- The law of definite proportions means that
 - all compounds have similar compositions
 - all samples of the same compound have the same composition
 - sample of a compound from different sources can have different compositions
 - a compound might have different amounts of atoms
- True or False: If 125 grams of reactants are used in a reaction, only 50 grams of product are expected.
- True or False: Some elements can combine to form more than one compound.
- True or False: The idea of the atom did not develop until the Renaissance.
- True or False: Atoms combine in small whole-number ratios to form compounds.
- True or False: Copper can form three different compounds with chlorine.
- An _____ is a substance that cannot be broken down into simpler substances by ordinary chemical means.
- Compounds are more _____ than the elements from which they are formed.

13. Carbon monoxide has a carbon/oxygen ratio of _____.
14. "Atomos" is a Greek word that means _____.
15. The ideas of Democritus were not useful in his day because there was no _____ evidence to support them.
16. H_2O (water) has a composition which is 2 grams H to 16 grams O. In H_2O_2 (hydrogen peroxide), the composition is 2 grams H to 32 grams O. What is the mass ratio for O in the two compounds?

Lesson Objective: Describe John Dalton's atomic theory.

17. John Dalton proposed that all matter was composed of
 - a. elements
 - b. atoms
 - c. compounds
 - d. particles
18. All atoms of a specific element must have all but one of the following
 - a. same size
 - b. same mass
 - c. same color
 - d. same reactivity
19. Chemical reactions do not involve the _____ of atoms.
 - a. transmutation
 - b. separation
 - c. combination
 - d. rearrangement
20. John Dalton was a chemist and a
 - a. farmer
 - b. college professor
 - c. schoolteacher
 - d. physician
21. Dalton's atomic theory was based on all of the following except
 - a. law of conservation of mass
 - b. law of definite proportions
 - c. law of multiple proportions
 - d. law of conservation of energy
22. True or False: Dalton proposed his atomic theory in 1808.
23. True or False: Atoms of one element may be the same as atoms of a different element.
24. True or False: Atoms of different elements can combine in fractional ratios to form compounds.
25. True or False: A scanning tunneling microscope is needed to see atoms.
26. True or False: Chemical reactions can change atoms of one element into atoms of a different element.
27. All _____ is composed of extremely small particles called atoms.
28. An _____ is the smallest particle of an element that retains the properties of that element.
29. Atoms of one element are _____ from atoms of another element.
30. Democritus was a Greek philosopher who lived from 460 B.C. until 370 B.C. He was among the first known individuals to suggest the idea of the atom as the basic unit of matter. The word atom comes from the Greek word *atomos*, meaning "indivisible". Democritus (Figure 1.1) was indeed ahead of his time, but his ideas were not useful at that time in describing chemical behavior because there was no experimental evidence to support it. His approach was a philosophical one rather than a truly scientific one. Many centuries would pass before the notion of atoms was merged with modern scientific experimentation and thought.

- a. Would the ideas of Democritus be considered a hypothesis or a theory?
- b. Why is Dalton's atomic theory still referred to as a theory?

4.2 The Nuclear Model of the Atom

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Distinguish between the three main subatomic particles.

- One of the following is not a main subatomic particle
 - proton
 - positron
 - electron
 - neutron
- True or False: Dalton assumed that atoms could not be broken down further.
- Investigations in the late _____ proved that atoms can indeed be broken down into smaller particles.

Lesson Objective: Understand the contributions of J. J. Thomson, Robert Millikan, and Ernest Rutherford to atomic theory.

- J.J. Thomson carried out experiments with
 - an X-ray tube
 - a test tube
 - a cathode ray tube
 - a fluorescent tube
- Cathode rays were shown to be streams of _____ particles.
 - positively charge particles
 - electrically neutral particles
 - a mixed of charged particles
 - negatively charged particles
- Millikan's oil drop experiment determined the mass of the
 - electron
 - neutron
 - proton
 - atomic nucleus
- Complete atoms are electrically
 - positive
 - negative
 - neutral
 - complex
- _____ discovered the proton in 1886.
 - Ernest Rutherford

- b. Eugene Goldstein
 - c. James Chadwick
 - d. Robert Millikan
9. True or False: The cathode ray was composed of particles.
 10. True or False: The cathode rays were deflected towards a negatively charged metal plate.
 11. True or False: The value for the charge on the electron depends on the gas used in the cathode ray tube.
 12. True or False: The neutron has approximately the same mass as the proton.
 13. True or False: The proton was first called a canal ray.
 14. The electron weighs about _____ the mass of the hydrogen atom.
 15. The _____ is the positively charged electrode in the cathode ray tube.
 16. The cathode ray deflection in a _____ field showed it was a _____ charged particle.

Lesson Objective: Describe the structure of the nuclear atom.

17. J.J. Thomson developed the _____ model of the atom
 - a. apple pie
 - b. rice pudding
 - c. cherry pie
 - d. plum pudding
18. Rutherford performed the _____ foil experiment to study atomic structure.
 - a. alpha
 - b. beta
 - c. silver
 - d. gold
19. Rutherford's experiment used _____ to bombard the foil.
 - a. alpha particles
 - b. beta particles
 - c. gamma particles
 - d. delta particles
20. The nucleus of the atom is composed of
 - a. protons and electrons
 - b. protons and neutrons
 - c. neutrons and electrons
 - d. protons and positrons
21. True or False: Rutherford felt that all the atom's positive charge was in the nucleus.
22. True or False: Rutherford's experiments showed the location of the electron.
23. True or False: In Rutherford's experiment, most of the particles passed through the gold foil.
24. True or False: None of the particles in Rutherford's experiment bounced back toward the source.

TABLE 4.1: Relative Charge and Mass

Particle	Relative Charge	Relative Mass
proton		
	0	
		1/1840

26. Rutherford's experiment showed that most of the atom was _____ .

27. The particles used in Rutherford's experiment were about _____ times the size of the hydrogen atom and were _____ charged.
28. Rutherford's atomic model became known as the nuclear model. In this model, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in a nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.
- What data from Rutherford's experiment suggested the atom was mostly empty space?
 - Describe Rutherford's model of the atom. Why was it incomplete?

TABLE 4.2: Relative Charge and Mass

Particle	Relative Charge	Relative Mass
proton	+1	1
neutron	0	1
electron	-1	1/1840

4.3 Isotopes and Atomic Mass

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define atomic number.

- The atomic number for an element indicates the number of
 - protons
 - neutrons
 - electrons
 - muons
- An atom can be classified as a particular element based on
 - atomic mass
 - number of neutrons
 - molar mass
 - number of protons
- True or False: The number of positively charged protons must equal the number of negatively charged electrons.
- True or False: All atoms of the same element must have the same atomic number.
- The atomic number for nickel (Ni) is _____ .
- Scandium (Sc) has _____ protons in its nucleus.

Lesson Objective: Define mass number.

- The mass number for an element is the
 - total number of subatomic particles
 - total number of protons and electrons
 - total number of protons and neutrons
 - total number of neutrons and electrons
- The number of neutrons in an atom equals
 - mass number – number of electrons
 - number of protons – mass number
 - number of protons
 - mass number – number of protons
- True or False: ${}^{42}_{20}\text{Ca}$ The subscript always stands for the number of neutrons in the atom.
- True or False: The number of neutrons will not always equal the number of electrons in the atom.
- True or False: Electrons contribute a large amount to the atomic mass of an element.

TABLE 4.3: Elements Table

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
chromium	Cr	24				51
sodium	Na		12	11		23
silicon	Si				14	30
indium	In			49		99

Lesson Objective: Understand how isotopes differ from one another and be able to designate them by various methods.

13. Isotopes of an element differ only by
 - a. number of electrons
 - b. number of neutrons
 - c. number of protons
 - d. number of neutrons + electrons
14. Chemical behavior is determined by the number of
 - a. protons + neutrons
 - b. protons + electrons
 - c. neutrons + electrons
 - d. protons alone
15. Only one of the following is an appropriate way to indicate an isotope of iodine (I) with an atomic mass of 131
 - a. I
 - b. ^{131}I
 - c. 131-I
 - d. I-131
16. Naturally occurring carbon has _____ isotopes.
 - a. one
 - b. four
 - c. three
 - d. two
17. True or False: All atoms of a given element are identical.
18. True or False: The presence of isotopes does not affect chemical reactivity for an element.
19. True or False: Isotopes of a given element have the same number of electrons.
20. True or False: Most elements do not have isotopes.
21. The term _____ refers to the nucleus of a given isotope of an element.
22. Not all atoms of a given element are _____ .

Lesson Objective: Be able to calculate the average atomic mass of an element.

23. An atomic mass unit is defined in relation to the _____ atom.
 - a. hydrogen
 - b. fluorine
 - c. carbon
 - d. nitrogen

24. Masses of individual atoms can be determined by
- mass spectrometer
 - atomic balance
 - mass fluorometer
 - fluorescence radiation
25. True or False: The reference standard for atomic masses is C-12.
26. Calculate the average atomic mass of the following elements:
27. ⁶
28. Li = 7.59%, 6.015 amu
29. ⁷
30. Li = 92.41% , 7.016 amu
31. ⁶³
32. Cu = 69.17% 62.29 amu
33. ⁶⁵
34. Cu = 30.83% 64.927 amu.
35. Atoms are the fundamental building blocks of all matter and are composed of protons, neutrons, and electrons. Because atoms are electrically neutral, the number of positively charged protons must be equal to the number of negatively charged electrons. One of Dalton's points in his atomic theory was that all atoms of a given element are identical. In this section, we will see how this is not strictly true, thanks to variability in the number of neutrons that an atom may contain.
36. How would the atomic theory be different if Dalton knew about neutrons?
37. Some isotopes are very unstable and will break down (decay). What do you think happens in a nuclear decay process?

TABLE 4.4: Elements Table

Element	Symbol	Protons	Neutrons	Electrons	Atomic Number	Atomic Mass
chromium	Cr	24	27	24	24	51
sodium	Na	11	12	11	11	23
silicon	Si	14	16	14	14	30
indium	In	49	50	49	49	99

CHAPTER **5**

Electrons in Atoms Worksheets

Chapter Outline

- 5.1 LIGHT
 - 5.2 THE QUANTUM MECHANICAL MODEL
 - 5.3 ELECTRON ARRANGEMENTS IN ATOMS
-

5.1 Light

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the relationships between speed, wavelength, and frequency of light.

- Light travels at a speed of _____ m/s
 - 2.998×10^5
 - 3×10^8
 - 3×10^6
 - 2.998×10^7
- One of the following is not true of a wave
 - a wave with a short wavelength has a high frequency
 - frequency is usually measured in cycles/second
 - the amplitude is the width of the wave
 - $c = \lambda\nu$
- True or False: Electromagnetic radiation is a form of energy that exhibits wavelike behavior.
- True or False: The wavelengths of visible light are longer than those of infrared light.
- Waves are characterized by their _____ motion.
- A _____ is characterized as the distance between any two corresponding points on adjacent waves.
- Blue light in the visible spectrum has a frequency of approximately 630 THertz (1 THertz = 10^{12} cycles/sec). What is the wavelength of this light?

Lesson Objective: Understand the photoelectric effect and how it is related to the wave-particle duality of light.

- Light from a heated body is emitted
 - at all wavelengths
 - at only one or two wavelengths
 - at certain specific wavelengths
 - in no discernible pattern
- A _____ is the minimum amount of energy that can be gained or lost by an electron.
 - qualim
 - quanta
 - quentum
 - quantum
- True or False: A photon has the mass of one electron.
- True or False: Einstein proposed that light could behave as a particle.
- Light shined onto a _____ surface causes the ejection of _____ from that metal.

13. The _____ frequency needed to cause the ejection of electrons is referred to as the _____ frequency.
14. What is the energy of a photon of light with a frequency of 690 THertz?
15. What is the frequency of a photon with a E of $3.51 \times 10^{-19}\text{J}$?

Lesson Objective: Describe how changes in electron energies lead to atomic emission spectra.

16. Energy levels of electrons in an atom are
 - a. equal
 - b. quantumized
 - c. continuous
 - d. quantized
17. There are _____ visible lines in the hydrogen atomic emission spectrum.
 - a. four
 - b. three
 - c. two
 - d. one
18. True or False: Electrons move to higher energy levels when they absorb energy.
19. True or False: Light emitted from electrons is composed of continuous wavelengths.
20. Light is emitted when electrons _____ to _____ energy levels.
21. White light viewed through a prism give a _____ spectrum.
22. Atomic emission spectra led a new _____ of the atom based on _____.

Lesson Objective: Describe the Bohr model of the atom.

23. The Balmer series of hydrogen emission lines appear in the _____ region.
 - a. infrared
 - b. visible
 - c. ultraviolet
 - d. far infrared
24. The Bohr model of the atom is sometimes referred to as the _____ model.
 - a. orbital
 - b. quantum
 - c. planetary
 - d. solar system
25. The Lyman series has _____ different possible electron transitions.
 - a. seven
 - b. three
 - c. four
 - d. six
26. True or False: All the lines of the hydrogen emission spectrum appear in the visible range.
27. True or False: The Bohr model of the atom only works for the hydrogen atom.
28. True or False: The Brackett series of hydrogen electron transitions drop electrons from higher energy levels to energy level 4.
29. The series of emission lines from transitions of hydrogen electrons to energy level three are called the _____ series.
30. The hydrogen emission lines indicate that the electrons were arranged into specific _____ levels.
31. Low frequency light (red) is unable to cause ejection of electrons from the metal surface. At or above the threshold frequency (green), electrons are ejected. Incoming light that has the same intensity but a higher

frequency (blue) causes the ejection of the same number of electrons, but the ejected electrons move at a greater speed.

32. Why would higher frequency light cause ejected electrons to move at a greater speed?
33. Why are there more emission lines for helium than for hydrogen? Why is the emission spectrum of iron so complex?

5.2 The Quantum Mechanical Model

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Understand the de Broglie wave equation and how it illustrates the wave nature of the electron.

- The de Broglie wave equation showed that the wavelength of any moving object is equal to
 - $\frac{h}{m}$
 - $\frac{h}{mv}$
 - hm
 - $\frac{m}{h}$
- One of the following statements will not be true if quantum mechanics are used in explanations of particle behavior
 - sub-atomic particles display quantum behavior
 - wave-particle duality must exist
 - some quantum behavior can be observed visually
 - discrete energy levels must be shown
- True or False: The Bohr model of the electron clearly explained why atomic energy states were quantized.
- True or False: Electrons can only exist in an atom at specific energies.
- Acceptance of the _____ nature of the electron allowed a new way of understanding how electrons functioned.
- The Heisenberg Uncertainty Principle states that the _____ and _____ of a particle cannot be determined simultaneously.
- In an eight-ball pool game, the cue ball can travel 5m/sec. If the cue ball has a mass of 160 g, what is the wavelength it demonstrates?

Lesson Objective: Explain the difference between quantum mechanics and classical mechanics.

- All of the following statements about classical mechanics are true except one.
 - all objects move with wavelike motion
 - objects gain or lose energy in any amount
 - the materials demonstrate wave-particle duality
 - size and mass of objects obscures any quantum effects
- The Heisenberg Uncertainty Principle states
 - velocity and position can both be determined at the same time
 - a beam of light can interfere with the path and speed of an electron
 - photons do not affect the position of an electron
 - the velocity of a car cannot accurately be measured by a radar beam
- True or False: Particles described by quantum theory can only lose or gain energy in discrete units.

11. True or False: The Heisenberg Uncertainty Principle says that we can never measure the position and velocity of the moon accurately.
12. Particles whose motions are better described by quantum mechanics can only gain or lose energy in discrete units called _____.
13. The study of motion in large objects is called _____ .

Lesson Objective: Understand how the Heisenberg uncertainty principle and Schrödinger's wave equation led to the idea of atomic orbitals.

14. Which of the following statements about the Schrödinger wave equation is not true?
 - a. the quantum mechanical model of the electron comes from this equation
 - b. this equation supports the Bohr model of electron orbits
 - c. solutions to this equation are called wave functions
 - d. quantization of electron energies is required to solve the equation
15. An electron cloud
 - a. represents a 50% probability of finding an electron
 - b. has uniform density
 - c. can be called an orbital
 - d. can only show where the electron is not located
16. True or False: An orbital represents a three-dimensional area of space.
17. True or False: Electrons travel around the nucleus in circular orbits.
18. The electron cloud represents an area of space where there is a _____% probability of finding an electron.
19. The _____ wave equation accurately described the behavior of the electron in a hydrogen atom.

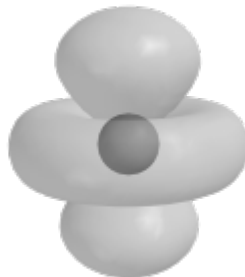
Lesson Objective: Know the four quantum numbers and how they are related to the arrangement of electrons in an atom.

20. Each electron can be described by _____ quantum numbers.
 - a. 4
 - b. 3
 - c. 2
 - d. 1
21. The quantum numbers describe all the characteristics of an electron except
 - a. velocity
 - b. magnetic spin
 - c. energy level
 - d. orbital shape
22. True or False: The principal quantum number describes the magnetic spin of an electron.
23. True or False: The magnetic quantum number describes the orientation of an orbital.
24. Each orbital can hold a maximum of _____ electrons.
25. The possible values for the spin quantum number are _____ and _____ .

Lesson Objective: Describe the interrelationships between principal energy level, sublevel, orbital and electron spin and how they relate to the number of electrons of an atom.

26. One of the following is not an allowable sublevel for $n = 3$
 - a. p
 - b. d

- c. f
 - d. s
27. For $n = 2$, the number of orbitals per principal energy level is
- a. 3
 - b. 6
 - c. 1
 - d. 2
28. True or False: The p orbitals are dumbbell-shaped.
29. True or False: All s orbitals have complex shapes.
30. The following image represents a _____ orbital.



- 31.
32. The total number of allowable orbitals for $n = 4$ is _____.
33. The Heisenberg Uncertainty Principle states that it is impossible to determine simultaneously both the position and the velocity of a particle. The detection of an electron, for example, would be made by way of its interaction with photons of light. Since photons and electrons have nearly the same energy, any attempt to locate an electron with a photon will knock the electron off course, resulting in uncertainty about where the electron is headed.
34. A running back collides with a defensive tackle and is knocked down. Explain how the Heisenberg Uncertainty Principle applies to this situation.
35. Do a unit analysis of the de Broglie equation to show how an increase in mass affects the wavelength of the moving object.

5.3 Electron Arrangements in Atoms

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

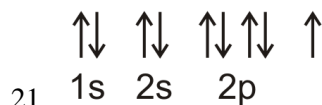
Lesson Objective: Understand how to apply the Aufbau principle, the Pauli exclusion principle, and Hund's rule to determine ground state electron configurations.

- An atom's electron configuration describes
 - the shape of the electrons
 - the arrangement of electrons
 - the shape of electrons
 - the shape of the orbitals
- The Aufbau principle states
 - lower orbitals are filled last
 - orbitals are filled in random order
 - lower orbitals are filled before upper orbitals
 - no lower energy orbitals are filled
- The Pauli exclusion principle states that
 - all electrons have four quantum numbers
 - quantum numbers are excluded for some elements
 - two electrons in an element may have identical quantum numbers
 - no two electrons in an element can have the same four quantum numbers
- Hund's rule states that
 - orbitals of equal energy must all have one electron before further filling.
 - all orbitals must first fill with one electron
 - energy levels will change as orbitals are filled
 - orbital filling is a random process
- True or False: Each element has a unique electron configuration.
- True or False: The 4f orbital fills just after the 4d orbital.
- True or False: There is one element whose electrons all have the same spin quantum number.
- True or False: Carbon has one empty 2p orbital.
- True or False: Nitrogen has one 2p orbital that contains two electrons.
- The 2p orbital must be filled before the ____ orbital can be filled.
- The ____ orbital is filled just before the 3p orbital is allowed to fill.
- Each orbital can contain up to ____ electrons.
- Boron has ____ empty 2p orbitals.

Lesson Objective: Be able to write correct orbital filling diagrams and electron configurations for all elements.

- The arrows in orbital filling diagrams indicate

- orbital rotation of the electrons
 - electron spin
 - magnetic rotation
 - orientation of orbital
- Electron configuration superscripts indicate all of the following except
 - number of electrons
 - number of protons
 - atomic number
 - orbital shapes
 - The electron configuration $1s^22s^22p$ is the electron configuration for
 - carbon
 - oxygen
 - boron
 - beryllium
 - True or False: The electron configuration for lithium is $1s^22s^2$.
 - True or False: Oxygen will have one paired set of 2p electrons.
 - True or False: The orbital filling diagram for He is $\uparrow\uparrow$.
 - The orbital filling diagram for sodium (atomic number 11) is _____.



- is the orbital filling diagram for _____.
- The electron configuration for Al is _____.
 - $1s^22s^22p^6$ is the electron configuration for _____.

Lesson Objective: Know how to use the noble gas notation shorthand method.

- The noble gas shorthand method
 - indicates reactivity
 - requires writing all the electrons in the diagram
 - allows abbreviation of the electron configuration
 - shows relationships among the noble gases
- True or False: The noble gas configuration for calcium is $[\text{Kr}]4s^2$.

TABLE 5.1:

Element	Atomic Number	Electron Configuration
Ni		$[\text{Ar}]3d^84s^2$
Mg	12	
Be	4	

Lesson Objective: Be able to determine the number of valence electrons and the number of unpaired electrons in any atom.

- Valence electrons are
 - unpaired electrons in the outermost principal energy level
 - all electrons in the outermost orbital
 - electrons in the outermost principal energy level

- d. electrons with paired spin in the outermost principal energy level
28. *Note: You may use the electron configuration tables for questions 28-30.*
29. Sc (atomic number 21) has _____ valence electrons.
30. Chlorine (atomic number 17) would be expected to have _____ unpaired electrons.
31. Nitrogen (atomic number 7) has _____ valence electrons and _____ unpaired electrons.

Lesson Objective: Understand that some electron configurations are exceptions to the normal Aufbau process.

31. The first element to show an exception to the Aufbau process is
- Cr
 - Pt
 - Nb
 - Cs
32. The last of the three rules for determining ground state electron configurations gives information on how to arrange electrons in a set of orbitals that are all within the same sublevel. Hund's rule states that orbitals of equal energy are each occupied by one electron before any orbital is occupied by a second electron, and all of the unpaired electrons must have the same spin. A slight preference for keeping electrons in separate orbitals helps minimize the natural repulsive forces that exist between any two electrons.
33. Why do all unpaired electrons need to have the same spin?
34. Develop a method using only the electron configuration tables for determining how many unpaired electrons a given atom has. Confine your discussion to elements 1-20.

TABLE 5.2:

Element	Atomic Number	Electron Configuration
Ni	28	$[\text{Ar}]3d^84s^2$
Mg	12	$[\text{Ne}]3s^2$
Be	4	$1s^22s^2$

CHAPTER **6**

The Periodic Table Worksheets

Chapter Outline

- 6.1 HISTORY OF THE PERIODIC TABLE
 - 6.2 ELECTRON CONFIGURATION AND THE PERIODIC TABLE
 - 6.3 PERIODIC TRENDS
-

6.1 History of the Periodic Table

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe some of the early attempts to organize the chemical elements.

Multiple Choice

- One of the following metals was known since ancient times.
 - gallium
 - tungsten
 - copper
 - cadmium
- The triad system was based on all of the following except
 - similar chemical reactivity
 - equal numbers of protons in the nucleus
 - pattern of atomic masses
 - reactivity with other triads
- True or False: John Newlands developed a system for organizing elements known as the Law of Octaves.
- True or False: Zinc (Zn) was discovered between 1844-1886.
- Newland noticed that every _____ element exhibited similar properties.
- Newland's law did not seem to hold for elements heavier than _____.

Lesson Objective: Understand how Mendeleev organized his periodic table.

- Mendeleev noticed a trend in elements when organized in order of
 - increasing atomic mass
 - increasing atomic number
 - increasing electron number
 - increasing reactivity
- Mendeleev placed _____ before iodine based on chemical properties.
 - bromine
 - selenium
 - tellurium
 - sulfur
- True or False: There are three missing elements between zinc and arsenic.
- True or False: Lothar Meyer published the first periodic table.
- Similarities in chemical behavior show a pattern known as _____ .

12. Mendeleev arranged his table so that atomic mass increases from _____ to _____ .
13. The element _____ was discovered four years after Mendeleev published his table.

Lesson Objective: Know the improvements that Moseley made to Mendeleev's table.

14. Moseley based his periodic table modifications on data obtained from
 - a. X-ray spectra
 - b. atomic emission spectra
 - c. atomic absorption spectra
 - d. atomic X-ray spectra
15. Moseley defined atomic number as
 - a. number of neutrons in the nucleus
 - b. number of neutrons plus protons
 - c. number of protons
 - d. number of electrons
16. True or False: Rutherford was the first to propose modifications of Mendeleev's periodic table.
17. True or False: Moseley used atomic number for his modifications of the periodic table.
18. Moseley defined atomic _____ as the number of protons in the nucleus of an atom.

Lesson Objective: Know the periodic law.

19. Tellurium is placed before iodine in the periodic table because
 - a. it has a larger atomic mass
 - b. it has more protons
 - c. it has a lower atomic number
 - d. it has more electrons
20. True or False: The periodic law describes the arrangement of elements based on atomic number.
21. When elements are arranged in order of _____ atomic number, there is a _____ repetition of their _____ and _____ properties.

Lesson Objective: Describe various components of the modern periodic table, including periods, groups, metals, nonmetals, and metalloids.

22. Some chemical symbols are based on _____ for the elements.
 - a. Latin names
 - b. trivial names
 - c. Greek names
 - d. Chinese names
23. A period is a
 - a. horizontal column
 - b. vertical column
 - c. vertical row
 - d. horizontal row
24. The current numbering system for groups was developed by the
 - a. American Chemical Society

- b. European Union of Scientists
 - c. International Union of Pure and Applied Chemistry
 - d. International Association of Chemistry
25. True or False: A metal is a good conductor of heat and electricity.
26. True or False: Aluminum is considered a metalloid.
27. True or False: Nonmetals can exist in all three states of matter at room temperature.
28. A _____ is generally a poor conductor of heat and electricity.
29. A _____ has properties _____ between those of a metal and a nonmetal.
30. A _____ is a horizontal row of the periodic table.
31. A group is a vertical _____ of the periodic table.
32. In the solid state _____ are brittle.
33. Approximately _____% of the elements are metals.
34. English chemist John Newlands (1838-1898) arranged the elements in increasing order of atomic mass and noticed that every eighth element exhibited similar properties. He called this relationship the *Law of Octaves*. Unfortunately, some elements were missing, and the law did not seem to hold for elements that were heavier than calcium. Newlands' work was largely ignored and even ridiculed by the scientific community in his day. It was not until years later that another more extensive effort to construct a periodic table would gain much greater acceptance, and the pioneering work of John Newlands would be appreciated.
35. Why was Newland ridiculed by his scientific colleagues?
36. Why is the organization of the periodic table by physical properties not very useful?

6.2 Electron Configuration and the Periodic Table

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Understand the relationship between the number of orbitals in various energy sublevels and the length of the periods in the periodic table.

- Period 3 has _____ elements
 - 6
 - 8
 - 18
 - 32
- The 6p orbitals begin in period _____.
 - 2
 - 4
 - 6
 - 8
- True or False: All the seventh period elements have been found in nature.
- True or False: 3s and 3p sublevels are being filled in the third period.
- The electron configuration for barium is $[\text{Xe}]6s^2$. Barium will be found in the _____ period.
- An element is found in period four. Is this element Be (atomic number 2) or Mn (atomic number 25)? The element is _____.

Lesson Objective: Identify each block of the periodic table and be able to determine which block each element belongs to based on its electron configuration.

- The d sublevel is always _____ behind the period in which the sublevel occurs.
 - two principal sublevels
 - one principal level
 - four principal sublevels
 - three principal sublevels
- The 4f sublevel belongs in the _____ period.
 - second
 - fourth
 - sixth
 - eighth
- True or False: With one exception, all the s sublevels being filled are in groups 1 and 2.

10. True or False: The d block is comprised of elements in groups 4-9.
11. Elements in groups 13-18 are primarily filling ____ orbitals.
12. The f orbitals are primarily filled in periods ____ and ____ .

Lesson Objective: Describe the relationship between outer electron configuration and group number. Be able to determine the number of valence electrons for any element.

13. The number of valence electrons in p-block elements is
 - a. group number - 6
 - b. group number - 8
 - c. group number - 10
 - d. group number - 12
14. Alkaline earth metals have ____ valence electrons.
 - a. 1
 - b. 2
 - c. 3
 - d. 4
15. True or False: Strontium has the electron configuration $[\text{Kr}]5s^2$, giving it two valence electrons.
16. True or False: The noble gases all have s^2p^5 electron configurations.
17. Polonium has the electron configuration $[\text{Hg}]6p^4$, which gives it ____ valence electrons.
18. O is in group ____ and has six valence electrons.
19. Iron has the electron configuration $[\text{Ar}]3d^64s^2$, putting it in group ____

Lesson Objective: Locate the following groups on the periodic table: alkali metals, alkaline earth metals, halogens, and noble gases.

20. The halogens are found in group ____ .
 - a. fifteen
 - b. sixteen
 - c. seventeen
 - d. eighteen
21. The period 4 group 1 element is
 - a. Li
 - b. K
 - c. Ca
 - d. Rb
22. True or False: Radon is a noble gas.
23. True or False: Atomic number 34 is a halogen.
24. Radium is located in group two and is considered to be an _____ metal.
25. Cesium is in group 1 and has ____ electron(s) in its outer orbital.
26. Xenon (atomic number 54) is classified as group ____ element.
27. The transition element ${}_{28}\text{Ni}$ is in group
 - a. 10
 - b. 9
 - c. 8
 - d. 7

28. The actinides are all located in group _____ .
- 1
 - 2
 - 3
 - 4
29. True or False: Tungsten (atomic number) 74 is a transition element.
30. True or False: The lanthanide element samarium (atomic number 62) is found in group 4.

TABLE 6.1: Table

Element	Atomic Number	Group	Period	Category
Palladium (Pd)	46			transition
Neodymium (Nd)				
	92			

32. Looking at the block diagram, you may be wondering why hydrogen and helium were not included in our discussion of the alkali metal and alkaline earth metal groups. Though hydrogen, with its $1s^1$ configuration, appears as though it should be similar to the rest of Group 1, it does not share the properties of that group. Hydrogen is a unique element that cannot be reasonably included in any single group of the periodic table. Some periodic tables even separate hydrogen's square from the rest of Group 1 to indicate its solitary status.
33. Do hydrogen and helium have anything in common?
34. How do we know where to put astatine in the periodic table if most of its properties are unknown?

TABLE 6.2: Table

Element	Atomic Number	Group	Period	Category
Palladium (Pd)	46	10	5	transition
Neodymium (Nd)	60	3	6	lanthanide
Uranium (U)	92	3	7	actinide

6.3 Periodic Trends

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Learn the periodic trends for atomic radius.

- The atomic radius is determined by measuring the
 - circumference of the atomic cloud
 - radius of the atomic cloud
 - distance between two nuclei
 - distance from nucleus to outer edge of cloud
- The atomic radius increases when
 - a new principal energy level is added
 - the atomic number increases within a group
 - more protons are added to the nucleus
 - electrons are lost
- True or False: Ionic radius increases from right to left in a period.
- True or False: Ionic radius decreases from top to bottom in a group.
- The distance between two H atoms is 74 pm. The ionic radius is _____ pm.
- The abbreviation pm stands for _____, which is 1×10 _____ meters.

Lesson Objective: Know the relationship between group number and valence electrons.

- The maximum number of valence electrons for an atom in its ground state is
 - 10
 - 8
 - 6
 - 4
- The outer electron configuration for group 16 is
 - ns^2
 - ns^2np^6
 - ns^2np^4
 - ns^2np^1
- True or False: Valence electrons are those in the outermost occupied principal energy level.
- True or False: Group number 13 has three valence electrons.
- Valence electrons will always be in the ____ and ____ orbitals.
- The number of valence electrons and the outer electron configuration is constant within a _____.

Lesson Objective: Describe how ions are formed.

13. Ions are formed by the gain or loss of
 - a. protons
 - b. electrons
 - c. neutrons
 - d. protons + electrons
14. Names of monoatomic anions end with
 - a. id
 - b. ic
 - c. ide
 - d. ied
15. True or False: The proper superscript for an ion with a single unit of charge is +1.
16. True or False: When an atom gains one or more electrons, it becomes negatively charged.
17. A positively charged ion is called _____ .
18. A negatively charged ion is called _____ .

Lesson Objective: Learn the periodic trends for ionization energy.

19. Ionization energies are measured in
 - a. kJ/mol
 - b. $J \times sec/kg$
 - c. J/mol
 - d. $kJ/gram$
20. True or False: Ionization energy tend to decrease from left to right across a period.
21. Ionization energy is the energy required to _____ an _____ .
22. An _____ in ionization energy results from an increase in nuclear charge.

Lesson Objective: Explain how multiple ionization energies are related to noble gas electron configurations.

23. Why is the second ionization energy for sodium so much higher than its first ionization energy?

Lesson Objective: Describe electron affinity.

24. The elements of the _____ group gain electrons most readily.
 - a. alkali metals
 - b. alkaline earth metals
 - c. transition elements
 - d. halogens
25. True or False: Electron affinities increase from top to bottom down a group
26. When energy is released in a chemical reaction or process, the energy is expressed as a _____ number.
27. Addition of one more electron to a halogen results in a _____ electron configuration.

Lesson Objective: Predict the effect that ion formation has on the size of an atom.

28. Removal of valence electrons from an atom results in
 - a. increase in ionic radius

- b. increase in occupied principal energy levels
- c. decrease in ionic radius
- d. formation of anions

29. True or False: The radius of selenium increases when it forms an ion.

Lesson Objective: Learn the periodic trends for electronegativity.

30. The most electronegative element is

- a. O
- b. F
- c. Br
- d. Si

31. True or False: Electronegativity deals with the ability to repel electrons.

32. The least electronegative element is _____.

33. Another anomaly can be found between Groups 15 and 16. Atoms of Group 16 (O, S, etc.) often have lower ionization energies than atoms of Group 15 (N, P, etc.). This can be explained in terms of Hund's rule. In a nitrogen atom, there are three unpaired electrons in the 2p sublevel. In an oxygen atom, there are four electrons in the 2p sublevel, so one orbital must contain a pair of electrons. One electron from this pair will be removed first during the ionization of an oxygen atom. Since electrons repel each other, it is slightly easier to remove an electron from a doubly occupied orbital than to remove an unpaired electron. Consequently, removing a paired electron from oxygen requires slightly less energy than removing an unpaired electron from the nitrogen atom.

34. Why is it easier to remove an electron from a paired orbital?

35. What problems could exist with the current way of determining atomic radii?

CHAPTER **7**

Chemical Nomenclature Worksheets

Chapter Outline

- 7.1 IONIC COMPOUNDS
 - 7.2 MOLECULAR COMPOUNDS
 - 7.3 ACIDS AND BASES
-

7.1 Ionic Compounds

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Distinguish between inorganic and organic chemistry.

1. Naming of organic compounds is
 1. easy because there are so few of them
 2. complex because there are so many of them
 3. based on exceptions to the rules
 4. the same as rules for inorganic compounds
2. True/False: All compounds containing carbon are organic compounds.

Lesson Objective: Interpret a molecular formula.

3. One of the following statements about the molecule shown below is incorrect $\text{CH}_3\text{CH}_2\text{CH}_2\text{OCH}_2\text{COOH}$
 1. the molecule contains five carbons
 2. the molecule contains ten hydrogens
 3. the molecule contains four oxygens
 4. the molecule contains eighteen atoms
4. True/False: If there is no subscript below a symbol, that means there is only one of that element being represented.
5. The molecule serotonin ($\text{C}_{10}\text{H}_{12}\text{N}_2\text{O}$) contains ____ nitrogen atoms and ____ hydrogen atoms.

Lesson Objective: Explain why an ionic compound is represented by an empirical formula.

6. An empirical formula represents
 1. the three-dimensional structure of the molecule.
 2. the small whole-number ratio among the atoms
 3. the actual number of atoms in the three-dimensional array
 4. the crystal structure made up of the different atoms
7. True/False: Ionic compounds exist as discrete molecular units.

Lesson Objective: Be able to determine the charges of monatomic ions formed by the representative elements from the position of each element on the periodic table.

8. Monatomic cations are formed by
 1. adding electrons
 2. losing protons
 3. losing electrons

4. adding protons
9. True/False: A monatomic anion is an atom with a -2 charge.
10. Fill in the missing information in the table below:

TABLE 7.1: Table

Atom	Ion	Name
Na	Na ⁺	sodium
Sr		
		bromide
	Ga ³⁺	
S		

Lesson Objective: Use the Stock system to identify the charge of transition metal ions.

11. The Stock system uses _____ to indicate ionic charge
 1. Roman numerals
 2. + and - to show cations and anions
 3. -ic and -ous to show charge
 4. Arabic numerals
12. True/False: The Stock system is only used for cations.
13. True/False: Cd(II) is used to designate a cadmium 2+ compound.
14. Fill in the missing information in the table below:

TABLE 7.2: Table

Element	Ion	Stock Name
Fe	Fe ³⁺	iron(III)
	Ni ²⁺	
Ag		
		chromium(II)
	Sn ⁴⁺	

Lesson Objective: Name an ionic compound given its formula.

15. A binary compound
 1. contains two types of atoms
 2. contains atoms combined in ratios of two
 3. can be a compound made of two metals
 4. can be a compound made of two nonmetals
16. The Stock system is used in naming
 1. all compounds containing cations
 2. only group 2 cations
 3. compounds containing transition metals
 4. compounds containing transition nonmetals
17. True/False: The overall charge of a binary ionic compound can vary.
18. Name the following compounds:

1. KBr
2. CaI₂
3. Na₂S
4. CoCl₂
5. Pt₃P₂
6. (NH₄)₃PO₄

Lesson Objective: Write the correct formula for an ionic compound given its name.

19. The criss-cross method is useful for
 1. balancing equations
 2. assigning charges
 3. determining the ratio of ions
 4. naming complicated compounds
20. The proper formula for potassium oxide is
 1. K₂O
 2. KO
 3. KO₂
 4. K₂O₃
21. The correct name for the CN⁻ ion is
 1. cyanate
 2. cyanous
 3. cyanic
 4. cyanide
22. True/False: The CO₃²⁻ ion is known as the bicarbonate ion.
23. True/False: A polyatomic ion is composed of more than one atom.
24. True/False: FeCl₃ is a ternary ionic compound
25. Write the formulas for the following compounds:
 1. copper(II) oxide
 2. magnesium bromide
 3. potassium dichromate
 4. magnesium phosphate
 5. iron(III) bromate

TABLE 7.3: Table

Atom	Ion	Name
Na	Na ⁺	sodium
Sr	Sr ²⁺	strontium
Br	Br ⁻	bromide
Ga	Ga ³⁺	gallium
S	S ²⁻	sulfide

TABLE 7.4: Table

Element	Ion	Stock Name
Fe	Fe ³⁺	iron(III)

TABLE 7.4: (continued)

Element	Ion	Stock Name
Ni	Ni ²⁺	nickel(II)
Ag	Ag ⁺	no Stock number
Cr	Cr ²⁺	chromium(II)
Sn	Sn ⁴⁺	tin(IV)

7.2 Molecular Compounds

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the difference between an ionic compound and a molecular compound.

1. One of the following statements about ionic compounds is not correct.
 1. ionic compounds are represented by empirical formulas
 2. ionic compounds take the form of individual molecules
 3. ionic compounds exists in extended arrays
 4. ionic compounds are made up of alternating positive and negative ions
2. One of the following statements about molecular compounds is not true
 1. molecular compounds share valence electrons
 2. molecular compounds form bonds between pairs of atoms
 3. molecular compounds lose electrons to nonmetal atoms
 4. all the molecules in a sample of a given molecular compound are identical
3. True/False: Carbon dioxide consists of a central carbon atom bonded to two oxygen atoms.
4. True/False: Water is an ionic compound.

Lesson Objective: Name a molecular compound given its formula.

5. Binary molecular compounds are usually composed of
 1. two metals
 2. a metal and a non-metal
 3. two nonmetals
 4. three or more different elements
6. The prefix _____ could be used to indicate a molecule with eight carbon atoms
 1. penta
 2. hexa
 3. hepta
 4. octa
7. The compound ethane pentabromide would contain _____ bromine atoms.
 1. five
 2. three
 3. seven
 4. four
8. The nonmetal ____ would follow S when writing binary compound formulas
 1. Cl
 2. I

3. H
4. P
9. True/False: P_2O_5 is named phosphorus pentoxide.
10. True/False: N and O can combine with each other in a variety of ratios.
11. True/False: CO would be named carbon monoxide.
12. True/False: CCl_4 is the formula for carbon trichloride.
13. Name the following compounds
 1. NO
 2. N_2O_5
 3. H_2O (the technical name)
 4. C_2F_6
 5. SiO_2
 6. NH_3 (the technical name)

Lesson Objective: Write the correct formula for a molecular compound given its name.

14. The formula for xenon hexachloride is
 1. $XeCl_4$
 2. $XeCl_6$
 3. $XeCl_8$
 4. $XeCl_5$
15. The formula for carbon tetrachloride is
 1. CCl
 2. CCl_2
 3. CCl_3
 4. CCl_4
16. Boron trihydride has the following formula
 1. BH_3
 2. B_2H_3
 3. B_2H_6
 4. B_3H_6
17. True/False: The formula for dinitrogen tetroxide is N_2O_2
18. True/False: The formula for sulfur trioxide is SO_3 .
19. True/False: The formula for selenium tetrahydride is SeH_5 .
20. Write formulas for the following compounds
 1. bromochloride
 2. arsenic triiodide
 3. disilicon tetroxide
 4. tellurium difluoride
 5. diboron hexahydride.
 6. triphosphorus octoxide
21. Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O . A molecule of octane, which is a component of gasoline, contains eight atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18} .
22. Does the name octane follow the rules that are discussed in this section? Why would the shorter name be used?
23. Would you expect ionic compounds and molecular compounds to react differently? Explain your answer.

7.3 Acids and Bases

Worksheet

{{Teachers lbody= Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define acids and bases in terms of the ions that are produced when each type of compound is dissolved in water.

- An acid produces _____ when dissolved in water.
 - OH^-
 - hydrogen ions
 - electrons
 - hydroxide ions
- An acid contains one or more _____ atoms.
 - oxygen
 - halogen
 - hydrogen
 - chlorine
- A base produces _____ when dissolved in water.
 - H^+
 - Cl^-
 - hydronium ions
 - hydroxide ions
- True/False: The bonds in an acid are ionic bonds.
- True/False: A base is an ionic compound.
- True/False: A binary acid consists of hydrogen and two other elements.
- Label each compound as acid (A), base (B), or other (O).
 - HCl
 - CH_3CH_3
 - KOH
 - HOH
 - HNO_3

Lesson Objective: Be able to name acids when given their chemical formulas.

- One of the following suffixes is not used in naming acids
 - ite
 - ote
 - ate
 - ide

9. Which of the following is the correct name for H_2CO_3 ?
1. carbonic acid
 2. carbolic acid
 3. carbonatic acid
 4. hydrocarbonic acid
10. What is the name for the PO_4^{3-} ion?
1. phosphite
 2. phosphorous
 3. phosphate
 4. phosphoric
11. True/False: The suffix of the anion determines how the acid is named.
12. True/False: All monoatomic anions end in *ite*.
13. True/False: The name of an acid is determined by the anion.
14. Name the following acids:
1. HCN
 2. HNO_3
 3. HMnO_4
 4. HClO_3
 5. H_2CrO_4

Lesson Objective: Write the correct formula for an acid when given its name.

15. The acid prefix for an anion ending in *ide* is
1. hypo-
 2. hydric-
 3. hydro-
 4. hydra-
16. If the anion ends in *ite*, the acid name ends in _____.
1. ic acid
 2. ate acid
 3. ite acid
 4. ous acid
17. True/False: The names of all monoatomic anions end in *-ite*.
18. True/False: NO_2^- is the nitrite anion.
19. Write the formulas for the following acids:
1. phosphoric acid
 2. nitrous acid
 3. hydroiodic acid
 4. arsenic acid
 5. perchloric acid

Lesson Objective: Name and write formulas for bases.

20. Bases exist as
1. extended three-dimensional arrays
 2. molecular compounds
 3. discrete molecules
 4. molecular arrays

21. The names of all bases end in
 1. hydroxyl
 2. hydroxide
 3. hydride
 4. hydronium
22. True/False: OH^+ is the hydroxide ion.
23. True/False: Hydroxide ions are formed from oxygen and hydrogen atoms that have bonded together.
24. Name the following bases
 1. $\text{Ca}(\text{OH})_2$
 2. LiOH
 3. $\text{Cu}(\text{OH})_2$
 4. $\text{Fe}(\text{OH})_3$
25. Write the formulas for the following bases
 1. potassium hydroxide
 2. iron(II) hydroxide
 3. magnesium hydroxide
 4. lead (IV) hydroxide
26. Acids and bases are two very important classes of chemicals that you probably use every day. Many foods and beverages contain acids, such as citrus fruits and juices, soda, and vinegar. Bases are very prevalent in cleaners such as ammonia, and most drain cleaners use a strong base as their active ingredient.
27. How does our knowledge of chemistry affect our understanding of acids and bases?
28. Why do we need to develop special naming for acids instead of just naming them as cations?

CHAPTER

8

Ionic and Metallic Bonding Worksheets

Chapter Outline

- 8.1 IONS
 - 8.2 IONIC BONDS AND IONIC COMPOUNDS
 - 8.3 METALLIC BONDS
-

8.1 Ions

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Be able to determine the number of valence electrons for any element and draw an electron dot diagram for any atom.

- The number of valence electrons for p-block elements is
 - group number minus 8
 - group number minus 10
 - group number minus 12
 - group number minus 6
- An electron dot diagram for an element shows
 - all electrons
 - p-sublevel electrons
 - valence electrons
 - bondable electrons
- True/False: Beryllium has two valence electrons.
- True/False: The number of valence electrons for carbon is three.
- Draw the electron dot diagram for
 - calcium
 - phosphorus
 - xenon
 - potassium

Lesson Objective: Use the octet rule to predict the charges of the most common ions formed by the representative elements.

- In order to achieve a full octet, the silicon atom will most likely
 - lose two electrons
 - lose four electrons
 - gain two electrons
 - gain four electrons
- Atoms of _____ elements tend to lose valence electrons to satisfy the octet rule.
 - gases
 - halogens
 - metals
 - noble gases
- True/False: Gilbert Lewis first proposed the octet rule.

9. True/False: All elements obtain eight valence electrons, either by electron loss or electron gain.
10. For each of the following elements, indicate whether they will gain or lose electrons to satisfy the octet rule:
1. cesium
 2. arsenic
 3. astatine
 4. radium
 5. magnesium

Lesson Objective: Write electron configurations for ions.

11. Cations are formed when
1. atoms lose an electron
 2. atoms lose all their valence electrons
 3. atoms add one electron
 4. atoms add electrons to fill the valence level
12. Phosphorus has the electron configuration of $[\text{Ne}]3s^2p^3$. This atom will most likely achieve a complete outer principal energy level by
1. losing 3 electrons
 2. losing 5 electrons
 3. gaining 1 electron
 4. gaining 3 electrons
13. True/False: The halogens each need to add one electron to achieve an outermost principal energy level.
14. True/False: Beryllium can achieve a complete outermost principal energy level by losing four electrons.
15. Write electron configurations for the following ions:
1. carbon 4+
 2. magnesium 2+
 3. oxygen 2-
 4. chloride 1-
 5. manganese 2+

Lesson Objective: Identify other atoms or ions that are isoelectronic with a particular ion.

16. The Sr^{2+} ion is isoelectronic with
1. krypton
 2. argon
 3. xenon
 4. rubidium
17. Isoelectronic ions have the following characteristic
1. they react the same
 2. they have the same number of electrons
 3. they have the same electron configuration
 4. they have the same number of protons
18. True/False: Na^+ and Cl^- are isoelectronic.
19. True/False: Mg^{2+} and S^{2-} are isoelectronic.
20. The Ba^{2+} ion is isoelectronic with the 1- _____ ion
21. Which cation would be isoelectronic with O^{2-} ? _____

Lesson Objective: Know that transition metal ions with either half-filled or completely filled *d* sublevels are particularly stable.

22. Transition metals form 2+ ions by
 1. losing s electrons
 2. filling incomplete d levels
 3. gaining p electrons
 4. losing d electrons
23. The pseudo noble gas electron configuration for some transition elements has _____ electrons in the outermost principal energy level.
 1. 12
 2. 8
 3. 10
 4. 18
24. True/False: Zinc and copper ions have full d sublevels.
25. True/False: Scandium achieves a noble gas electron configuration by losing its two 4s electrons.
26. Explain why the following ions would be particularly stable.
27. a. Mn
28. $^{2+}$
29. b. Cd
30. $^{2+}$
31. There are two ways in which atoms can satisfy the octet rule. One way is by sharing their valence electrons with other atoms, which will be covered in the next chapter, *Covalent Bonding*. The second way is by transferring valence electrons from one atom to another. Atoms of metallic elements tend to lose all of their valence electrons, which leaves them with an octet from the next lowest principal energy level. Atoms of nonmetallic elements tend to gain electrons in order to fill their outermost principal energy level with an octet.
32. Why would metallic elements tend to lose electrons and nonmetallic elements gain electrons to form ions?
33. What is the benefit of having a rule such as the octet rule?

8.2 Ionic Bonds and Ionic Compounds

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Explain how an ionic bond is formed when electrons are transferred from one atom to another in terms of the resulting electrostatic attraction. Draw diagrams showing this process.

1. The strength of an ionic bond is directly dependent on
 1. the size of the ion
 2. the magnitude of the charges
 3. the distance between the ions
 4. the ionic polarities
2. In describing ions, the cation is shown without dots because
 1. electrons have filled the outer principal energy level
 2. the valence electrons have added more electrons to make a full octet
 3. the energy level that contained the original valence electrons is now empty.
 4. no change has taken place in the electron distribution pattern
3. In order for sulfur to form a cation. It needs to
 1. lose two protons
 2. lose four electrons
 3. gain two electrons
 4. gain one electron
4. True/False: A larger ion will form a stronger bond than a smaller ion.
5. True/False: The chlorine atom gains one electron to become an anion.
6. True/False: Calcium needs to lose two valence electrons to form an ion.
7. Use diagrams to show the formation of the following ion pairs:
8. a. K and Br
9. b. Mg and O

Lesson Objective: Describe the structural arrangements of ions in a crystal, including coordination number and its relationship to a given compound's formula unit.

8. The formula unit for an ionic compound is
 1. an indication of the number of individual ions in the crystal
 2. an empirical formula
 3. a molecular formula
 4. a diagram of the crystal structure
9. 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
10. Cs and Na have different coordination numbers because
1. the ion sizes are very different
 2. the ionic charges are different
 3. the chloride ion changes shape in contact with the two cations
 4. solubilities of the two compounds are different
11. True/False: One formula unit of FeCl_3 contains three Cl^- for every Fe^{3+} .
12. True/False: Coordination numbers are determined by ionic combinations.
13. True/False: All group one cations will have the same packing arrangement.
14. What type of bonds exist between ions in an ionic compound?
15. Would you expect rubidium chloride to have a coordination number more like Na or more like Cs? Explain your answer.

Lesson Objective: Explain how various physical properties result from the ionic crystal lattice, such as strength, hardness, high melting points, brittleness, and electrical conductivity.

16. Amethyst is
1. a copper-containing mineral
 2. the common name for mercury (II) sulfide
 3. a form of quartz
 4. a vanadium-containing mineral
17. High melting points of ionic compounds are due to
1. the presence of many ionic bonds
 2. the ordered crystal structure
 3. various shapes of the crystals
 4. mixtures of metals in the crystals
18. Addition of NaCl to distilled water
1. decreases the electrical conductivity
 2. generated electricity
 3. increases negative ion movement
 4. causes conduction of an electrical current
19. Ionic crystals break when struck because
1. the hammer forces all the ions together
 2. ions of like charge are forced closer together and repel one another
 3. ions of opposite charge set up an electrical current
 4. a chemical reaction is initiated due to the pressure of the hammer
20. True/False: Distilled water can conduct an electric current.
21. True/False: Azurite is a blue-purple mineral.
22. True/False: Solid NaCl at room temperature will conduct electricity
23. True/False: Transition metal ions are responsible for the color of minerals.
24. The _____ force between ions of the _____ charge cause a crystal to shatter when struck with a hammer.
25. An electric current is conducted when _____ and _____ ions can _____ .
26. The figure below shows two different ways of representing the ionic crystal lattice. A ball and stick model makes it easier to see how individual ions are oriented with respect to one another. A space filling diagram is a more accurate representation of how the ions pack together in the crystal.
27. How does a model help us understand chemical processes?

28. How do you think the d orbitals might affect the interactions between Ti and Cl as compared to Na and Cl?

8.3 Metallic Bonds

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the electron-sea model of metallic bonding.

1. The points in a metal crystal lattice are occupied by
 1. anions
 2. cations
 3. metal atoms
 4. metal array
2. Electrons moving through the empty orbitals of the metal are called
 1. mobile
 2. delocalized
 3. degenerate
 4. displaced
3. True/False: All the atoms in a metallic crystal are identical.
4. True/False: In a metal, the positive ions migrate from one place to another.
5. Electrons in the _____ energy levels are mobile.
6. A metal can be described as an array of _____ ions surrounded by a sea of _____ electrons.

Lesson Objective: Explain how metallic bonding is responsible for the conductivity and luster of metals.

7. Metals conduct electricity well because
 1. there are large spaces between atoms
 2. electrons flow freely through the metal
 3. the metal cations channel electrons
 4. electrons cluster together
8. The process that gives rise to luster in a metal is called
 1. photoelectric effect
 2. photolysis
 3. photochromatic effect
 4. photometry
9. True/False: Luster of a metal is due to electrons emitting light when they move to higher energy levels.
10. True/False: Metals are good conductors of heat.
11. When electrons flow into one end of a metal, an _____ number flow out the other end.

Lesson Objective: Explain why metals are malleable and ductile, while crystalline ionic compounds are not.

12. Ionic crystals are brittle because

1. ionic bonds are weak
 2. like charges come too close together under force
 3. electron flow increases under force
 4. ions separate further under force
13. The malleability of a metal is due to
1. positive ions repelling one another
 2. negative electrons repelling one another
 3. electrons blocking contact between adjacent cations
 4. attraction between cations and anions
14. True/False: Metals are malleable because the cations can move
15. True/False: Metals can be rolled into thin sheets.
16. A _____ material can easily be drawn into thin wires.
17. A _____ can easily be hammered into thin sheets.

Lesson Objective: Describe how metal atoms are arranged, including the three most common packing systems.

18. _____ is the most efficient way to pack spherical objects
1. hexagonal packing
 2. linear packing
 3. closest packing
 4. body-centered packing
19. One of the following is not a metallic crystal structure
1. cubic body centered
 2. cubic face centered
 3. hexagonal
 4. hexagonal faced
20. True/False: The purpose for packing is to minimize empty space between atoms.
21. True/False: The coordination number for atoms in the fcc structure is 10.
22. Metallic atoms are arranged in _____ structures.
23. The coordination number for atoms in the bcc structure is _____.

Lesson Objective: Identify some common alloys and explain their importance.

24. Bronze is an alloy of
1. tin and gold
 2. copper and tin
 3. copper and zinc
 4. tin and aluminum
25. One of the following is not a constituent of steel
1. molybdenum
 2. manganese
 3. magnesium
 4. cobalt
26. True/False: The major constituent of steel is iron.
27. True/False: Bronze is used to make many musical instruments
28. Steels are used in building construction because of their _____, _____ and resistance to _____.
29. In a _____ alloy, various atoms just replace each other in the crystal structure.

30. In an _____ alloy, carbon atoms fit in between other atoms.
31. When light is shone onto the surface of a metal, its electrons absorb small amounts of energy and become excited into one of its many empty orbitals. The electrons immediately fall back down to lower energy levels and emit light. This process is responsible for the high luster of metals.
32. What does this process tell us about the properties of light?
33. What would be a benefit of incorporating a small atom like carbon into the crystal structure of iron metal?

CHAPTER **9**

Covalent Bonding Worksheets

Chapter Outline

- 9.1 LEWIS ELECTRON DOT STRUCTURES
 - 9.2 MOLECULAR GEOMETRY
 - 9.3 POLARITY AND INTERMOLECULAR FORCES
 - 9.4 HYBRIDIZATION OF ATOMIC ORBITALS
-

9.1 Lewis Electron Dot Structures

Worksheets

Name _____ Class _____ Date _____

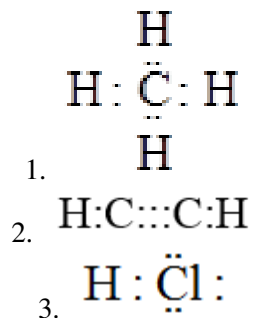
Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe how a covalent bond forms, including the energy change involved in the process.

1. One of the following is not a covalent compound
 1. NH_3
 2. MgBr_2
 3. CO_2
 4. H_2O
2. A covalent bond occurs when
 1. an atom donates an electron to another atom
 2. two atoms donate electrons
 3. two atoms share electrons
 4. two atoms lose an electron
3. True/False: A lower potential energy indicates a more stable system.
4. True/False: When two atoms approach, their electron clouds begin to repel each other.
5. A _____ molecule is one that contains only two atoms.

Lesson Objective: Use the octet rule to draw Lewis electron dot structures for simple molecules. Know how and when to incorporate double and triple bonds into the structures.

6. A covalent bond forms when
 1. two pairs of electrons are shared
 2. two electrons of the same spin overlap
 3. two full orbitals overlap
 4. two singly occupied orbitals overlap
7. A lone pair is a pair of electrons that are
 1. not shared between atoms
 2. shared equally between two atoms
 3. shared more with one atom than the other
 4. alternate between one atom and another
8. True/False: A double covalent bond is a bond shared between two pairs of atoms.
9. Draw the Lewis dot structures for the following compounds:
 1. hydrofluoric acid
 2. CH_3OH
 3. nitrogen trifluoride
10. Draw the structural formulas for the following Lewis dot structures:



Lesson Objective: Understand how a coordinate covalent bond differs from other covalent bonds.

11. In a coordinate covalent bond
 1. no electrons are shared
 2. a double set of electrons are shared
 3. one atom provides both electrons
 4. the electrons alternate between the two atoms
12. True/False: A coordinate covalent bond is stronger than a regular covalent bond.
13. What structural feature is needed for an atom to provide the electrons for a coordinate covalent bond?

Lesson Objective: Be able to draw Lewis structures for polyatomic ions.

14. A polyatomic ion
 1. has more than one electrical charge.
 2. is a group of covalently bonded atoms with an overall charge
 3. is a group of covalently bonded atoms with a distributed charge
 4. is a group of ions held together by an electrical charge
15. One of the following is not a polyatomic ion
 1. NH_4^+
 2. SO_3^-
 3. SO_4^{2-}
 4. O^{2-}
16. True/False: The electrical charge in a polyatomic ion is not localized.
17. True/False: Coordinate covalent bonds are never seen in polyatomic ions.
18. In a polyatomic ion, the charge can be determined by counting the total number of _____ electrons in the structure.
19. Determine the number of valence electrons in the polyatomic ion OH^- .

Lesson Objective: Understand the concept of resonance.

20. A resonance structure
 1. represents electron shifts among atoms
 2. shows multiple structures of the same compound
 3. gives a partial idea of the actual structure
 4. violates the octet rule
21. True/False: a “half-bond” in a resonance structure is shown by a dotted line.
22. How do we know that the electrons in ozone do not rapidly shift back and forth among the three O atoms?

Lesson Objective: Know some common exceptions to the octet rule.

23. One of the following is not an exception to the octet rule.
1. even-electron molecules
 2. odd-electron molecules
 3. incomplete octet
 4. expanded octet
24. True/False: The d sublevel becomes important in covalent bonds for elements of the third period and beyond.
25. Nitrous oxide (N-O) is a covalently bonded compound that is an exception to the octet rule. Which exception? Explain your answer.

Lesson Objective: Relate bond energy to the stability and reactivity of molecules.

26. One of the following has the highest bond energy:
1. H-H
 2. C=C
 3. Cl-Cl
 4. C-H
27. True/False: A high-energy bond is more reactive than a low-energy bond.
28. True/False: Bond energy is the energy required to break a covalent bond.
29. Why would a triple bond be stronger than a single bond?
30. Resonance is the use of two or more Lewis structures to represent the covalent bonding in a molecule. Each of the valid structures is referred to as a resonance structure. It is now understood that the true structure of a molecule that displays resonance is an average or a hybrid of all the resonance structure.
31. Why is it difficult to represent resonance structures on paper?
32. Why do boron and aluminum not form complete octets of valence electrons when forming covalent compounds?

9.2 Molecular Geometry

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

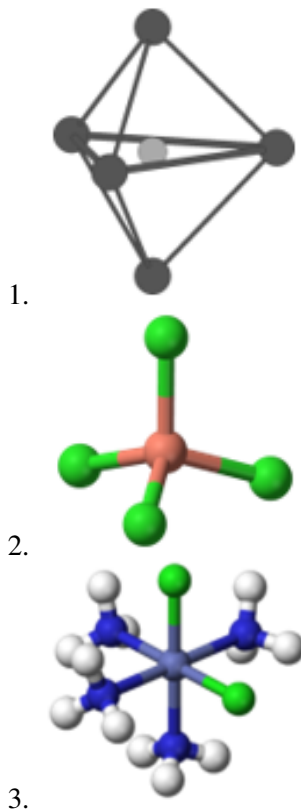
Lesson Objective: Explain the basis of VSEPR theory.

1. Molecular geometry deals with
 1. the size of a molecule
 2. the shape of a molecule
 3. the formula mass of a molecule
 4. the electrical charge on the molecule
2. The VSEPR model states that
 1. electrons will cluster in orbitals
 2. valence electrons can form covalent bonds
 3. valence electron pairs stay far apart
 4. valence electrons bring atoms closer together
3. True/False: Double and triple bonds are no different in repulsion than single bonds.
4. True/False: Valence bond electrons are shown in Lewis structures.
5. VSEPR stands for _____ - _____.
6. For VSEPR purposes, molecules are classified according to the number of _____ and the number of _____ around the _____ atom.

Lesson Objective: Predict the shapes of molecules and polyatomic ions using VSEPR theory.

7. The AB_2 VSEPR configuration is
 1. planar
 2. tetrahedral
 3. linear
 4. polygonal
8. PCl_5 contains two Cl atoms with an _____ orientation.
 1. equilateral
 2. equidistant
 3. equimetric
 4. equatorial
9. The molecular geometry of methane is
 1. tetrahedral
 2. trigonal
 3. pyramidal
 4. equilateral

10. True/False: Dotted line bonds project out of the page.
11. True/False: In PCl_5 , three of the Cl atoms are in the same plane.
12. True/False: An octahedral shape would have seven sides.
13. Name the following VSEPR shape



Lesson Objective: Account for variations in bond angles based on the relative repulsive forces exerted by lone pairs and bonding pairs of electrons.

14. Lone pair electrons are _____ repulsive than bonding pair electrons.
 1. slightly more
 2. much more
 3. slightly less
 4. much less
15. The molecular geometry of the water molecule is
 1. crooked
 2. bent
 3. trigonal
 4. tetrahedral
16. A distorted tetrahedron configuration is seen in the _____ molecule.
 1. SF_5
 2. ClF_3
 3. SF_4
 4. Cl_2F_5
17. True/False: The H-O-H bond angle is 104.5° .
18. True/False: The ammonia molecule contains two lone pairs.
19. True/False: VSEPR treats double and single bonds in different ways.

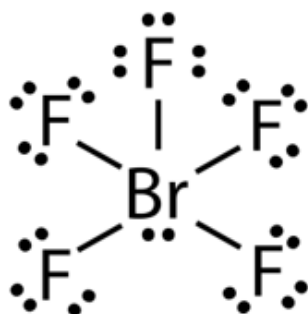
20. Why is the shape of the ammonia molecule trigonal instead of tetrahedral?
 21. Use the diagram below to explain how the lone pair electrons influence the shape of the water molecule.



22.

Lesson Objective: Describe the relationship between molecular geometry and electron domain geometry.

22. Electron domain geometry deals with
1. both bonding electrons and lone pair electrons
 2. lone pair electrons only
 3. bonding electrons only
 4. resonance electrons
23. Lone pair electrons use _____ than bonding pair electrons
1. more space
 2. the same amount of space
 3. less space
 4. variable space
24. True/False: Lone pairs on the atoms surrounding the central atom do not affect the geometry.
 25. True/False: Lone pairs on the central atom are not considered in determining geometry.
 26. The water molecule has a _____ molecular geometry and a _____ electron domain geometry.
 27. Why is the electron domain geometry of ozone called trigonal planar?
 28. In XeF_4 , two of the atoms are replaced by lone pairs, forming a _____ geometry.
 29. The bromine pentafluoride molecule (**Figure 9.1**) has a central bromine atom, five single bonds to fluorine atoms, and one lone pair.



30.

FIGURE 9.1

31. The six groups of electrons around the central atom give it an octahedral domain geometry. Since all of the peripheral atoms in an octahedral molecule are equivalent, any one of them could equally be replaced by the lone pair. The resulting geometry is called
32. *square pyramidal*
33. . A surface covering a square pyramidal molecule is a four sided pyramid on a flat base.
34. What would be the expected molecular geometry of BrF_5

35. 4
36. ? Explain your answer.
37. What is the major difference in the molecular geometry and the electron domain geometry in terms of assigning shapes to molecules?

9.3 Polarity and Intermolecular Forces

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe how the electronegativity difference between two atoms in a covalent bond results in the formation of a nonpolar covalent, polar covalent, or ionic bond.

- Electronegativity is the
 - number of positive charges on an ion
 - relative ability of an atom to attract electrons
 - ability of a cation to attract anions
 - conductivity of a molecule
- A non-polar covalent bond involves
 - equal sharing of electrons
 - strong attraction of electrons by one atom
 - balanced charge distribution in the bond
 - unpaired electrons in orbitals
- The Greek letter δ represents
 - change in state
 - rate of reaction
 - partial charge
 - partial bond energy
- The C-I bond is
 - mostly ionic
 - partially covalent
 - polar covalent
 - mostly covalent
- True/False: A difference in electronegativity greater than 1.8 indicates a mainly ionic bond.
- True/False: The Si-Cl bond is mostly covalent
- True/False: The bond between Cl and Br is non-polar covalent.
- True/False: Bonds between a metal atom and a non-metal atom are usually ionic in nature.
- Define the following terms:
- a. non-covalent polar bond
- b. diatomic element
- c. polar covalent bond
- The atom with the highest electronegativity is _____.
- The atom with the lowest electronegativity is _____.
- Arrange the following bonds in terms of increasing polarity, beginning with the lowest polarity:
 - Ni-P

2. Si-Cl
3. Al-Cl
4. Cu-Br
5. Pb-I

Lesson Objective: Describe how molecular geometry plays a role in determining whether a molecule is polar or nonpolar

13. A polar molecule has
 1. two negative poles
 2. two positive poles
 3. one end is slightly negative
 4. no charge distribution
14. A dipole has
 1. two like charges
 2. two diffuse charges
 3. two poles
 4. two negative poles
15. True/False: Carbon tetrachloride is a non-polar molecule.
16. True/False: The dipoles in the water molecule cancel each other out.
17. Why is BF_3 non-polar?
18. What would be the effect on the polarity of the molecule formed when CCl_4 has one Cl replaced by an H?

Lesson Objective: Distinguish between the following three types of intermolecular forces: dipole-dipole forces, London dispersion forces, and hydrogen bonds.

19. London dispersion forces
 1. are seen between polar atoms
 2. are a result of differing electronegativities between two atoms
 3. are due to random motions of molecules
 4. are due to random motions of electrons
20. Hydrogen bonds involve
 1. H covalently bonding to carbon
 2. H interacting with halogens
 3. H attached to an electronegative atom attracted to a lone pair of electrons
 4. H attached to an electronegative atom attracted to an electronegative atom
21. True/False: A H attached to a Si atom can form a hydrogen bond.
22. True/False: Dispersion forces are a form of van der Waals interactions.
23. _____ are attractions between molecules.
24. Why would London dispersion forces be so weak?

Lesson Objective: Describe how chemical bonding and intermolecular forces influence the properties of various compounds.

25. One of the statements below about water solubility of molecular compounds is not correct
 1. molecules that form hydrogen bonds are usually soluble
 2. molecules that form dispersion forces are usually soluble
 3. molecules that form dipole-dipole forces are usually soluble

4. molecules that form ions are usually soluble
26. When heated, diamond
 1. melts at high temperature
 2. vaporizes to a gas
 3. does not change
 4. converts to graphite
 27. True/False: At room temperature, molecular compounds can be solid, liquid, or gas.
 28. True/False: HBr has a higher boiling point than HF.
 29. Melting and boiling points of molecular compounds are _____ in comparison with ionic compounds.
 30. Why are lone pair electrons needed to form H bonds?
 31. The degree to which a given bond is ionic or covalent is determined by calculating the difference in electronegativity between the two atoms involved in the bond. As an example, consider the bond that occurs between an atom of potassium and an atom of fluorine. Using the table, the difference in electronegativity is $4.0 - 0.8 = 3.2$. Because the difference in electronegativity is relatively large, the bond between the two atoms is primarily ionic. Since the fluorine atom has a much larger attraction for electrons than the potassium atom does, the valence electron from the potassium atom is considered to have completely transferred to the fluorine atom. The figure below shows how the difference in electronegativity relates to the ionic or covalent character of a chemical bond.
 32. Is there a clear difference between an ionic bond and a covalent bond?
 33. Offer one reason why a knowledge of bond polarity might be useful in predicting chemical reactions.

9.4 Hybridization of Atomic Orbitals

Worksheets

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe valence bond theory as it pertains to the formation of a covalent bond between atoms.

1. Valence bonds are formed between _____ electrons in two atoms
 1. matched
 2. equal spin
 3. unpaired
 4. matched
2. The bond length is
 1. the distance between two radii
 2. twice the distance of the atomic nuclei
 3. the optimum distance between the two nuclei
 4. the distance between outer energy levels
3. True/False: At the bond length, the potential energy becomes a minimum.
4. True/False: Overlapping orbitals must be of the same type.
5. State the valence bond theory.
6. Would group one atoms be likely to form covalent bonds? Explain your answer.

Lesson Objective: Describe the process of electron promotion and hybridization during the formation of hybrid orbitals.

7. The H-C-H bond angle in methane is
 1. 107°
 2. 90°
 3. 109.5°
 4. 106.5°
8. The symbol \uparrow indicates
 1. electron magnetic field
 2. unpaired electron
 3. electron orbital shape
 4. electron principal energy level
9. Hybrid orbitals
 1. form as a result of electron repulsion
 2. act to stabilize orbitals
 3. form in preparation for bond formation

4. act to reorient bonds
10. True/False: The process of bond formation is accompanied by a decrease in energy.
11. True/False: Oxygen in compounds has one lone pair of hybridized electrons.
12. True/False: The molecular geometry of ammonia is tetrahedral.
13. Carbon forms a set of hybrid orbitals called _____.
14. Why does helium not form covalent bonds?
15. Explain how oxygen in compounds gets two lone pairs of electrons.

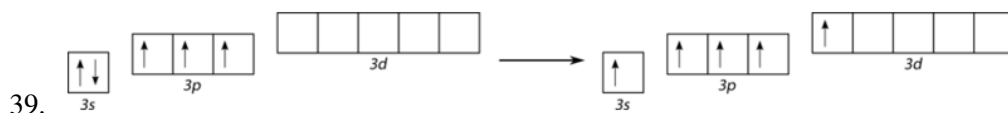
Lesson Objective: Explain the relationship between electron domain geometry and the various types of hybrid orbitals.

16. The geometry of sp orbitals is
 1. tetrahedral
 2. trigonal planar
 3. linear
 4. trigonal pyramidal
17. Trigonal bipyramidal hybridization involves _____ orbitals
 1. two
 2. three
 3. four
 4. five
18. VSEPR can be used for all of the following except
 1. predicting reactivity
 2. predicting molecular geometry
 3. predicting electron domain geometry
 4. hybridization
19. True/False: Hybridization can involve d-orbitals.
20. True/False: Trigonal bipyramidal geometry is associated with sp^3 hybridization.
21. True/False: The sp^2 hybridization is seen in BF_3 .
22. Predict the hybridization expected in AlH_3
23. What type of hybridization would you expect for $SiCl_4$?
24. Does phosphorus have d electrons to use in bonding for PCl_5 ?

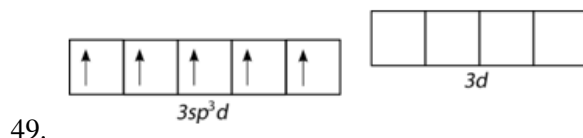
Lesson Objective: Distinguish between sigma and pi bonding.

25. A triple bond consists of
 1. three sigma bonds
 2. two sigma bonds and one pi bond
 3. one sigma bond and two pi bonds
 4. three pi bonds
26. A double bond is formed from
 1. two s orbitals
 2. one s orbital and one p orbital
 3. two p orbitals
 4. one p orbital and one d orbital
27. True/False: The electron domain geometry around each carbon in a double bond is trigonal bipyramidal.
28. True/False: A sigma bond is formed by orbitals overlapping end-to-end.
29. How is the unhybridized $2p$ orbital in a carbon-carbon double bond oriented?

30. In a triple bond, how are the sets of unhybridized p orbitals oriented with relation to one another?
31. Elements in the third period and beyond are capable of expanding their octet to form molecules with either trigonal bipyramidal or octahedral electron domain geometries. In order to accomplish this, the previously unoccupied orbitals in the d sublevel of the central atom are involved in the hybridization process.
32. For phosphorus pentachloride (PCl_5)
33. 5
34.), the electron domain and molecular geometries are trigonal bipyramidal. The electron promotion is from the
35. $3s$
36. orbital to an empty
37. $3d$
38. orbital.



40. The hybridization then occurs with a single
41. s
42. orbital, three
43. p
44. orbitals, and a single
45. d
46. orbital. The set of five hybrid orbitals are called
47. sp^3d
48. hybrids.



50. The overlap of the five hybrid
51. sp^3d
52. orbitals with the
53. $3p$
54. orbital of each chlorine atom results in the five covalent bonds of PCl_5
55. 5
56. . Other molecular geometries derived from trigonal bipyramidal electron domain geometry (seesaw, T-shaped, linear) also display the same hybridization. Is the promotion of an
57. s
58. orbital electron to a
59. d
60. orbital a violation of the Aufbau principle.
61. Would you expect a bond formed by overlapping p orbitals to be more reactive than a sigma bond on the same carbon? Explain your answer.

CHAPTER

10

The Mole Worksheets

Chapter Outline

10.1 THE MOLE CONCEPT

10.2 MASS, VOLUME AND THE MOLE

10.3 CHEMICAL FORMULAS

10.1 The Mole Concept

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Identify three methods for measuring the amount of matter in a sample.

1. One of the following is not a way to determine the amount of a material.
 1. determine the mass
 2. determine the molecular structure
 3. determine the volume
 4. determine the number of molecules
2. An exact quantity
 1. is an approximation
 2. can vary from one place to another
 3. has unlimited significant figures
 4. has ten significant figures
3. True/False: Dimensional analysis involves measuring the size of a material.
4. True/False: Conversion factors allow us to determine the amount of a material in different units.
5. A box of 56 candy bars is purchased. They are to be shared among eight students. How many candy bars does each student get?
6. Two cups of oil are needed for baking. If there are 237 mL/cup, will a 30-ounce bottle of oil be sufficient for the recipe (30 mL/ounce)?

Lesson Objective: Define the mole and its relationship to Avogadro's number.

7. The mole is the SI unit for
 1. volume of a gas
 2. mass of material
 3. number of representative particles
 4. volume of liquid
8. One of the following cannot serve as a representative particle
 1. formula unit
 2. ion
 3. atom
 4. molecule
9. True/False: The representative particle for KBr is the formula unit.
10. True/False: A representative particle can vary depending on the size of the sample.
11. State Avogadro's number
12. Define representative particle.

Lesson Objective: Use Avogadro's number to convert between moles and the number of representative particles of a substance.

13. The molecule $\text{CH}_3\text{CH}_2\text{CHBrCH}_2(\text{CH}_3)_2\text{CCH}_3$ contains ____ carbon atoms
1. seven
 2. nine
 3. ten
 4. eight
14. True/False: A mole of the above compound contains 18 moles of H atoms.
15. How many particles are there in 2.4 moles of CO_2 ?
16. How many moles of NaCl are there in 1.14×10^{24} particles?
17. Nitric acid has the formula HNO_3 . If there are 3.55×10^{23} atoms of O in the sample, how many moles of nitric acid are there?

Lesson Objective: Relate the atomic mass of an element to its molar mass.

18. The atomic mass unit is based on
1. C-13
 2. C-9
 3. C-12
 4. C-11
19. One of the following is not true about the atomic mass of an element
1. is affected by the number of isotopes
 2. is a weighted average of all the isotopes
 3. does not count the unstable isotopes
 4. considers each isotope to be important
20. True/False: Atomic masses are generally not whole numbers.
21. True/False: The atomic mass of an element is usually expressed in grams.
22. How many moles of gold are there in 14.2×10^{23} atoms of gold?
23. There are 4.7 moles of helium in a container. How many atoms of He does the container hold?

Lesson Objective: Calculate the molar mass of a given compound.

24. The molecular mass of a compound is
1. the mass of one molecule
 2. the mass of 6.02×10^{23} molecules
 3. the mass of the central atom
 4. the mass of one ion
25. True/False: The term "formula mass" only applies to ions.
26. True/False: Ionic compounds exist in an extended three-dimensional lattice.
27. Calculate the formula masses of the following compounds
1. KBr
 2. NaMnO_4
 3. $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl}$
 4. $(\text{NH}_4)_2\text{SO}_4$

10.2 Mass, Volume and the Mole

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Vocabulary

- Avogadro's hypothesis
- molar volume
- standard temperature and pressure (STP)

Lesson Objective: Use molar mass to make conversions between mass and moles of a substance.

1. The molar mass of NaCl is _____ g/mol
 1. 58.5
 2. 74.6
 3. 42.0
 4. 42.4
2. The compound whose molar mass is 133.5 is
 1. MgCl_2
 2. AlCl_3
 3. TiBr_2
 4. CsI
3. True/False: The number of moles of product can be determined by measuring the mass of the product.
4. Find the mass of the following materials:
 1. 0.45 moles Au
 2. 1.33 moles Mn.
 3. 3.7 moles Fe_2O_3
 4. 2.9 moles CuCl_2
5. Determine the number of moles in each of the following:
 1. 224 grams IrI
 2. 748 grams NaNO_3
 3. 136 grams $\text{CH}_3\text{CHClCH}_2\text{CH}_3$
 4. 12.7 grams $\text{C}_6\text{H}_2(\text{NO}_2)_3\text{CH}_3$

Lesson Objective: Explain Avogadro's hypothesis and how it relates to the volume of a gas at standard temperature and pressure.

6. All of the following except one are properties of gases.
 1. compressible

2. expands when heated
 3. little space between molecules
 4. actual size of gas particles negligible
7. Conditions for standard temperature and pressure of a gas are
1. 0°C and 1 atm
 2. 0°K and 1 atm
 3. 32°C and 2 atm
 4. 273°K and 2 atm
8. True/False: At STP, one mole of any gas occupies 22.4 L.
9. True/False: As the temperature increases, the gas volume decreases.

Lesson Objective: Convert between moles and volume of a gas at STP.

10. True/False: one mole of gas = 22.4 liters at any temperature.
1. How many moles are there in 45.7 L of gas at STP?
 2. A reaction produced 235 moles of gas. What volume did the gas occupy?
11. Why must conditions be at STP in order for the 22.4 L conversion factor to be used?

Lesson Objective: Calculate the density of gases at STP.

13. One of the following statements about gas density is not true. The density of a gas is affected by
1. molar mass
 2. temperature
 3. pressure
 4. reactivity
14. Gas density can be calculated using the formula
1. $D = \frac{\text{molar mass}}{22.4}$
 2. $D = \text{molar mass} \times 22.4$
 3. $D = \frac{22.4}{\text{molar mass}}$
 4. $D = \frac{\text{molar mass}}{\text{volume}}$
15. True/False: Units for gas density are grams/L
16. True/False: The density of helium is less than the density of air
17. Calculate the densities of the following gases at STP:
1. N₂
 2. SO₂
 3. CH₄
18. Calculate the molar mass of the gases when given the density at STP:
1. 1.43 g/L
 2. 1.25 g/L
 3. 5.40 g/L

Lesson Objective: Use the mole road map to make two-step conversions between mass, number of particles, and gas volume.

19. The ____ is at the center of any calculations in the road map
1. mole

2. mass
 3. density
 4. volume
20. What is the molar mass of a gas that has a mass of 6.5 grams at a volume of 12.4 L at STP?
 21. What volume will be occupied by 3.2 moles of O_2 at STP?
 22. How many particles are in a 45.7 L volume of gas at STP?

10.3 Chemical Formulas

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Calculate the percent composition of a compound either from mass data or from the chemical formula. Use percent composition to calculate the mass of an element in a certain sample of a compound.

- The percent composition is defined as percent by _____.
 - volume
 - mass
 - ions
 - molecular weight
- In determining percent composition, it is assumed that
 - analytical methods are variable
 - different answers are acceptable
 - the composition will vary
 - the material is pure
- True/False: Percent mass = mass of compound/mass of element.
- True/False: The sum of percent masses of the elements must equal 100%.
- Determine the percent composition of the following compounds:
 - NaOH
 - CuBr₂
 - N₂S₂
- Determine the percent composition from the following analytical data:
 - Analysis shows that 2.56 grams of iron reacts with 1.08 grams oxygen.
 - An experiment was performed in which 6.1 g table salt produced 2.6 g Cl in the sample.
 - A reaction uses 18.06 grams Mg to react completely with 6.96 grams N

Lesson Objective: Calculate the percentage of a hydrate's mass that is due to water.

- A hydrate is
 - a molecule containing H and O
 - a molecule that dissolves in water
 - a molecule with one or more water molecules as part of the structure
 - a molecule that produces water when reacted with acids
- Anhydrous CoCl₂ has a _____ color
 - red
 - pink

3. blue
4. green
9. True/False: Many ionic compounds contain water as part of the crystal lattice.
10. True/False: Cobalt (II) chloride hexahydrate contains seven molecules of water.
11. Calculate the percent by mass of water in the following compounds:
 1. $\text{SnCl}_5 \cdot 5\text{H}_2\text{O}$
 2. $\text{FeCl}_3 \cdot 6\text{H}_2\text{O}$
 3. $\text{Ba}(\text{OH})_2 \cdot 8\text{H}_2\text{O}$

Lesson Objective: Determine the empirical formula of a compound from percent composition data.

12. Glucose has the molecular formula of $\text{C}_6\text{H}_{12}\text{O}_6$. The correct empirical formula for glucose is
 1. $\text{C}_3\text{H}_6\text{O}_3$
 2. $\text{C}_2\text{H}_4\text{O}_2$
 3. CH_2O
 4. C_2HO
13. True/False: Only empirical formulas are used to describe ionic compounds.
14. True/False: Elemental analysis tells us the atomic masses of the elements in the compound.
15. Calculate the empirical formula of each compound from the data listed below:
 1. 39.9% C, 6.7% H, 53.4% O
 2. 65.5% C, 5.5% H, 29.0% O
 3. 34.95% C, 6.844% H, 46.56% O, 13.59% N

Lesson Objective: Determine the molecular formula of a compound from the empirical formula and the molar mass.

16. The molar mass of a compound can be determined using a
 1. mass spectrometer
 2. atomic absorption spectrometer
 3. X-ray spectrometer
 4. infrared spectrometer
17. True/False: The molecular formula and the empirical formula for a compound may be the same.
18. Determine the molecular formula from the following data:
 1. empirical formula C_2OH_4 molar mass 88 g/mol
 2. empirical formula $\text{C}_4\text{H}_4\text{O}$ molar mass 136 g/mol
 3. empirical formula CFBrO molar mass 254.7 g/mol

CHAPTER **11**

Chemical Reactions Worksheets

Chapter Outline

11.1 CHEMICAL EQUATIONS

11.2 TYPES OF CHEMICAL REACTIONS

11.1 Chemical Equations

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Describe chemical reactions using word equations.

- In a chemical equation, the arrow stands for
 - yields
 - comes from
 - equals
 - is
- Which of the following is not a word equation?
 - silver + sulfur \rightarrow silver sulfide
 - burning wood \rightarrow carbon dioxide and water
 - drive north \rightarrow turn left at the light
 - water forms hydrogen and oxygen
- A reactant is
 - the material formed in the reaction
 - produced in the reaction
 - the material acted upon
 - the process of reacting
- True/False: Burning is a chemical reaction.
- True/False: Tarnish is formed when silver sulfide reacts with air.
- True/False: Some chemical equations use an equal sign to show products.
- The generic chemical equation: _____ \rightarrow _____ .
- Methane + _____ \rightarrow _____ + _____ .

Lesson Objective: Know the correct symbols to use in order to write skeleton equations for chemical reactions.

- A skeleton equation shows
 - formulas of reactants
 - formulas of products
 - formulas of reactants and products
 - the process of the reaction
- One of the following is not used to describe the physical state of a material in a reaction

- a. s
- b. l
- c. g
- d. w

11. Only one of the following is true: In a skeleton equation, all the formulas must
- a. be written correctly
 - b. have balanced structures
 - c. use names of materials
 - d. use arrows to show direction
12. True/False: A chemical equation is a presentation of a chemical reaction.
13. True/False: The abbreviation (s) stands for “solution”.
14. True/False: Skeleton equations show the relative amounts of materials to be used.
15. Comment on the following equations
- a. $CH_4 \leftarrow CO_2 + H_2O$
 - b. $H_2O_{(s)} \rightarrow H_{2(s)} + O_{2(l)}$
16. What symbols could be written above the arrow and what do they represent?

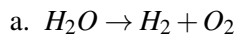
Lesson Objective: Use coefficients to balance chemical equations so that the law of conservation of mass is followed.

17. A balanced equation assumes all of the following except
- a. the law of conservation of mass is true
 - b. amount of reactants does not equal amount of product
 - c. matter can neither be created nor destroyed
 - d. equal numbers of atoms of each kind exist on both sides of the arrow
18. One of the following is not used in balancing equations
- a. change subscripts on formulas
 - b. change coefficients
 - c. double-check accuracy of formulas
 - d. make sure correct products are produced in the reaction
19. If more than one reactant or product are present, they are separated with
- a. dash
 - b. comma
 - c. plus
 - d. bracket
20. Which of the following represents the correct way to balance the equation for the formation of methane from carbon and hydrogen?
- a. $C + H_2 \rightarrow CH_4$
 - b. $C + 2H_2 \rightarrow CH_4$
 - c. $2C + H_2 \rightarrow C_2H_2$

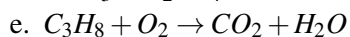
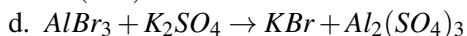
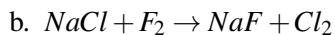
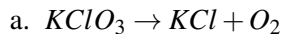


21. True/False: Mass does not need to be conserved in a chemical reaction.
22. True/False: A coefficient is used in front of a chemical to indicate number of atoms in balanced equation.
23. True/False: Subscripts can be changed in order to balance an equation.
24. True/False: If a polyatomic ion is unchanged, it can be treated as a unit.

25. What is the problem with each of the following equations?



26. Balance the following equations:



27. Write balanced equations from the word equations below.

a. sodium + water \rightarrow sodium hydroxide + hydrogen

b. hydrochloric acid + calcium carbonate \rightarrow calcium chloride + water + carbon dioxide

c. iron(III) chloride + sodium hydroxide \rightarrow iron(III) hydroxide + sodium chloride

d. phosphorous + oxygen \rightarrow phosphorus pentoxide

e. silver oxide \rightarrow silver + oxygen

11.2 Types of Chemical Reactions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Define and give general equations for combination, decomposition, single-replacement, and double-replacement reactions.

- The general equation for a decomposition reaction is
 - $A + B \rightarrow AB$
 - $AB \rightarrow A + B$
 - $A + BC \rightarrow AC + B$
 - $AB + CD \rightarrow AD + CB$
- In a decomposition reaction, one of the products can be a
 - compound
 - element
 - atom
 - ion
- Zn reacts with HCl in a _____ reaction.
 - decomposition
 - halogen replacement
 - hydrogen replacement
 - combination
- True/False: There are five basic types of chemical reactions.
- True/False: In a single-replacement reaction, a non-metal can displace a metal.
- True/False: Many metals react easily with acids.
- Metal _____ decompose on heating to form _____ and _____.
- The reaction of sulfur trioxide with water leads to the formation of _____.
- Most decomposition reactions involve the input of _____ in the form of _____, _____, or _____.

Lesson Objective: Classify a reaction as combination, decomposition, single-replacement, double-replacement, or combustion.

- Many metals react with acids to form
 - metal oxides
 - halogens
 - hydrogen gas

- d. bases
11. A metal carbonate decomposes to form
- metal oxide
 - metal oxide + oxygen
 - metal hydroxide
 - metal oxide + carbon dioxide
12. In a double displacement reaction, one product could be all of the following except
- a gas
 - water
 - a precipitate
 - an element
13. True/False: Lead iodide forms a yellow precipitate.
14. True/False: Combustion reactions involve oxygen
15. True/False: One new compound is formed in a double-replacement reaction.
16. Combination reactions can occur when an _____ reacts with a compound to form a compound with a _____ number of atoms.
17. The products of hydrocarbon combustion are _____ and _____.
18. One type of double-displacement reaction leads to the formation of a _____ compound.

Lesson Objective: Use the activity series to correctly predict whether a given reaction will occur.

19. The activity series
- is a list of possible chemical reactions
 - ranks metals in terms of reactivity with water
 - tells which metals will ionize in solution
 - shows which metals displace halogens
20. Only one of the metals below will react with cold water
- Fe
 - Ca
 - Co
 - Mg
21. True/False: The activity series can be used to predict the likelihood of a single-replacement reaction taking place.
22. True/False: Copper reacts with water, but not with acids.
23. In the following equations, indicate “reaction” or “no reaction”. Do not list a product
- $Mg + Co(NO_3)_2 \rightarrow$
 - $Cu + Co(NO_3)_2 \rightarrow$
 - $Pb + Ni(NO_3)_2 \rightarrow$

Lesson Objective: Predict the products of simple reactions, given only the reactants.

24. Zn does not react with one of the following:
- steam
 - acids
 - $Sr(NO_3)_2$
 - $Fe(NO_3)_2$
25. The activity series allows the prediction of products of _____ reactions.
- decomposition
 - single-replacement reactions
 - combustion
 - combination
26. True/False: Silver reacts with acids.
27. True/False: Two new compounds are formed in double-replacement reactions.
28. Write the formulas for the products of the following reactions:
- $Li + Cr(NO_3)_2 \rightarrow$
 - $NaCl + AgNO_3 \rightarrow$
 - $CH_3CH_2CH_2OH + O_2 \xrightarrow{\Delta}$
 - $Mn + O_2 \rightarrow$
 - $CaCO_3 \xrightarrow{\Delta}$
 - $Ba + HCl \rightarrow$

CHAPTER **12** Stoichiometry Worksheets

Chapter Outline

- 12.1 MOLE RATIOS**
 - 12.2 STOICHIOMETRIC CALCULATIONS**
 - 12.3 LIMITING REACTANT AND PERCENT YIELD**
-

12.1 Mole Ratios

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Relate balanced chemical equations to everyday analogies, such as a recipe.

- A recipe for lasagna that serves four requires $\frac{1}{4}$ cup grated Parmesan cheese. You have eight people coming for dinner and want to expand the recipe to feed them. How much Parmesan cheese do you need?
 - $\frac{1}{4}$ cup
 - $\frac{3}{4}$ cup
 - $\frac{1}{2}$ cup
 - $\frac{2}{3}$ cup
- If four extra people show up for dinner, how much more Parmesan cheese do you need in addition to what you used in problem one?
 - $\frac{1}{2}$ cup
 - $\frac{1}{4}$ cup
 - $\frac{3}{4}$ cup
 - $\frac{1}{3}$ cup
- A peanut butter and jelly sandwich (PBJ) requires two slices of bread (B), one serving of peanut butter (PB), and one serving of jelly (J). Write the equation for this sandwich.

Lesson Objective: Define stoichiometry.

- Stoichiometry requires the use of $a(n)$
 - balanced equation
 - unbalanced equation
 - skeleton equation
 - word equation
- True/False: _____ Stoichiometry allows calculations of needed amounts of chemicals in a reaction.
- You have fifteen slices of bread and five servings of peanut butter. How many sandwiches can you make?
- Each morning for breakfast, you have two slices of toast, one bowl of Crunchy Puffs (30 g), one glass of milk (8 oz), and one banana. You have twelve slices of bread, a full box of cereal (250 g), 36 oz milk, and four bananas. How many full servings of breakfast do you have?
- An acoustic guitar has six strings, six tuning pegs, twenty frets, and one pick guard. A guitar maker has 90 frets to use. How many guitars can be made?
- For each of the following reactions, determine the mass of reactants and products:
 - $8CO + 17H_2 \rightarrow C_8H_{18} + 8H_2O$
 - $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
 - $TiCl_4 + 2H_2O \rightarrow TiO_2 + 4HCl$

Lesson Objective: Use mole ratios to convert between amounts of substances in a chemical reaction.

10. In a laboratory situation, it is common to express the amounts of materials in all of the ways below except
1. mL
 2. moles
 3. grams
 4. number of particles
11. Mole ratio is a conversion factor that converts
1. moles to grams
 2. grams to moles
 3. ratio of one material to another
 4. ratio of atoms in the equation
12. True/False: Chemical equations express the relative amounts of reactants and products in a reaction.
13. True/False: In a chemical reaction, the number of moles of reactants will equal the number of moles of product.
14. True/False: In a mole ratio problem, the given amount is written first.
15. $C_2H_6O + 3O_2 \rightarrow 2CO_2 + 3H_2O$
16. Write the mole ratios of
1. $O_2 \rightarrow CO_2$
 2. $C_2H_6O \rightarrow CO_2$
 3. $C_2H_6O \rightarrow O_2$
16. $12HClO_4 + P_4O_{10} \rightarrow 4H_3PO_4 + 6Cl_2O_7$
17. Write the mole ratios of
1. $HClO_4 \rightarrow H_3PO_4$
 2. $HClO_4 \rightarrow Cl_2O_7$
 3. $P_4O_{10} \rightarrow H_3PO_4$
17. $2NaCl + F_2 \rightarrow 2NaF + Cl_2$
1. How many moles of NaF will be produced from 1.2 moles F_2 ?
 2. How many moles of $NaCl$ are needed to form 3.7 moles Cl_2 ?
18. $2C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
1. How many moles of C_3H_8 are needed to produce 4.9 moles CO_2
 2. How many moles O_2 are needed for the complete combustion of 4.2 moles C_3H_8 ?
19. $Pb(OH)_2 + 2HCl \rightarrow 2H_2O + PbCl_2$
1. How much $Pb(OH)_2$ is required to form 3.6 moles of $PbCl_2$
 2. How much HCl is needed for that reaction?
 3. 4.7 moles $Pb(OH)_2$ will require how many moles HCl ?
20. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
1. How much ammonia is needed to produce 3.2 moles NO ?
 2. How much oxygen is needed for this process?
 3. 6.9 moles ammonia will produce how much H_2O ?

12.2 Stoichiometric Calculations

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Calculate the amount in moles of a reactant or product from the mass of another reactant or product. Calculate the mass of a reactant or product from the moles of another reactant or product.

Lesson Objective: Calculate the mass of a reactant or product from the mass of another reactant or product.

3. 76 g SnO_2 are used in the following reaction: $\text{SnO}_2 + 2\text{H}_2 \rightarrow \text{Sn} + 2\text{H}_2\text{O}$
 4. How many grams of
 5. Sn
 6. are formed?
-
4. The following reaction produces 35.1 g of NaCl : $\text{Na}_3\text{PO}_4 + 3\text{HCl} \rightarrow 3\text{NaCl} + \text{H}_3\text{PO}_4$
 - a. How many grams of Na_3PO_4 are needed?
 - b. How many grams H_3PO_4 are formed?

Lesson Objective: Create volume ratios from a balanced chemical equation.

5. The combustion of methane gas is described by the equation $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$.
 - a. What volume of oxygen is needed at STP to combust 4.7 L methane?
 - b. How much carbon dioxide will be formed?
6. Carbon monoxide can be converted to carbon dioxide in the presence of oxygen as seen in the equation $2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$.
 - a. How many liters of oxygen at STP are needed to react with 17.3 L CO ?
 - b. How much carbon dioxide will be formed?

Lesson Objective: Use volume ratios and other stoichiometric principles to solve problems involving mass, molar amounts, or volumes of gases.

8. In the reaction $\text{Ba}_3\text{N}_2 + 6\text{H}_2\text{O} \rightarrow 3\text{Ba}(\text{OH})_2 + 2\text{NH}_3$, there are 84.3 grams $\text{Ba}(\text{OH})_2$ formed.
 - a. What is the volume of ammonia produced at STP?
 - b. How many moles of Ba_3N_2 are needed?

12.3 Limiting Reactant and Percent Yield

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Analyze a chemical reaction in order to determine which reactant is the limiting reactant and which is the excess reactant.

- The reagent which determines how much product will be made is the _____ reagent
 - deficient
 - rate-limiting
 - limiting
 - restricted
- The excess reagent is the material that is
 - used up in the reaction
 - left over
 - brings stoichiometric balance
 - makes extra product
- A balanced equation tells us
 - the masses of materials needed
 - the masses of the products
 - the mole ratio needed for the reaction to occur
 - the number of representative units of the products
- True/False: All reactions need an excess reagent.
- True/False: Limiting reagents produce multiple products.
- You find six frozen pizza crusts in the freezer and decide to make pizza. For each pizza, you need 8oz pizza sauce, $\frac{1}{2}$ cup sliced pepperoni, and 4 oz shredded mozzarella cheese. You check your available ingredients and find you have 64 oz pizza sauce, 2.5 cups sliced pepperoni and 30 oz mozzarella cheese. What is the limiting reagent and how many pizzas can you actually make?
- In the reaction $2Al + 3I_2 \rightarrow 2AlI_3$, determine the limiting reagent if 1.6 mol Al and 2.9 mol I_2 are used
- Propane burns in oxygen according to the following reaction: $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$
- If 14.g grams propane reacts with 3.44 grams oxygen, identify the limiting reagent.
- The reaction shown below uses 30 g ammonium nitrate and 50 g sodium phosphate. Determine the limiting reagent
- $3NH_4NO_3 + Na_3PO_4 \rightarrow (NH_4)_3PO_4 + 3NaNO_3$

Lesson Objective: Calculate the amount of excess reactant remaining after a reaction is complete.

- To identify the excess reagent, all data must be in _____ quantities
 - molal
 - molecular

3. mass
4. molar
11. True/False: A balanced equation allows us to determine the excess reagent.
12. Iron reacts with sulfur in the following process: $Fe(s) + S(l) \rightarrow FeS(s)$. If 7.5 g iron reacts with 9.1 g sulfur, what is the excess reagent and how much of that reagent is left at the end of the reaction?
13. Calcium carbonate and iron(III) phosphate react as shown in the equation below:
14. $3CaCO_3 + 2FePO_4 \rightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$
15. In one experiment, 95.5 g calcium carbonate are combined with 43.2 g iron(III) phosphate. What is the excess reagent in the process and how much of that reagent remains when the reaction is completed?

Lesson Objective: Calculate the theoretical yield of a reaction when the available amounts of each reactant are known.

14. The theoretical yield is
 1. the amount you hope to get
 2. the maximum amount that can be formed from the given amounts of reactants
 3. The amount you need for the next reaction
 4. the actual amount you obtain in the reaction
15. True/False: The theoretical yield is often below 90%.
16. Calculate the theoretical yield in each of the following situations:
 - a. $C_6H_6 + Cl_2 \rightarrow C_6H_5Cl + HCl$ calculate the yield of C_6H_5Cl from 40.7 g benzene.
 - b. $CS_2 + 3O_2 \rightarrow CO_2 + 2SO_2$ calculate the yield of SO_2 from 27 g carbon disulfide.
 - c. $H_3PO_4 + 3KOH \rightarrow K_3PO_4 + 3H_2O$ calculate the yield of K_3PO_4 from 86.3 g KOH .

Lesson Objective: Calculate the percent yield of a reaction based on the theoretical and actual yields.

17. Actual yield is
 1. the amount you obtain in the reaction
 2. the amount you want to obtain in the reaction
 3. the amount of product using excess reagent
 4. the amount of product before purification
18. Percent yield can be determined using
 1. the fraction you actually obtain
 2. actual yield minus theoretical yield
 3. theoretical yield divided by actual yield
 4. actual yield divided by theoretical yield
19. One of the following is not a cause for percent yield being less than 100%
 1. incomplete reaction
 2. uncalibrated balance
 3. side reactions
 4. impure product
20. True/False: Actual yield can be affected by products that are not dry.
21. True/False: Impurities should be calculated as part of actual yield.
22. True/False: A high percent yield indicates a more efficient and economical reaction
23. Carry out the percent yield calculations in the following situations:

- a. Theoretical yield: 14.7 g
- b. Actual yield: 12.9 g Percent yield: ?
- c. Theoretical yield: 89.5 g
- d. Actual yield: ? Percent yield: 56.4%
- e. Theoretical yield: ?
- f. Actual yield: 27.4 g Percent yield: 68.3%

CHAPTER 13 States of Matter Worksheets

Chapter Outline

13.1 KINETIC - MOLECULAR THEORY AND GASES

13.2 LIQUIDS

13.3 SOLIDS

13.4 CHANGES OF STATE

13.1 Kinetic - Molecular Theory and Gases

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: State the main points of the kinetic molecular theory, and describe how it relates to the properties of an ideal gas.

- All of the following except one are forms of matter
 - gas
 - solid
 - electronegativity
 - liquid
- Which of the following is the correct statement about the effects of temperature (T) on the average kinetic energy of gas particles (KE)?
 - increase in T → decrease in KE
 - increase in T → no effect on KE
 - decrease in T → increase in KE
 - decrease in T → decrease in KE
- The forces of attraction between molecules of an ideal gas are
 - ionic
 - non-existent
 - dipole-dipole
 - electrostatic
- The kinetic-molecular theory assumes that
 - molecules are always in motion
 - above room temperature, molecules are in motion
 - molecules are not in motion at temperatures below freezing
 - temperature has no effect on molecular motion
- True/False: Gas molecules are far apart.
- True/False: An ideal gas is an imaginary gas.
- True/False: The kinetic-molecular theory only applies to gases.
- True/False: Collisions between gas molecules result in loss of kinetic energy.
- Gas particles are in constant _____ motion in _____ directions.
- There are no forces of _____ or _____ between gas particles.
- The kinetic-molecular theory as applied to gases has _____ basic assumptions.

Lesson Objective: Define pressure and describe how gases exert pressure.

12. Pressure equals
- $\frac{\text{force}}{\text{area}}$
 - $\frac{\text{area}}{\text{force}}$
 - force \times area
 - force + area
13. The earth's atmosphere exerts pressure because
- high temperatures increase gas motions
 - gravity pulls gas molecules down
 - more gas molecules = more pressure
 - higher altitude = higher pressure
14. True/False: A heavier rock exerts more pressure on the ground than a lighter rock.
15. True/False: Gas pressure results from collisions of gas molecules with the sides of a container.
16. Pressure is defined as the _____ per unit area on a _____.

Lesson Objective: Understand the barometer and how it measures atmospheric pressure. Convert between units of gas pressure.

17. A barometer is used to measure
- gas pressure for experiments
 - atmospheric pressure
 - vacuum pressure
 - gas pressure in enclosed systems
18. Only one of the following is not used to measure air pressure
- manometer
 - aneroid barometer
 - mercury barometer
 - sphygmomanometer
19. True/False: Evangelista Torricelli invented the barometer.
20. True/False: The mm Hg is the standard unit of pressure.
21. Perform the following conversions:
- 698 torr to kPa
 - 107 kPa to atm
 - 0.9 atm to mm Hg
 - 93.2 kPa to mm Hg

Lesson Objective: Relate temperature to average kinetic energy.

22. One of the following statements about absolute zero is not true
- absolute zero = zero degrees K

- b. absolute zero is the temperature at which there is no electron movement
 - c. absolute zero is the temperature at which there is no molecular motion
 - d. absolute zero is based on a theoretical limit
23. The kinetic energy of particles in a sample of NaCl
- a. increase with an increase in temperature
 - b. decrease with an increase in temperature
 - c. increase with a decrease in temperature
 - d. is not affected by a change in temperature
24. True/False: At room temperature, the average kinetic energy of ice is the same as the average kinetic energy of steam.
25. True/False: Absolute zero has been reached in the laboratory.
26. As temperature _____, the range of kinetic energies increases and the distribution curve _____.
27. When a person stands on the floor, his feet exert pressure on the surface. That pressure is related to both the mass of the person and the surface area of his feet. If the person were holding a heavy object, the pressure would increase because of a greater force. Alternatively, if the person stands on his toes, the pressure also increases because of a decrease in the surface area.
28. A ballerina has a mass of 54.5 kg. Her feet occupy a space of 220 cm
29. ²
30. . Calculate the pressure exerted on the floor when she is standing flat-footed. When dancing
31. *en pointe*
32. (on the tips of her toes), her shoes occupy a space of 30 cm
33. ²
34. . Calculate the new pressure.
35. One drop of water has a volume of about 0.05 mL. How many atoms of helium at STP would occupy that volume?

13.2 Liquids

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe a liquid according to the kinetic-molecular theory.

- One of the following statements about gases is incorrect
 - gas particles are very far apart
 - gases cannot be compressed
 - gas particles do not interact with one another
 - intermolecular forces can be ignored for gases
- One of the following statements about liquids is false
 - intermolecular forces hold liquid particles together
 - liquids and gases are fluids
 - liquids can be compressed
 - there is very little empty space between particles of a liquid
- True/False: Liquids are much more dense than gases
- True/False: A liquid has a defined shape.
- A _____ is a substance that is capable of _____ from one place to another.

Lesson Objective: Describe how a liquid exhibits surface tension.

- One of the following statements about liquids is not true
 - particles at the surface are pulled downward
 - particles at the surface are pulled upward
 - particles at the surface are pulled sideways
 - particles at the surface are pulled into the liquid
- Only one of the following effects is produced in water because of its surface tension
 - floating of rock
 - floating of paper clip
 - floating of rowboat
 - floating of carbon dioxide
- True/False: Liquids with strong intermolecular forces show high surface tensions.
- True/False: Surface tension is a measure of the elastic force on the surface of the liquid.
- The _____ of a liquid is a measure of the _____ force in the liquid's surface.
- Which of the molecules shown below will exhibit the highest surface tension? Explain your answer.
 - $\text{CH}_3\text{CH}_2\text{CH}_3$
 - $\text{CH}_3\text{-O-CH}_3$
 - $\text{CH}_3\text{CH}_2\text{OH}$

Lesson Objective: Describe the evaporation of a liquid and its relationship to the kinetic energy of the evaporating particles.

12. During the process of evaporation
 1. molecules with lower kinetic energy escape first
 2. as evaporation proceeds, the temperature of the liquid increases
 3. an increase in temperature of the liquid will increase the rate of evaporation
 4. liquid in a closed container will evaporate more rapidly than liquid in an open container.
13. Evaporation of a liquid requires all of the following except
 1. high enough temperatures
 2. high enough kinetic energy so that particles can escape the surface
 3. an open system
 4. very high atmospheric pressure
14. True/False: A liquid whose particles have higher kinetic energies will evaporate more slowly than a liquid whose particles have a lower kinetic energy.
15. Define the following terms
 1. vaporization
 2. evaporation
 3. condensation

Lesson Objective: Define vapor pressure and understand its relationship to intermolecular forces and to the temperature of the liquid.

16. Liquid in a closed container behaves in the following way:
 1. rate of evaporation equals rate of condensation
 2. rate of condensation is greater than rate of evaporation
 3. collisions with the side of the container increases the kinetic energy of the liquid
 4. rate of evaporation is greater than rate of condensation
17. One of the following statements about vapor pressure is not true
 1. based on strength of intermolecular forces
 2. constant for a given substance at a set temperature
 3. can be measured using a manometer
 4. a liquid with weak intermolecular forces has a low vapor pressure
18. True/False: Liquid in a closed system demonstrates a dynamic equilibrium between vapor and liquid.
19. True/False: A given liquid will evaporate less quickly when it is heated.
20. True/False: The temperature dependence of vapor pressure for a liquid is linear.

Lesson Objective: Describe the process of boiling and differentiate between boiling point and normal boiling point.

21. In the four settings below, the boiling point of water is highest at
 1. Mauna Kea, Hawaii (13,803 feet above sea level)
 2. Denver, Colorado (5183 feet above sea level)
 3. New Orleans, Louisiana (4 feet above sea level)
 4. Death Valley, California (282 feet below sea level)
22. The standard pressure for measuring boiling points is
 1. 750 torr

2. 760 mm Hg
 3. 725 torr
 4. 745 mm Hg
23. True/False: At the boiling point of a liquid, all molecules in the liquid have enough kinetic energy to vaporize.
24. True/False: Evaporation of a liquid does not occur below the boiling point.
25. The _____ is the temperature at which the vapor pressure of a liquid is equal to the _____ pressure.

Lesson Objective: Use a vapor pressure curve to determine boiling points at different atmospheric pressures.

26. The vapor pressure curve shows all of the following except
1. an increase in temperature produces an increase in pressure
 2. water has the highest boiling point
 3. diethyl ether has stronger intermolecular forces than water
 4. at 200 mm Hg, the boiling point of ethanol is less than that of ether at 760 mm Hg
27. True/False: Pressure cookers increase the boiling point of water because the water vapor cannot escape.
28. Predict the boiling point of the following materials at a pressure of 400 mm
1. diethyl ether
 2. ethyl alcohol
 3. water
29. Liquids and solids are referred to as the condensed states of matter. One way in which liquids and gases are similar is that they are both fluids. A **fluid** is a substance that is capable of flowing from one place to another and takes the shape of its container.
30. Give an example of a gas flowing from one place to another.
31. Would you be able to brew a decent cup of coffee on top of Mt. Everest?

13.3 Solids

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe a solid according to the kinetic-molecular theory.

1. One of the following is not a characteristic of a solid
 1. particles packed tightly
 2. properties of a fluid
 3. densest of three states of matter
 4. almost completely incompressible
2. One of the following is not true about melting point of a solid
 1. same as freezing point of corresponding liquid
 2. dependent upon strength of attractive forces
 3. at melting point, solid and liquid are in equilibrium
 4. ionic bonds contribute to low melting points
3. True/False: Vibrational forces disrupt attractive forces at melting point.
4. True/False: Solid particles are organized in orderly arrangements.
5. The melting point is the temperature at which a _____ changes into a _____.

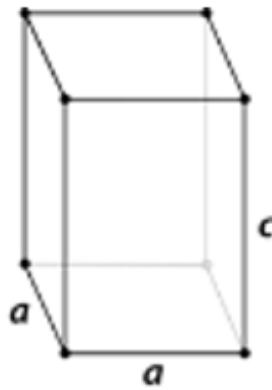
Lesson Objective: Understand that a solid also has a vapor pressure, and describe the relationship between the vapor pressure of a solid and sublimation.

6. One of the following does not undergo sublimation
 1. iodine crystals
 2. solid carbon dioxide
 3. gold nuggets
 4. ice cubes
7. _____ is the term used for the change of state from a gas to a solid.
 1. freezing
 2. crystallization
 3. deposition
 4. precipitation
8. True/False: Dry ice is not a good coolant because it readily melts into liquid slush.
9. True/False: Iodine can sublime at room temperature.
10. Sublimation is the change of state from a _____ to a _____ without passing through the _____ state.

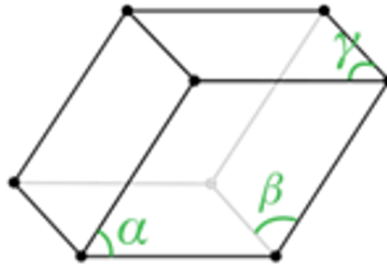
Lesson Objective: Describe the features of the seven basic crystal systems.

11. One of the following is not a crystal system

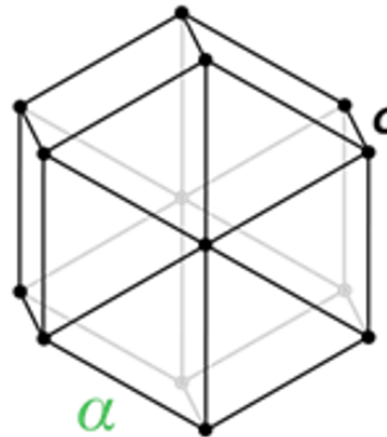
1. pyramidal
 2. monoclinic
 3. tetragonal
 4. hexagonal
12. In the cubic crystal structure
1. $a \neq b = c$
 2. $a = b = c$
 3. $a \neq b \neq c$
 4. $a = b \neq c$
13. True/False: In the orthorhombic crystal, the angles between the faces are all 90° .
14. True/False: Particles of a solid crystal are either ions or molecules.
15. A crystal is a substance in which the particles are arranged in an _____, repeating, _____-
pattern.
16. Name the crystal systems illustrated below.



1.



2.



3.

Lesson Objective: Define a unit cell.

17. One of the following is not a cubic crystal system unit cell

1. simple cubic
 2. rhombic cubic
 3. face-centered cubic
 4. body-centered cubic
18. True/False: A crystal is a unit cell repeated over and over again in three dimensions.
19. A _____ is the smallest portion of a crystal lattice that shows the three-dimensional pattern of the entire crystal.

Lesson Objective: List the four classes of crystalline solids and describe the properties of each.

20. Molecular crystals are held together by all the following forces except
1. hydrogen bonds
 2. dispersion forces
 3. ionic forces
 4. dipole-dipole forces
21. Only one of the following solids is an example of a covalent network crystal
1. NaCl
 2. SiO₂
 3. Na
 4. NH₃
22. True/False: Ionic crystals may be composed of either monatomic or polyatomic ions.
23. True/False: Metallic crystals are poor conductors of electricity.
24. List two examples of covalent network crystals.
25. Why do covalent network crystals have very high melting points?

Lesson Objective: Describe an amorphous solid.

26. All of the following are characteristics of glass except
1. distinct melting point
 2. shatters in an irregular way
 3. made by cooling a mixture of materials
 4. soften when heated
27. True/False: Rubber and plastic are two examples of amorphous solids.
28. Define an amorphous solid.
29. When one of the noble gases is cooled and solidified, the lattice points are individual atoms rather than molecules. However, because the atoms are held together by dispersion forces and not by covalent or metallic bonds, the properties of such a crystal are most similar to the crystals of molecular substances.
30. What properties would you expect to see in the noble gas crystals?
31. Why would metallic crystals not be classified as ionic crystals?

13.4 Changes of State

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Interpret heating and cooling curves.

- Water molecules in a block of ice that is being heated
 - absorb energy and vibrate faster
 - absorb energy and vibrate slower
 - release energy and vibrate faster
 - release energy and vibrate slower
- During the melting process, the solid and liquid states
 - lose energy
 - are in equilibrium with one another
 - are constantly changing
 - exist independently of one another
- _____ bonds keep the water molecules together in the solid and liquid states
 - dipole-dipole
 - covalent
 - van derWaals
 - hydrogen
- The transition from liquid water to ice involves
 - addition of energy
 - breaking of hydrogen bonds
 - loss of energy
 - breaking of covalent bonds
- True/False: The plateau portions of a heating curve indicate melting and boiling points for the material.
- True/False: The heating curve for carbon dioxide has two plateaus.
- True/False: An increase in kinetic energy of water molecules produces the phase changes in a heating curve.
- True/False: At the boiling point, the temperature will continue to increase as the liquid water changes into steam.
- The melting point of mercury is 234.3 K and the boiling point is 629.7 K.
 - Draw the heating curve for mercury
 - Why were K temperature units used instead of Celsius?
- Draw a heating curve for carbon dioxide.

Lesson Objective: Know the terms for the six different changes of state.

- At 100°C, water undergoes _____ to become steam

1. melting
 2. condensation
 3. vaporization
 4. sublimation
12. Ice undergoes _____ at 273 K.
1. deposition
 2. condensation
 3. melting
 4. vaporization
13. At 373 K, steam undergoes _____ to become liquid.
1. condensation
 2. deposition
 3. freezing
 4. sublimation
14. For water, the transition from liquid to solid occurs at _____ .
1. 32°C
 2. 279 K
 3. 0°C
 4. 0 K
15. True/False: The transition from solid to liquid is called melting.
16. True/False: Carbon dioxide only transitions from solid to liquid.
17. True/False: The transition from liquid to vapor is called condensation.
18. True/False: Water undergoes three transition processes.
19. Define the following terms:
1. sublimation
 2. melting
 3. evaporation
 4. condensation
 5. freezing
 6. deposition

Lesson Objective: Describe the general features of a phase diagram, including the triple point and the critical point.

20. If the pressure on a gas is increased, all of the following can happen except
1. the gas can form a liquid
 2. the gas molecules are closer together
 3. the gas will sublime
 4. a solid can form
21. If the pressure on liquid water at 100° is increased,
1. the water converts more easily to gas
 2. the water remains in the liquid form
 3. more solid water is formed
 4. density of the water decreases
22. True/False: The solid state for water is less dense than the liquid state.
23. True/False: At a high enough pressure, a gas can be converted to a solid
24. Define the following terms:

1. phase diagram
 2. triple point
 3. critical temperature
 4. critical pressure
25. At 373.99°C, particles of water in the gas phase are moving very, very rapidly. At any temperature higher than that, the gas phase cannot be made to liquefy, no matter how much pressure is applied to the gas.
26. Why can't water condense at temperatures higher than this?
27. When solid water melts, energy must be put into the system until that process is complete. Why does the temperature remain constant during this process even though additional energy is being added?

CHAPTER

14**The Behavior of Gases
Worksheets****Chapter Outline**

- 14.1 GAS PROPERTIES**
 - 14.2 GAS LAWS**
 - 14.3 IDEAL GASES**
 - 14.4 GAS MIXTURES AND MOLECULAR SPEEDS**
-

14.1 Gas Properties

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Describe how a gas can be compressed.

- The pressure in a typical scuba tank is
 - 700-800 atm
 - 700-800 torr
 - 200-300 atm
 - 200-300 kPa
- At STP, the typical scuba tank will hold approximately _____.
 - 2500 L
 - 3500 L
 - 2500 mL
 - 3500 mL
- Gases can be compressed because
 - gas molecules are more flexible
 - gas molecules are farther apart
 - gas molecules have less energy
 - gas molecules move more slowly
- True/False: Gases are more compressible than liquids and solids.
- True/False: A given volume of gas has a definite shape.
- True/False: When gases are compressed, the particles are forced closer together.
- Gases are unlike other states of matter in that a gas _____ to fill the _____ and _____ of its container.
- Compressibility is a measure of how much a given _____ of matter _____ when placed under _____.

Lesson Objective: Identify three factors that affect gas pressure. Describe the effects according to the kinetic-molecular theory.

- Gas pressure is affected by all of the following except
 - molecular structure of gas
 - temperature
 - volume of container
 - amount of gas
- If the amount of gas in a rigid container is increased,

- a. the temperature will decrease
 - b. the pressure will increase
 - c. the pressure will decrease
 - d. the pressure will not change
11. If the volume of the container is decreased,
- a. the pressure will decrease
 - b. there will be fewer gas collisions with the walls of the container
 - c. the pressure will increase
 - d. the pressure will not be affected
12. Increasing the volume of a container has the same effect on pressure as
- a. decreasing the number of particles
 - b. increasing the temperature
 - c. increasing the number of particles
 - d. changing the container structure
13. An unvented can placed in a fire will
- a. freeze
 - b. explode
 - c. ignite
 - d. collapse
14. An increase in temperature causes air molecules to
- a. increase in velocity
 - b. decrease in velocity
 - c. react with their surroundings
 - d. become converted to other structures
15. One of the following does not produce an increase in pressure
- a. increase in temperature
 - b. decrease in volume
 - c. increase in number of particles
 - d. increase in volume
16. True/False: Basketballs are checked for pressure before a game.
17. True/False: Adding more gas particles to a rigid container will lower the pressure.
18. True/False: The air in a fully inflated car tire is at a higher pressure than the outside air.
19. True/False: A decrease in the volume of a container causes gas particles to collide less frequently with the container wall.
20. True/False: An increase in temperature increases the kinetic energy of gas particles.
21. True/False: Doubling the Kelvin temperature doubles the pressure on a container.
22. True/False: Doubling the number of particles in a container decreases the number of collisions with the vessel walls.
23. Gas particles move _____ and in _____ lines until they _____ collide with other gas particles or with one of the walls of their container.
24. The collisions with the walls of the container define the _____ of the gas.
25. List the four variables used to describe the condition of a gas.

26. Gases are compressible because most of the volume of a gas is composed of the large amounts of empty space between the gas particles. At room temperature and standard pressure, the average distance between gas molecules is about ten times the diameter of the molecules themselves. When a gas is compressed, as when the scuba tank is being filled, the gas particles are forced closer together.
27. If the gas pressure is doubled, what would be the distance between gas molecules?
28. Summarize the behavior of a gas as a function of the rate of collision with the container walls.

14.2 Gas Laws

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Use Boyle's law to calculate pressure-volume changes at constant temperature.

- One of the following statements about the kinetic-molecular theory is incorrect
 - the volume of gas particles is negligible
 - attractive forces between gas molecules can be ignored
 - there is little space between gas particles
 - gas particles have elastic collisions with container walls
- Robert Boyle
 - developed many gas laws
 - is considered a founder of the science of chemistry
 - was a French scientist
 - studied the relationship between temperature and amount of gas
- Boyle's Law deals with the relationship between
 - pressure and temperature
 - volume and temperature
 - pressure and volume
 - temperature and amount of gas
- True/False: In an inverse relationship, as one variable increases, the other variable decreases.
- True/False: Temperature is a variable when using Boyle's Law.
- True/False: A plot of data for Boyle's Law gives a straight line.
- The volume of a given mass of gas varies _____ with the _____ when the _____ is kept constant.
- Perform the following calculations:
 - What is the new volume of 500 mL of gas when the pressure changes from 750 torr to 980 torr? Assume a constant temperature.
 - A constant T, a 650 mL container of helium at 500 torr expands to 900 mL. What is the new pressure?

Lesson Objective: Use Charles's law to calculate volume-temperature changes at constant pressure.

- The temperature scale used for Charles' Law calculation is
 - Celsius
 - Fahrenheit
 - Kelvin

- d. Centigrade
10. A gas brought to extremely low temperatures will
- become liquid
 - increase in volume
 - lose molecules
 - remain a gas
11. True/False: As temperature increases at constant pressure, the volume decreases.
12. True/False: A plot of data for Charles' Law gives a straight line.
13. Charles' Law states that the volume of a given mass of gas varies directly with the _____ temperature of the gas when the _____ is kept constant
14. Make the following calculations:
- At constant pressure, the temperature of 350 mL of hydrogen is increased from 345K to 525K. What is the new volume?
 - At constant pressure, the volume of nitrogen gas decreases from 980 mL at 350K to 670 mL when the temperature is changed. What is the new temperature for the gas?

Lesson Objective: Use Gay-Lussac's law to calculate pressure-temperature changes at constant volume.

15. Gay-Lussac's Law requires a _____ container
- flexible
 - expanding
 - rigid
 - variable
16. Gay-Lussac was a _____ chemist.
- Belgian
 - Spanish
 - Italian
 - French
17. True/False: A plot of data for Gay-Lussac's Law give a straight line.
18. True/False: An increase in temperature produces an increase in kinetic energy of gas particles.
19. As a gas is cooled at constant _____, its _____ continually decreases until the gas condenses to a liquid.
20. Make the following calculations:
- A steel cylinder of oxygen gas shows a pressure of 3400 torr at 30°C. What is the pressure when the tank is at the top of Annapurna where the temperature is -10°C.
 - A rigid cylinder of nitrogen shows a pressure of 4000 torr at 25°C. The cylinder is then transported to the desert where it shows a pressure of 4300 torr. What is the temperature in the desert?

Lesson Objective: Use the combined gas law to solve problems in which pressure, volume, and temperature all change.

21. In combined gas law problems, only the _____ is held constant.
- volume of gas
 - amount of gas
 - pressure of gas

- d. temperature of gas
22. In combined gas law calculations
- pressures can only be in torr
 - volumes can only be in liters
 - temperatures can only be in Kelvin
 - pressures can only be in kPa
23. True/False: In the combined gas law expression, T is in the numerator.
24. True/False: In the combined gas law expression, V is in the numerator.
25. The combined gas law expresses the relationship between the _____, _____, and _____ temperature of a _____ amount of gas
26. Carry out the following calculations:
- $P_1 = 230 \text{ torr}$ $T_1 = 473 \text{ K}$ $V_1 = 65 \text{ mL}$
 - $P_2 = ?$ $T_2 = 690 \text{ K}$ $V_2 = 37 \text{ mL}$
 - $P_1 = 876 \text{ torr}$ $T_1 = 235 \text{ K}$ $V_1 = 450 \text{ mL}$
 - $P_2 = 395 \text{ torr}$ $T_2 = 350 \text{ K}$ $V_2 = ?$

Lesson Objective: Use Avogadro's law to calculate volume-number of moles changes at constant temperature and pressure.

27. Avogadro's Law relates volume and amount of gas at
- constant temperature and pressure
 - constant temperature
 - constant pressure
 - constant velocity
28. True/False: Avogadro's Law allows determination of the quantity of gas.
29. Perform the following calculations:
30. A 4.5 L balloon holds 2.5 mol of gas. If 1.6 mol of gas are added, what is the new volume of the balloon?

14.3 Ideal Gases

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Know the ideal gas law, and know which of the different values for the ideal gas constant to use in a given situation.

- One of the following is not dealt with in the ideal gas law
 - pressure
 - number of moles
 - molecular mass
 - temperature
- The parameter R in the ideal gas equation stands for
 - ideal gas variable
 - ideal gas constant
 - ideal gas value
 - ideal gas correction
- True/False: The ideal gas equation is $PV = mRT$.
- True/False: R is the ideal gas constant.
- To evaluate the gas constant, it is necessary to use _____ for the temperature and it is conventional to use units of _____ for volume.
- When pressure is measured in kPa, the value for R is _____.

Lesson Objective: Use the ideal gas law to calculate the pressure, volume, temperature, or number of moles of an ideal gas when the other three quantities are known.

- A joule equals
 - kPa/liter
 - $kPa \cdot L$
 - kPa/ml
 - $kPa \cdot mL$
- One of the following is not a value for R
 - $8.314 J/K \cdot mol$
 - $62.36 L \cdot mm Hg/K \cdot mol$
 - $5.34 L \cdot kPa/K \cdot mol$
 - $0.08206 L \cdot atm/K \cdot mol$

9. True/False: Kelvin temperatures must be used in evaluating the gas constant.
10. True/False: All pressure units give the same value for R .
11. What volume is occupied by 23.4 mol hydrogen at a temperature of 317 K and a pressure of 719 mm Hg?
12. What is the pressure in atm of 16.7 mol helium that occupies a volume of 7.2 L at a temperature of 279 K?

Lesson Objective: Use the ideal gas law to calculate the molar mass or the density of a gas.

13. The molar mass of an unknown gas can be determined by
 - a. determining mass at STP
 - b. determining volume at STP
 - c. determining mass and volume at known temperature and pressure
 - d. determining mass and volume at STP
14. The number of moles of an unknown gas can be calculated using
 - a. $n = \frac{PV}{RT}$
 - b. $n = PVRT$
 - c. $n = \frac{RT}{PV}$
 - d. $n = \frac{PVR}{T}$
15. True/False: In calculating molar mass, it is assumed that the gas is ideal.
16. True/False: moles of gas/mass = molar mass.
17. A gas had a mass of 0.8396 g with a volume of 0.35 L at a pressure of 745 mm Hg and a temperature of 370 K. Calculate the molar mass.
18. What is the density of carbon tetrachloride vapor at 714 torr and 398 K?

Lesson Objective: Use the ideal gas law in stoichiometry problems involving gases that are not at STP.

19. One of the following is not needed to perform gas stoichiometry problems
 - a. a balanced equation
 - b. conditions of standard temperature and pressure
 - c. the amount of the starting material
 - d. the amount of product formed
20. Gas stoichiometry calculations require all of the following except
 - a. identity of reactants
 - b. formation of a gaseous product
 - c. formula mass of reactants
 - d. formula mass of products
21. True/False: Gas stoichiometry calculations can be performed using any set of conditions.
22. True/False: In combustion reactions, the products are carbon dioxide and water.
23. What volume of CO_2 is produced from the reaction of 36.7 g acetone at $76^\circ C$ and a pressure of 645 mm Hg?.

$$CH_3-\underset{\begin{array}{c} || \\ O \end{array}}{C}-CH_3 + 4 O_2 \rightarrow 3 CO_2 + 3 H_2O$$
24.

$$2 CH \equiv CH + 5 O_2 \rightarrow 4 CO_2 + 2 H_2O$$
25. What volume of O_2 is needed to produce 112 g CO_2 from the combustion of acetylene at 303 K and 695 torr?

Lesson Objective: Explain the conditions under which real gases are most ideal or least ideal.

25. A gas deviates from ideal behavior at
- very low pressures
 - very high temperatures
 - very low concentrations of gas particles
 - very high pressures
26. The ideal gas
- has a $\frac{PV}{RT}$ ratio very close to one at all temperatures
 - has strong intermolecular interactions
 - has gas particles crowded close together
 - has little empty space between gas particles
27. True/False: The gas particles of a truly ideal gas can occupy volume.
28. True/False: Continued cooling of a gas converts it to a liquid.
29. Gases whose attractive forces are _____ are more _____ than those with strong attractive forces.
30. An _____ gas is one that follows the gas laws at all conditions of _____ and _____.
31. If the particles are moving at slower _____, the _____ forces between them are more prominent.

14.4 Gas Mixtures and Molecular Speeds

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use Dalton's law and mole fraction to calculate the partial pressure of a gas in a mixture.

- Gas pressure for a mixture of nitrogen and helium is determined by
 - the total number of particles present
 - the pressure of the gas in higher quantity
 - pressure of nitrogen/pressure of helium
 - pressure of helium/pressure of nitrogen
- Mole fraction is defined as
 - moles of one substance/grams of total material
 - moles of one substance/ total moles of mixture
 - total moles of mixture/moles of one substance
 - grams of one substance/total grams of mixture
- Gases in a mixture
 - exert pressure affected by the structure of other gases
 - exert pressure affected by the pressure of other gases
 - exert pressure independently of other gases
 - exert a lower pressure when other gases are present
- True/False: The pressure exerted by one gas in a mixture is called its individual pressure.
- True/False: Nitrogen exerts more pressure in air than does oxygen.
- True/False: Gas pressure is influenced by the number of particles in the container.
- Define partial pressure
- A mixture of gases contains 2.7 moles gas A and 4.1 moles gas B. What is the mole fraction of each gas?
- One liter volumes of two gases are mixed in a one-liter container. Gas A has a pressure of 175 mm Hg and the total pressure of the mixture is 890 mm Hg. What is the pressure of gas B?
- A mixture of 34.5 moles nitrogen and 14.7 moles argon produce a combined pressure of 1420 mm Hg. What is the partial pressure of each gas?

Lesson Objective: Calculate the pressure of a gas that has been collected by water displacement in order to determine the volume of the dry gas.

- Collecting a gas by water displacement involves
 - bubbling water through the gas
 - dissolving the gas in water

- c. passing the gas through water to collect it
 - d. reacting the gas with water
12. The vapor pressure of water
- a. increases with an increase in temperature
 - b. decreases with an increase in temperature
 - c. increases with a decrease in temperature
 - d. is not affected by temperature
13. The actual volume of a gas collected over water
- a. is higher than the volume observed
 - b. is not affected by water vapor
 - c. is lower than the volume observed
 - d. depends on the gas
14. True/False: Water displacement is often used to collect gases in the lab.
15. True/False: The water vapor in a gas will be due to evaporation.
16. True/False: Water vapor can be ignored in calculating gas volumes.
17. _____ law can be used to calculate the amount of the desired gas by _____ the contribution of the water vapor.
18. A gas is collected over water at 25°C . The volume is 3.5 L and the pressure is 785 mm Hg. What is the corrected pressure of the gas?
19. An experiment generates 1.8 L chlorine gas over water at a temperature of 30°C and 698 mm Hg. What is the volume of chlorine gas at STP?

Lesson Objective: Define diffusion and effusion.

20. One of the following statements about diffusion is not true
- a. gases diffuse more rapidly than liquids and solids
 - b. gases diffuse less rapidly than liquids or solids
 - c. liquids diffuse less rapidly than gases
 - d. solids do not diffuse
21. A hydrogen balloon will stop floating
- a. after heavier gases have effused into the balloon
 - b. after nitrogen replaces the hydrogen
 - c. after hydrogen has effused from the balloon
 - d. after hydrogen combines with oxygen to form water in the balloon
22. True/False: Solids diffuse fairly readily.
23. True/False: Effusion involves gas escaping through a tiny hole.
24. _____ is the tendency of molecules to move from an area of _____ concentration to an area of _____ concentration until the concentration is _____.
25. _____ is the process of a _____ gas escaping through a tiny hole in its container.

Lesson Objective: Use Graham's law to calculate the velocity ratios of two gases based on their molar masses.

26. The rate of diffusion for a gas is a function of

- a. the square of the molar mass
 - b. the square root of the molar mass
 - c. the square of the kinetic energy
 - d. the square root of the kinetic energy
27. Graham's Law was developed by
- a. The Welch chemist Thomas Graham
 - b. The British chemist Alexander Graham
 - c. the Scottish chemist Thomas Graham
 - d. the Scottish chemist Alexander Graham
28. True/False: Gases with lower molar masses diffuse more rapidly than gases with higher masses.
29. True/False: Graham's Law applies only to diffusion of a gas.
30. Calculate the relative rates of diffusion of H_2 and He at STP.
31. A gas diffuses 3.98 times slower than H_2 at the same temperature and pressure. What is the molar mass of the gas?

CHAPTER **15**

Water Worksheets

Chapter Outline

- 15.1 PROPERTIES OF WATER
 - 15.2 AQUEOUS SOLUTIONS
 - 15.3 COLLOIDS AND SUSPENSIONS
-

15.1 Properties of Water

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the structure and polarity of a water molecule.

- Bonds in a water molecule are
 - polar
 - polar covalent
 - covalent
 - ionic
- One of the following statements about the structure of water molecules is not true
 - oxygen is more electronegative than hydrogen
 - one oxygen atom is connected to two hydrogen atoms
 - the oxygen atom has a δ^+ charge
 - the H-O-H bond angle is 105°
- The three-dimensional configuration of the water molecule is due to
 - repulsions between the two H atoms
 - attractions between O and H
 - sp^3 hybridization
 - two sets of lone pair O electrons
- True/False: In a water molecule, O has a δ^- charge.
- True/False: The water molecule has a bent shape.
- True/False: The water molecule as a whole is non-polar.
- Because of the _____ electronegativity of the oxygen atom, the bonds are _____ covalent.
- The _____ shape of the water molecule is critical because the _____ O-H bonds do not _____ one another.
- Polar molecules attract one another by _____ forces.

Lesson Objective: Describe the hydrogen bonding that occurs in water and ice.

- In a water molecule, the hydrogen atoms
 - have little electron density
 - are attracted to each other
 - function as dipoles
 - are unreactive
- The bent structure of the water molecule
 - puts strain on the H-O bonds
 - makes the molecule polar
 - decreases interactions with other molecules

4. makes the O atom polar
12. The geometry around each oxygen atom in water is approximately
 1. tetrahedral
 2. square planar
 3. biplanar
 4. octahedral
13. True/False: Hydrogen atoms are strongly attracted to lone pair O electrons on adjacent water molecules.
14. True/False: Hydrogen bonds are weaker than conventional dipole-dipole interactions.
15. True/False: Each O atom can form hydrogen bonds with one other water molecule.
16. The geometry around each oxygen atom consists of two _____ bonds and two _____ bonds.
17. The geometry of the water molecule shows bond angles of _____ around each O atom. The bond angle for a “pure” tetrahedron is _____ .

Lesson Objective: Discuss the unique properties of water and ice.

18. Water reaches its highest density at approximately
 1. 0°C
 2. 4°C
 3. 12°C
 4. 6°C
19. As water freezes, the hydrogen bonds gradually form a _____ network.
 1. tetragonal
 2. octahedral
 3. hexagonal
 4. tetrahedral
20. Hydrogen bonding is responsible for all the following properties of water except
 1. high surface tension
 2. high melting point
 3. low vapor pressure
 4. low boiling point
21. When ice forms, one of the following occurs
 1. hydrogen bonds rearrange to produce molecules with less space between them
 2. hydrogen bonds no longer exist
 3. hydrogen bonds rearrange to produce molecules with more space between them
 4. hydrogen bonds move closer together
22. True/False: Between 4°C and 0°C, the density of water gradually increases.
23. True/False: Most molecular compounds with a molar mass similar to water are gases at room temperature.
24. True/False: Water’s high vapor pressure keeps lakes and rivers from evaporating too fast.
25. True/False: The hydrogen bonds in liquid water constantly break and reform.
26. Liquid water at 0°C has a density of _____ g/cm³, while solid water (ice) at the same temperature has a density of _____ g/cm³.
27. A great deal of _____ is required to break apart the _____ network of _____ and return it to the _____ state.
28. Surface tension can be seen by the curved _____ that forms when water is in a thin _____, such as a graduated _____ or a _____.

15.2 Aqueous Solutions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define a solution and describe the parts of a solution.

1. A true solution has all of the following characteristics except one:
 1. solute does not settle out of solution.
 2. solute can be removed by filtration.
 3. solute can be solid, liquid, or gas
 4. solution is stable
2. Solute particles are
 1. limited to ions
 2. very large, greater than 1 mm
 3. too small to measure
 4. usually less than 1 nm
3. True/False: Aqueous solutions can be made only with solids.
4. True/False: Solute particles can be atoms, ions, or molecules
5. Define the following terms:
 1. solution
 2. solvent
 3. solute

Lesson Objective: Describe how an aqueous solution is formed from both ionic compounds and molecular compounds.

6. All of the following are true statements about aqueous solutions except
 1. water is a polar material
 2. water molecules are in continuous motion
 3. polar water molecules surround the crystal lattice
 4. water is attracted to the ions in NaCl
7. When a sugar molecule dissolves in water
 1. water forms hydrogen bonds with the sugar molecules
 2. water interacts with the sugar ions
 3. the sugar molecule bonds are broken
 4. the intermolecular forces for the sugar molecules are not affected.
8. True/False: The Na^+ is attracted to the oxygen end of the water molecule.
9. True/False: Solute particles surround the water molecules.
10. Hydration is the process of _____ particles being _____ by water molecules arranged in a specific manner.

11. _____ helps to _____ aqueous solutions by preventing the _____ and _____ ions from coming back together and forming a _____.

Lesson Objective: Explain the reasons why some compounds are insoluble in water.

12. When a nonpolar liquid is mixed with water
1. the water molecules surround the nonpolar molecules
 2. the two liquids mix together
 3. two layers are formed
 4. the nonpolar liquid dissolves in the water
13. Ethanol dissolves in water because
1. the ethanol molecule can ionize
 2. ethanol is a polar molecule
 3. ethanol can dissociate into smaller components
 4. ethanol is light enough to dissolve
14. True/False: A nonpolar molecule dissolves in water by breaking up some of the hydrogen bonds between adjacent water molecules.
15. True/False: All ionic compounds are soluble in water.
16. Define the following terms
1. miscible
 2. immiscible
17. Calcium carbonate and silver chloride are nearly insoluble in water because the attractions between the ions in the crystal _____ are stronger than the _____ that the water molecules have for the ions.

Lesson Objective: Write equations for the dissociation of ionic compounds in water.

18. In writing equations for polyatomic ion dissociation
1. the subscripts always indicate the ions to be formed
 2. all polyatomic ions dissociate into their components
 3. polyatomic ions do not dissociate
 4. compounds containing polyatomic ions do not dissociate
19. The equation for dissolving of a nonelectrolyte (NE) in water is
1. $NE \rightarrow NE(aq)$
 2. $NE \rightarrow NE(s)$
 3. $NE \rightarrow N + E$
 4. $NE \rightarrow N(aq) + E(aq)$
20. True/False: When NO_3^- dissolves in water, it forms N and O ions.
21. True/False: Ammonium phosphate forms ammonium ions and phosphate ions in aqueous solution.
22. Write dissociation equations for the following strong electrolytes:
1. KBr
 2. Li_2SO_4
 3. NaOH
 4. $\text{Ba}(\text{OH})_2$
 5. MgCl_2

Lesson Objective: Distinguish between strong electrolytes, weak electrolytes, and nonelectrolytes.

23. A strong electrolyte shows all of the following characteristics except

1. mostly ionized in water
 2. solution conducts electricity well
 3. can be a nonelectrolyte in the pure state
 4. seen in all polar compounds
24. Aqueous nitrous acid is
1. a weak electrolyte, forming only 5% ions
 2. a moderate electrolyte, forming 40% ions
 3. a strong electrolyte, forming 95% ions
 4. a nonelectrolyte
25. True/False: The equation for the ionization of a weak electrolyte includes a double arrow.
26. True/False: Some polar molecules become weak electrolytes when dissolved in water.
27. Define the following terms:
1. strong electrolyte
 2. weak electrolyte
28. HCL is a _____ in the ___ phase, but becomes an _____ when dissolved in _____.

15.3 Colloids and Suspensions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the properties of a suspension.

- Sand in ocean water is an example of
 - suspension
 - solution
 - colloid
 - emulsion
- Particles in a suspension are
 - smaller than those in a solution
 - smaller than colloid particles
 - about 100 times the size of a solution particle
 - about 1000 times the size of a solution particle
- Suspension particles are
 - molecules
 - large particles
 - large molecules
 - small molecules
- One of the following is not a property of a suspension
 - heterogeneous mixture
 - particles settle out
 - particles cannot be filtered out
 - particles scatter light
- True/False: A suspension is a homogeneous mixture.
- True/False: Suspensions require active mixing to stay suspended.
- True/False: Suspensions do not scatter light.
- True/False: Suspensions consist of large particles or aggregates.
- A _____ is a _____ mixture in which some of the particles settle out of the mixture upon standing.
- The particles in a _____ are far larger than those of a _____, so _____ is able to pull them down out of the _____ (water).

Lesson Objective: Describe the properties of a colloid and distinguish from a solution or a suspension.

- Colloids have all of the following properties except
 - particles do not separate on standing
 - particles can be filtered out

3. heterogeneous mixture
 4. particles scatter light
12. All of the following are examples of colloids except
1. colorless solution
 2. fog
 3. smoke
 4. milk
13. An emulsion is
1. a mixture of a suspension and a colloid
 2. a colloidal dispersion of a gas in a solid
 3. a colloidal dispersion of a solid in a liquid
 4. a colloidal dispersion of a liquid in a liquid or a solid
14. The Tyndall effect is
1. the scattering of visible light by a solution
 2. the scattering of visible light by a colloid
 3. the scattering of visible light by a suspension
 4. the scattering of visible light by a solid
15. One of the following is not a colloidal system
1. fog
 2. clouds
 3. smoke
 4. snow
16. A liquid emulsion is a dispersion of
1. a liquid in a gas
 2. a liquid in a solid
 3. a liquid in a liquid
 4. a solid in a liquid
17. True/False: The Tyndall effect is seen when light passes through dust particles in the air.
18. True/False: Suspensions exhibit Brownian motion.
19. True/False: Egg yolk is used to stabilize an oil-vinegar mixture.
20. True/False: Marshmallow is an example of a foam colloid.
21. True/False: Colloidal particles are larger than suspension particles.
22. True/False: Dissolved particles in a solution are too small to scatter light.
23. Define the following terms:
1. colloid
 2. Tyndall effect
 3. emulsion
24. What is Brownian motion? What causes it?
25. Fill in the table below:

TABLE 15.1:

Class of Colloid	Dispersed Phase	Dispersion Medium	Examples
Sol and gel		liquid	paint, jellies, blood, gelatin, mud
	solid	gas	
Solid emulsion	liquid		cheese, butter

TABLE 15.1: (continued)

Class of Colloid	Dispersed Phase	Dispersion Medium	Examples
Liquid emulsion		liquid	
Liquid aerosol	liquid	gas	
	gas	solid	marshmallow
Foam	gas		whipped cream, shaving cream

TABLE 15.2:

Class of Colloid	Dispersed Phase	Dispersion Medium	Examples
Sol and gel	solid	liquid	paint, jellies, blood, gelatin, mud
Solid aerosol	solid	gas	smoke, dust in air
Solid emulsion	liquid	solid	cheese, butter
Liquid emulsion	liquid	liquid	milk, mayonnaise
Liquid aerosol	liquid	gas	fog, mist, clouds, aerosol spray
Foam	gas	solid	marshmallow
Foam	gas	liquid	whipped cream, shaving cream

CHAPTER

16**Solutions Worksheets****Chapter Outline**

- 16.1 SOLUBILITY**
 - 16.2 SOLUTION CONCENTRATION**
 - 16.3 COLLIGATIVE PROPERTIES**
 - 16.4 NET IONIC EQUATIONS**
-

16.1 Solubility

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: List examples of solutions made from different solute-solvent combinations.

- Alcohol dissolved in water is an example of a _____ solution.
 - solid in liquid
 - liquid in liquid
 - gas in liquid
 - liquid in gas
- An example of a liquid in solid solution is
 - water in air
 - zinc in copper
 - mercury in silver and tin
 - salt in water
- True/False: Carbon dioxide in water is an example of a liquid in gas solution.
- True/False: Many solutions do not involve water.
- Air is a _____ mixture of many different gases.
- Solid-solid solutions, such as brass, bronze, and _____ silver, are called _____.

Lesson Objective: List and explain three factors that affect the rate of dissolving of a solid solute in a liquid solvent.

- The rate at which a solute dissolves depends upon the _____ of the solute particles.
 - size
 - molecular mass
 - crystal structure
 - configuration
- Temperature affects all of the following parameters except
 - kinetic energy of the system
 - increased frequency of collision between solute and solvent
 - increased energy of collision
 - change in molecular configuration
- True/False: Stirring a solute affects the overall amount dissolved.
- True/False: A greater surface area decreases the amount of material dissolved.
- Dissolving is a _____ phenomenon, since it depends on _____ molecules colliding with the outer surface of the _____.
- A temperature change not only affects the _____ of dissolving, but it also affects the _____ of solute that can be dissolved.

Lesson Objective: Explain solution equilibrium and distinguish between saturated, unsaturated, and super-saturated solutions.

13. The addition of NaCl to water
 1. can be done indefinitely
 2. will reach a saturation point
 3. will reach an equilibrium
 4. allows NaCl to redissolve
14. When NaCl is added to water, all of the follow occur except
 1. ions break off from the salt crystal
 2. ions are hydrated by water molecules
 3. ions precipitate
 4. ions recombine with the salt crystal
15. True/False: Some dissolved ions can recombine with the crystal.
16. True/False: More NaCl can dissolve if more water is added to the solution.
17. Define the following terms:
 1. recrystallization
 2. saturated solution
 3. unsaturated solution

Lesson Objective: Explain the effects of temperature on the solubility of solids and gases. Use a solubility curve to determine the solubilities of substances at various temperatures.

18. The solubility of gases
 1. increases as temperature increases
 2. decreases as temperature increases
 3. is not affected by temperature
 4. varies depending on the gas
19. At 50°C, the solubility of KNO₃ is approximately _____ g/100 g water.
 1. 82
 2. 108
 3. 62
 4. 75
20. True/False: The solubility of most solid substances increases as the temperature increases.
21. True/False: Recrystallization from a supersaturated solution is very slow.
22. Define the following terms:
 1. solubility
 2. supersaturated
23. When a solvent with a _____ dissolved in it is heated, the _____ energy of both the _____ and _____- increases.
24. The seed crystal provides a _____ site on which the _____ dissolved crystals can begin to grow.

Lesson Objective: Use Henry's law and explain the effect of pressure on the solubility of gases.

25. The relationship between gas solubility and pressure was developed by
 1. the French chemist Guillaume Henri
 2. the German chemist Wilhelm Henry

3. the English chemist William Henry
 4. the Scottish chemist Henry Williams
26. As pressure above the solvent _____, the amount of dissolved gas _____.
1. increases, decreases
 2. increases, increases
 3. decreases, increases
 4. decreases, not affected
27. True/False: Pressure has very little effect on the solubility of solids or liquids.
28. True/False: Solubilities of gases are usually reported in grams/mL.
29. State Henry's Law.
30. At 20°C, argon has a solubility in water of 62 mg/L at 760 torr. How much argon will dissolve at a pressure of 4180 torr at 20°C?
31. Two sealed soda pop cans contain a soft drink under pressure. One of the cans is placed in the refrigerator at 5°C. The other can is placed on the outdoor table at 32°C. After two hours, both of the cans are opened. Which one is more likely to have foam from the soft drink spill out? Explain your answer.

16.2 Solution Concentration

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use the terms concentrated and dilute to describe the relative concentration of a solution.

1. A concentrated solution is one that
 1. has a relatively small amount of solute
 2. has an unknown amount of solute
 3. has a relatively large amount of solute
 4. thinks a lot
2. A dilute solution is one that
 1. has a relatively small amount of solute
 2. has an unknown amount of solute
 3. has a relatively large amount of solute
 4. has had solvent removed
3. True/False: A concentrated solution contains a specified amount of solute.
4. True/False: Adding more water to a solution dilutes it.
5. Define concentration
6. If you were to _____ a solution, causing the solvent to evaporate, you would be _____ it, because the ratio of solute to solvent would be _____.

Lesson Objective: Calculate the concentration of a solution as either a mass percent or a volume percent.

7. Mass percent is used when
 1. the solute is a liquid
 2. when the solvent is a solid
 3. when the solute is a solid
 4. when the solute is a gas
8. Volume percent is calculated using the following formula
 1. $\text{volume solute} \div \text{volume solvent}$
 2. $\text{volume solute} \div \text{volume solution}$
 3. $\text{volume solvent} \div \text{volume solute}$
 4. $\text{volume solution} \div \text{volume solute}$
9. True/False: Ingredients in medicines and foods are often expressed as percentages.
10. True/False: When adding liquid solute and solvent together, volume is always conserved.
11. Describe how to prepare 400 grams of a 15% (mass/mass) aqueous solution of KBr.
12. Describe how to prepare 500 mL of a 70% (volume/volume) aqueous solution of acetic acid.

Lesson Objective: Calculate the molarity of a solution. Use molarity to calculate the mass of solute needed to prepare a particular solution.

13. Molarity is determined by
 1. mass of solute \div volume of solvent
 2. moles of solute \div mL solvent
 3. moles of solute \div liters of solution
 4. moles of solute \div liters of solvent
14. The symbol for molarity is
 1. m
 2. mol
 3. Mol
 4. M
15. True/False: Hydrated ions take up space in a solution.
16. True/False: Solutions are best prepared in a calibrated beaker.
17. A solution is prepared by dissolving 22.7 g NaCN in enough water to make 250 mL of solution. What is the molarity?
18. Describe how to prepare 500 mL of 0.25 M KCl.

Lesson Objective: Calculate the molarity of a diluted solution. Use the dilution equation to calculate the volume of a concentrated stock solution required for a particular dilution.

19. Volume units for dilution calculations
 1. must be in L
 2. must be in mL
 3. both sides of equation must have same units
 4. units on both sides of equation do not need to agree
20. Use of a _____ solution is common in preparing dilutions of lower concentration.
 1. base
 2. stock
 3. preparative
 4. working
21. True/False: A calibrated pipet improves the accuracy of preparing dilutions.
22. True/False: The molarity of a diluted solution can be calculated using $\frac{M_1V_1}{V_2} = M_2$
23. 150 mL of 0.05 M HCl is diluted to 750 mL. What is the new molarity?
24. A solution of 0.25 M NaOH is used to prepare 0.1 M NaOH. How much of the 0.25 M solution is needed to prepare 500 mL of the diluted NaOH?

Lesson Objective: Calculate the molality of a solution and distinguish molality from molarity.

25. Molality is defined as moles of solute divided by
 1. kg solvent
 2. kg solution
 3. grams solvent
 4. grams solution
26. The symbol for molality is
 1. m
 2. mol

3. m
 4. mol
27. True/False: Molality values are not affected by temperature.
 28. True/False: The choice of solvent has no effect on the value for the molality of a solution.
 29. What is the molality of 35 g acetone dissolved in 0.85 kg water? The molar mass of acetone is 58.1 g/mol.
 30. Calculate the molality of a solution of 75.4 g glucose (molar mass 180.1 g/mol) dissolved in 460 mL water.

16.3 Colligative Properties

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define a colligative property and identify three colligative properties of solutions.

- One of the following is not a colligative property of water
 - vapor pressure
 - freezing point
 - solubility
 - boiling point
- When a liquid evaporates readily, it has
 - a low vapor pressure
 - few molecules in the gas phase
 - a high boiling point
 - a high vapor pressure
- True/False: The vapor pressure of water depends on the number of particles dissolved in it.
- True/False: Nonvolatile solutes readily escape into the vapor phase.
- A _____ is a property of a solution that depends only on the _____ of solute particles dissolved in the solution and not on their _____.

Lesson Objective: Differentiate between the effects that an electrolyte has on the colligative properties of a solution compared to a nonelectrolyte.

- Vapor pressure lowering is best accomplished by a material that is all of the following except
 - nonvolatile
 - does not undergo chemical reaction
 - slightly soluble
 - does not escape into the gaseous phase
- Which of the following materials will have the greatest effect on colligative properties of a solution?
 - CH_3Cl
 - NaCl
 - $(\text{NH}_4)_2\text{SO}_4$
 - $\text{CH}_3\text{CH}_2\text{OH}$
- True/False: The lowering of vapor pressure depends on the number of dissolved solute particles.
- True/False: The chemical nature of the solute is an important influence on vapor pressure lowering.
- Ionic compounds are strong _____ that _____ into ions when they dissolve.
- Methanol (CH_3OH) has a freezing point of -97.6°C . Would you expect methanol to be an effective material for lowering the freezing point of water? Explain your answer.

12. Methanol has a boiling point of 64.7°C . Would you expect methanol to be an effective material for raising the boiling point of water? Explain your answer.

Lesson Objective: Calculate the freezing and boiling points of a solution of known molality.

13. Molal solutions are used in freezing point and boiling point calculations because
1. they are easier to calculate than molar solutions
 2. the concentration of molar solutions is affected by temperature
 3. the concentration of molal solutions is affected by pressure
 4. molal concentrations are more commonly used in chemistry than molar concentrations
14. The solvent with the highest freezing point depression constant is
1. water
 2. naphthalene
 3. camphor
 4. phenol
15. A solute affects the freezing point of a solvent by
1. decreasing the kinetic energy of the solvent molecules
 2. increasing the order of the solvent molecules
 3. increasing the kinetic energy of the solvent molecules
 4. disrupting the order of solvent molecules
16. True/False: Methylene glycol is used in many commercial antifreeze products.
17. True/False: NaCl and methanol provide the same degree of boiling point elevation.
18. True/False: Addition of a solute decreases the vapor pressure of the solvent.
19. Ethylene glycol has a freezing point of -12.9°C and a boiling point of 197.3°C . Explain why this material is used both to prevent radiator water from freezing in the winter and lower the possibility of overheating in the summer.
20. Calculate the freezing point of a solution of 125 g KBr in 450 g water.
21. Calculate the boiling point of the solution described in problem 20.
22. Since a solution of KBr would probably cause significant damage to the car radiator and engine, you decide to use 125 g of the nonelectrolyte octane (molar mass 114 g/mole). Will this have a greater or less effect on freezing point depression of the solution?

Lesson Objective: Use freezing or boiling point information to calculate the molar mass of an unknown solute.

23. In order to calculate molar mass from freezing or boiling point data, we need to know all of the following except
1. the K_f or K_b of the solvent
 2. the molar mass of the solvent
 3. the type of solute –electrolyte or not
 4. the number of particles in solution
24. Once we measure the freezing point depression, the next step is to
1. calculate K_f
 2. calculate molar mass
 3. calculate molality of the solution
 4. determine moles of solute
25. If the value for the number of ions formed is in error, then the molar mass will
1. be lower than the real value
 2. be higher than the real value

3. not be affected by the real value
 4. depends on whether the value used is higher or lower than the real value
26. True/False: The mass of the solute is needed in order to calculate molar mass.
 27. True/False: Only freezing point data can be used to determine molar mass.
 28. True/False: Moles of solute can be calculated using $m \times kg$ solvent.
 29. The freezing point of a solution that contains 1.00 g of an unknown compound dissolved in 10.0 g of benzene is found to be 2.07 C° . The freezing point of pure benzene is 5.48 C° . The molal freezing point depression constant of benzene is $5.12\text{ C}^\circ/\text{molal}$. What is the molar mass of the unknown compound?
 30. Boiling point elevation could be used for determinations of molar mass. What are two problems with using boiling point instead of freezing point?
 31. Further investigation of the compound described in problem 30 shows that it is actually an electrolyte, although the number of particles formed is not yet known. What effect will this information have on acceptance of the calculated molar mass?

16.4 Net Ionic Equations

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Write net ionic equations for double-replacement reactions that produce precipitates, gases, or molecular compounds.

- A molecular equation is one in which
 - only molecules react
 - molecules are formed from ions
 - the equation is written as if all substances exist as molecules
 - ions are formed from molecules
- A dissolved ionic compound is indicated by
 - (AQ)
 - (aq)
 - (hyd)
 - (a)
- A net ionic equation
 - shows all the ions involved
 - eliminates ions that form precipitates
 - shows only product ions
 - only shows ions actually involved in the reaction
- True/False: A spectator ion does not participate in the reaction.
- True/False: Net ionic equations must be balanced by both mass and charge.
- True/False: The abbreviation (s) stands for solution.
- Define the following:
 - molecular equation
 - ionic equation
 - net ionic equation
- Classify the following reactions as molecular, ionic, or net ionic equations:
 - $CH_2 = CH_2 + H_2 \xrightarrow{Pt} CH_3CH_3$
 - $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
 - $Ag^+(aq) + NO_3^-(aq) + Rb^+(aq) + Cl^-(aq) \rightarrow AgCl(s) + Rb^+(aq) + NO_3^-(aq)$
- Write the molecular equation, the ionic equation, and the net ionic equation for the reaction between sodium phosphate and calcium chloride
- Write the molecular equation, the ionic equation, and the net ionic equation for the reaction between ammonium carbonate and aluminum nitrate.
- Write the molecular equation, the ionic equation, and the net ionic equation for the neutralization reaction between sodium hydroxide and sulfuric acid.

Lesson Objective: Write net ionic equations for single-replacement reactions.

12. A single replacement reaction involves a _____ as a reactant.
1. metal
 2. gas
 3. precipitate
 4. vapor
13. All of the following are single-replacement reactions except
1. metal replacement
 2. water replacement
 3. hydrogen replacement
 4. halogen replacement
14. The overall charge in a single-replacement reaction
1. is always zero
 2. is always +2
 3. depends on the reaction
 4. will be different for reactants and products
15. True/False: A metal in a single-replacement reaction is always a spectator ion.
16. True/False: A neutral element does not carry a charge.
17. True/False: A metal is always replaced in a single-replacement reaction.
18. Complete the following single-replacement reactions and balance the equations:
1. $Al + Pb(NO_3)_2 \rightarrow$
 2. $Br_2 + CaI_2 \rightarrow$
 3. $Al + HCl \rightarrow$
 4. $Zn + AgNO_3 \rightarrow$
 5. $Zn + H_2SO_4 \rightarrow$

Lesson Objective: Use the solubility rules to predict precipitate formation.

19. One of the following is completely soluble in water
1. AgCl
 2. AgNO₃
 3. AgS
 4. Ag₂SO₄
20. All of the following compounds are fairly insoluble except
1. Na₂S
 2. BaS
 3. CaS
 4. Fe₂S₃
21. One of the following compounds is soluble in water
1. Fe(OH)₃
 2. Zr(OH)₂
 3. Ba(OH)₂
 4. Mn(OH)₂
22. All of the following compounds are soluble except
1. NaBr
 2. KBr

3. CaCl_2
4. PbCl_2
23. True/False: Calcium carbonate is mostly insoluble in water.
24. True/False: All alkali metal compounds are soluble.
25. True/False: Mercury(I) chloride is fairly water-soluble.
26. True/False: Copper(II) carbonate dissolves readily in water.
27. Write the products of the following reactions and balance the equations. If an insoluble product is formed, indicate with (s):
 1. $\text{AlI}_3 + \text{HgCl}_2 \rightarrow$
 2. $\text{AgNO}_3 + \text{K}_3\text{PO}_4 \rightarrow$
 3. $\text{Ca}(\text{OH})_2 + \text{H}_3\text{PO}_4 \rightarrow$

CHAPTER **17**

Thermochemistry Worksheets

Chapter Outline

- 17.1 HEAT FLOW
 - 17.2 THERMOCHEMICAL EQUATIONS
 - 17.3 HEAT AND CHANGES OF STATE
 - 17.4 HESS'S LAW
-

17.1 Heat Flow

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe how chemical potential energy relates to heat and work.

- The release of chemical potential energy produces all of the following except
 - heat
 - work
 - motion
 - cooling
- Chemicals in gasoline contain a large amount of potential energy because
 - there are many chemical bonds present
 - gasoline is very volatile
 - gasoline components carry out several reactions
 - there are many compounds present in gasoline
- All of the following are examples of work except
 - moving a car
 - thinking about a problem
 - carrying a box
 - shoveling dirt
- True/False: Objects need to be in contact in order to transfer heat.
- True/False: Heat moves from an object at lower temperature to an object at higher temperature.
- True/False: The ability to monitor heat is important for an understanding of chemical processes.
- Define the following terms:
 - chemical potential energy
 - thermochemistry
 - heat
- Energy is the capacity for doing _____ or supplying _____.

Lesson Objective: Describe the law of conservation of energy and how heat flows between system and surroundings during both endothermic and exothermic processes.

- In carrying out a chemical reaction, all of the following are part of the system except
 - the chemicals in the reaction
 - the beaker holding the chemicals
 - the lab bench where the reaction is occurring
 - the heating device to warm the reaction
- An example of an exothermic reaction is

1. baking a cake
 2. burning a candle
 3. warming a room
 4. melting ice
11. An example of an endothermic reaction is
1. making coffee
 2. solar flares
 3. house fire
 4. sodium in water
12. True/False: The wood in a fireplace is part of the system.
13. True/False: Endothermic reactions do not need heat from the surroundings.
14. True/False: Cooking is an endothermic process.
15. Define the following:
1. endothermic
 2. exothermic
 3. law of conservation of energy
16. When _____ or _____ changes occur, they are generally accompanied by a transfer of _____.

Lesson Objective: Use the specific heat equation to perform calculations that relate mass, specific heat, change in temperature, and the amount of heat absorbed or released.

17. A unit of heat is the _____
1. *joule/sec*
 2. *joule/kg*
 3. *joule*
 4. *joule · gram/sec*
18. If a piece of chocolate cake contains 350 Calories, that is the equivalent of
1. 35 calories
 2. 3500 calories
 3. 35,000 calories
 4. 350,000 calories
19. One calorie equals _____ joules
1. 4.184
 2. 3.789
 3. 4.254
 4. 3.850
20. Which of the following will absorb the most heat from the sun in one hour?
1. Lake Erie
 2. Gulf of Mexico
 3. Pacific Ocean
 4. Mississippi River
21. True/False: The specific heat is the amount needed to raise one gram of material 1°C.
22. True/False: The specific heat of liquid water and of ice are the same.
23. True/False: Coastal climates are more moderate than climates inland.
24. True/False: Water is a poor coolant for engines.
25. How many joules are there in 145 calories?

26. A 150 g sample of water absorbs heat as the temperature increases from 14.7°C to 26.5°C. Calculate the number of joules of heat involved.
27. How much energy must be absorbed by 25.0 g of water to increase its temperature from 55°C to 96.0°C?
28. A sample of iron is heated from 15°C to 45°C and absorbs 1680 joules of energy. What is the mass of iron involved?

17.2 Thermochemical Equations

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define enthalpy, and know the conditions under which the enthalpy change in a reaction is equal to the heat absorbed or released.

- When doing thermochemical calculations, the pressure is
 - constant
 - lowered
 - increased
 - varied
- The enthalpy change is designated as
 - ΔG
 - ΔE
 - ΔH
 - ΔT
- The symbol q stands for
 - temperature
 - heat
 - energy
 - work
- True/False: Chemical reactions in the laboratory are run with controlled pressure.
- True/False: Changes in enthalpy are measured as reactants are converted to products.
- True/False: Heat is released or absorbed in chemical reactions.
- Define enthalpy.
- Why do we measure changes in enthalpy instead of enthalpy?

Lesson Objective: Describe the principles behind calorimetry, and be able to calculate the heat absorbed or released during a process that occurs in a calorimeter.

- An inexpensive calorimeter can be made out of
 - soda pop cans
 - plastic cups
 - foam cups
 - aluminum foil
- The easiest reactions to measure in an inexpensive calorimeter are
 - gases
 - solid-solid combinations

3. nonaqueous materials
 4. liquids
11. Using a calorimeter, the property measured is
1. change of state
 2. temperature change
 3. formation of a gas
 4. formation of a solid
12. True/False: A lid is used to limit heat exchange between materials in the cup and surrounding air.
13. True/False: The dissolved materials are the surroundings.
14. True/False: Temperatures of solutions are measured before they are mixed.
15. Define the following terms
1. calorimetry
 2. calorimeter
16. Calculate ΔH if a piece of metal with a specific heat of $0.98 \text{ J/g} \cdot ^\circ\text{C}$ and a mass of 20 g is heated from 22°C to 28°C .
17. A reaction was run between 50 mL of nitric acid and 50 mL of KOH, both at 0.5 M. A temperature change from 20°C to 34°C was observed. Assume the density of each solution is 1.00 g/mL and that the specific heat of each solution was that of pure water. Calculate the enthalpy change for this reaction.
18. A piece of metal is heated and dropped into a calorimeter holding 50.0 mL of water at 20°C . The temperature increases to 28.7°C . How many joules of energy were released?

Lesson Objective: Write and solve problems with thermochemical equations.

19. An exothermic reaction has all of the following except
1. heat is released
 2. heat released is written on the product side of the reaction
 3. energy is in kJ
 4. energy goes from surroundings to system
20. A thermochemical equation includes
1. specific temperature change
 2. enthalpy change for the reaction
 3. pressure change
 4. time of reaction
21. True/False: The enthalpy change for an exothermic reaction is negative.
22. True/False: Physical states of reactants and products do not affect enthalpy values.
23. Define the following:
1. thermochemical equation
 2. heat of reaction
- a. In the reaction $\text{B}_2\text{H}_6(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow \text{B}_2\text{O}_3(\text{s}) + 3\text{H}_2\text{O}(\text{g})$ $\Delta H = -2035 \text{ kJ}$
 - b. One reaction formed 1.5 moles
 - c. B_2O_3
 - d. . What is the
 - e. ΔH
 - f. for the reaction?
 - g. Another experiment used sufficient materials to form 46.7 g B_2O_3 . What is the ΔH for this reaction?
25. For the following reaction: $\text{NH}_4\text{NO}_3(\text{s}) \rightarrow \text{N}_2\text{O}(\text{g}) + 2\text{H}_2\text{O}(\text{g})$ $\Delta H = -99 \text{ kJ/mol}$

26. One experiment used 145 grams of reactant. What is the
27. ΔH
28. for the reaction?

17.3 Heat and Changes of State

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the enthalpy change that occurs as a substance changes between the solid and liquid states.

- As an ice cube melts, the temperature of the ice
 - increases
 - decreases
 - does not change
 - fluctuates
- Heat of fusion is strongly influenced by
 - intermolecular forces
 - temperature
 - pressure
 - solvent
- The heat of fusion for water is
 - 40.7 kJ/mol
 - 6.01 kJ/mol
 - 40.7 J/mol
 - 6.01 J/mol
- The heat of fusion deals with the process of
 - liquid converting to solid
 - solid converting to gas
 - solid converting to liquid
 - liquid converting to gas
- True/False: Energy is released when ice melts.
- True/False: The total energy released or absorbed depends on the amount of material involved.
- True/False: The conversion of vapor directly to solid is known as vaporization.
- True/False: The temperature during a phase change remains constant.
- Define the following terms:
 - molar heat of fusion
 - molar heat of solidification
- How much heat is absorbed when 25.7 grams of ice melts?
- What mass of ice can be melted by the application of 745 J of heat?

Lesson Objective: Describe the enthalpy change that occurs as a substance changes between the liquid and gas states.

12. Change from the liquid to the vapor state involves all of the following except
 1. absorption of energy
 2. more space between particles in the vapor state
 3. disruption of intermolecular forces
 4. increase in temperature
13. Oxygen has a lower heat of vaporization than the other materials in the table because
 1. oxygen is already a gas
 2. oxygen has a lower molecular mass
 3. oxygen does not have strong intermolecular interactions
 4. oxygen atoms are covalently bonded
14. True/False: Energy is absorbed when a liquid is converted into a gas.
15. True/False: Conversion of a gas to a liquid involves a release of energy.
16. Define the following:
 1. molar heat of vaporization
 2. molar heat of condensation
17. What mass of ethanol is converted from a liquid to a gas when 35.2 kJ of energy are absorbed?
18. How much energy is released when 45 g steam is converted to water at 100°C?

Lesson Objective: Calculate the enthalpy change involved in the change of state for any amount of a given substance.

19. To calculate the amount of energy involved in heating water from 20°C to 50°C, use the
 1. molar heat of fusion
 2. specific heat of water
 3. molar heat of vaporization
 4. specific heat of steam
20. The process of converting ice at -20°C to steam at 140°C requires _____ steps.
 1. three
 2. four
 3. five
 4. six
21. True/False: The heat absorbed in melting is found using the molar heat of fusion.
22. True/False: The specific heat of water is used to determine the amount of energy need to convert water to steam.

Note to reviewer - this type of problem requires five separate operations plus the final addition. I am counting it as five separate problems.

23. Calculate the heat released when 150 g of steam at 125°C is converted to ice at -19°C

Lesson Objective: Calculate the enthalpy changes involved as substances dissolve in water.

24. The molar heat of solution for NaOH is -445.1 kJ/mol. This value tells us that NaOH 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 temperature significantly while dissolving

25. True/False: Ammonium nitrate releases heat energy when it dissolves.
26. If 5.5 g of CaCl_2 are dissolved in 250 mL of water at 22°C , what will be the final temperature of the water? Assume that the heat capacity of the solution is the same as that of pure water.

17.4 Hess's Law

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use Hess's law of heat summation to add chemical reactions together in such a way as to produce a desired final equation. Calculate the enthalpy change for that final reaction.

- Problems in measuring enthalpy change for a reaction include all of the following except
 - slow reaction
 - reaction is an intermediate step
 - fast reaction
 - side reactions
- Heats of combustion are given for _____ mole(s) of substance
 - two
 - one
 - four
 - three
- Some reactions need to be "balanced" by all of the following except
 - multiply it by a factor
 - subtract a component
 - reverse the direction of the equation
 - double the value of ΔH if coefficients are doubled.
- The law dealing with adding thermochemical equations is called
 - Hessian's law
 - Hess' law of heat summation
 - Heser's Law
 - Hessian's law of heat summation
- Heats of combustion are
 - for one mole of a substance
 - are for two or more moles of a substance
 - are in units of kJ/mol
 - are in units of $kJ \cdot mol$
- To determine the ΔH for acetylene synthesis, all of the following needed to be done except
 - write balanced equations for each process
 - reverse the reaction where acetylene is a reactant
 - assign coefficients
 - reverse the reaction where graphite is a reactant
- True/False: Production of acetylene is an endothermic process.

8. True/False: When an equation is reversed, the sign for ΔH does not change.
9. True/False: If the final reaction has a coefficient of two, the combustion reaction for that ingredient also must have a coefficient of two.
10. True/False: Fractional coefficients cannot be used in combustion reactions.
11. True/False: Enthalpy changes for combustion reactions are easy to measure.
12. True/False: Enthalpy changes for reactions can be determined by indirect methods.
13. State Hess' law.
14. Define heat of combustion.
15. Calculate ΔH for each of the following reactions:
 - a. $2C(s) + H_2(g) \rightarrow C_2H_2(g)$
 - b. Given:
 - c. $C_2H_2(g) + (\frac{5}{2}) O_2(g) \rightarrow 2CO_2(g) + H_2O(l) \quad \Delta H^\circ = -1299.5 \text{ kJ}$
 - d. $C(s) + O_2(g) \rightarrow CO_2(g) \quad \Delta H^\circ = -393.5 \text{ kJ}$
 - e. $H_2(g) + (\frac{1}{2}) O_2(g) \rightarrow H_2O(l) \quad \Delta H^\circ = -285.8 \text{ kJ}$
 - f. $CS_2(l) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$
 - g. Given:
 - h. $C(s) + O_2(g) \rightarrow CO_2(g); \Delta H = -393.5 \text{ kJ/mol}$
 - i. $S(s) + O_2(g) \rightarrow SO_2(g); \Delta H = -296.8 \text{ kJ/mol}$
 - j. $C(s) + 2S(s) \rightarrow CS_2(l); \Delta H = +87.9 \text{ kJ/mol}$
 - k. $C(s, gr) + O_2(g) \rightarrow CO_2(g)$
 - l. Given:
 - m. $SrO(s) + CO_2(g) \rightarrow SrCO_3(s) \quad \Delta H = -234 \text{ kJ}$
 - n. $2SrO(s) \rightarrow 2Sr(s) + O_2(g) \quad \Delta H = +1184 \text{ kJ}$
 - o. $2SrCO_3(s) \rightarrow 2Sr(s) + 2C(s, gr) + 3O_2(g) \quad \Delta H = +2440 \text{ kJ}$
 - p. $3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(aq) + NO(g)$
 - q. Given
 - r. $2NO(g) + O_2(g) \rightarrow 2NO_2(g) \quad \Delta H = -116 \text{ kJ}$
 - s. $2N_2(g) + 5O_2(g) + 2H_2O(l) \rightarrow 4HNO_3(aq) \quad \Delta H = -256 \text{ kJ}$
 - t. $N_2(g) + O_2(g) \rightarrow 2NO(g) \quad \Delta H = +183 \text{ kJ}$

Lesson Objective: Define the standard heat of formation of a compound.

16. Standard conditions for thermochemistry are
 1. 20°C and 101.3 kPa
 2. 25°C and 101.3 kPa
 3. 25°C and 750 mmHg
 4. 20°C and 760 mm Hg
17. All of the following are true about standard states except
 1. the standard state for hydrogen is H_2
 2. the standard state for iron is solid
 3. the standard state for nitrogen is liquid
 4. the standard state for ammonia is gas
18. True/False: The standard state for bromine is liquid.
19. True/False: The standard state for carbon is diamond.
20. Define standard heat of formation.
21. The standard heat of formation for an element is _____.

Lesson Objective: Use known values for standard heats of formation to calculate a previously unknown standard heat of reaction.

22. The standard enthalpy of reaction is represented by
1. ΔH
 2. ΔH_r
 3. ΔH°
 4. ΔH^f
23. The Greek letter Σ stands for
1. standard
 2. sum
 3. significant
 4. signify
24. True/False: Each heat of formation must be multiplied by its coefficient in the equation.
25. True/False: Monoclinic and rhombic sulfur have the same standard heat of formation.
26. Use the attached table to calculate the standard heat of reaction for each of the following:
1. $\text{NaOH}(s) + \text{HCl}(g) \rightarrow \text{NaCl}(s) + \text{H}_2\text{O}(g)$
 2. $2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)$
 3. $\text{CH}_4(g) + 2\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{H}_2\text{O}(l)$
 4. $2\text{H}_2\text{S}(g) + 3\text{O}_2(g) \rightarrow 2\text{H}_2\text{O}(l) + 2\text{SO}_2(g)$
 5. $2\text{NO}(g) + \text{O}_2(g) \rightarrow 2\text{NO}_2(g)$

TABLE 17.1:

Compound	$\Delta H_f(kJ/mol)$	Compound	$\Delta H_f(kJ/mol)$
$\text{CH}_4(g)$	-74.8	$\text{HCl}(g)$	-92.3
$\text{CO}_2(g)$	-393.5	$\text{H}_2\text{O}(g)$	-241.8
$\text{NaCl}(s)$	-411.0	$\text{SO}_2(g)$	-296.1
$\text{H}_2\text{O}(l)$	-285.8	$\text{NH}_4\text{Cl}(s)$	-315.4
$\text{H}_2\text{S}(g)$	-20.1	$\text{NO}(g)$	+90.4
$\text{H}_2\text{SO}_4(l)$	-811.3	$\text{NO}_2(g)$	+33.9
$\text{MgSO}_4(s)$	-1278.2	$\text{SnCl}_4(l)$	-545.2
$\text{MnO}(s)$	-384.9	$\text{SnO}(s)$	-286.2
$\text{MnO}_2(s)$	-519.7	$\text{SnO}_2(s)$	-580.7
$\text{NaCl}(s)$	-411.0	$\text{SO}_2(g)$	-296.1
$\text{NaF}(s)$	-569.0	$\text{SO}_3(g)$	-395.2
$\text{NaOH}(s)$	-426.7	$\text{ZnO}(s)$	-348.0
$\text{NH}_3(g)$	-46.2	$\text{ZnS}(s)$	-202.9

CHAPTER

18**Kinetics Worksheets****Chapter Outline**

- 18.1 RATES OF REACTIONS**
 - 18.2 RATE LAWS**
 - 18.3 REACTION MECHANISMS**
-

18.1 Rates of Reactions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Be able to express the rate of a chemical reaction.

- Speed is determined by
 - $\frac{\text{time}}{\text{distance}}$
 - $\frac{\text{distance}}{\text{time}}$
 - distance \times time
 - distance + time
- Chemical reaction rates are measured using
 - $\frac{\text{molarity}}{\text{minute}}$
 - $\frac{\text{moles}}{\text{second}}$
 - $\frac{\text{molals}}{\text{minute}}$
 - $\frac{\text{molarity}}{\text{second}}$
- True/False: [A] means the number of moles of A.
- True/False: The term Δ refers to a change in concentration.
- Define reaction rate
- For a given reaction, [A] is 0.45 M at 5 seconds and 0.23 M at 17 seconds. Calculate the rate of the reaction.

Lesson Objective: Describe the collision theory as it relates to chemical reactions.

- An ineffective collision is one in which
 - the wrong compound is formed
 - a smaller than expected amount of the desired compound is formed
 - no rearrangement of atoms occurs
 - a rearrangement of atoms occurs
- Reactants form products when all of the following occur except
 - collisions occur
 - collisions have random orientation
 - collisions have enough kinetic energy
 - collisions have correct orientation
- True/False: Bonds are more likely to break at higher vibrational energies
- True/False: Some reactions can occur at room temperature
- Define
 - collision theory

2. activation energy
12. Supplying reactant _____ with _____ causes the _____ between the atoms to _____ with a greater frequency.

Lesson Objective: Draw and analyze a potential energy diagram for a reaction, including heat of reaction, activation energy, and the activated complex.

13. One of the following statements about potential energy diagrams is not true
 1. the potential energy of products is greater than reactants in an endothermic reaction
 2. the potential energy of products is greater than reactants in an exothermic reaction
 3. the potential energy of reactants is greater than products in an exothermic reaction
 4. the ΔH of the system decreases for an exothermic reaction
14. The activated complex
 1. represents the completed reaction
 2. forms at the start of the reaction
 3. indicates the energy needed for the reaction to occur
 4. exists for several minutes
15. True/False: The structures of most activated complexes have been thoroughly studied.
16. True/False: Collisions between reactants always lead to products.
17. Define
 1. potential energy diagram
 2. activated complex
18. The activated complex is equally likely to either _____ the original _____ or go on to form the _____ - _____.
19. Draw a potential energy diagram with a reactant potential energy of 50 kJ/mol. An activated complex at 150 kJ/mol and a product at 100 kJ/mol. The reaction is
20. $X_2 + Y_2 \rightarrow X_2Y_2 \rightarrow 2XY$
21. Is this reaction exothermic or endothermic?
22. Draw a potential energy diagram for an exothermic reaction
23. $AB + C \rightarrow ABC \rightarrow A + BC$

Lesson Objective: Describe and explain various factors that influence the rates of reactions.

21. When a catalyst is involved in a reaction, it is
 1. written over the arrow
 2. written as a reactant
 3. written under the arrow
 4. written as a product
22. Temperature affects the rate of reaction by all of the following except
 1. increasing the collision rate
 2. decreasing the collision rate
 3. increasing the energy of collision
 4. increasing the percentage of effective collisions
23. Granular sugar will react more rapidly with sulfuric acid than a sugar cube will because of
 1. greater surface area
 2. higher pressure
 3. increased temperature

4. presence of granular catalyst
24. Catalysts operate by
1. forming different compounds
 2. providing different reaction pathways
 3. reacting with the starting materials
 4. increasing the activation energy required
25. True/False: A catalyst is not involved in the reaction it catalyzes.
26. True/False: A decrease in temperature results in a higher rate of collision.
27. True/False: An increase in gas pressure is equivalent to increasing the concentration of the gas.
28. True/False: Collision theory does not explain factors that influence rate of reaction.
29. Define catalyst
30. When more _____ are present in a given amount of _____, a _____ number of _____ will naturally occur between those particles.
31. _____ gas pressure leads to a greater _____ of collisions between _____ particles.
32. Explain how finding better catalysts can help a chemical manufacturing company become more profitable.
33. We can express the rate of this reaction as the change in concentration of A divided by the time.
34. $rate = -\frac{\Delta[A]}{\Delta t} = -\frac{[A]_{final} - [A]_{initial}}{\Delta t}$
35. A bracket around a symbol or formula means the concentration in molarity of that substance. The change in concentration of A is its final concentration minus its initial concentration. Because the concentration of A is decreasing over time, the negative sign is used. Thus, the rate for the reaction is positive, and the units are molarity per second or M/s. How would we write the rate equation if we were measuring change in product concentration?
36. Would it be easier to measure an increase in the concentration of a product or a decrease in the concentration of a reactant?

18.2 Rate Laws

Worksheet

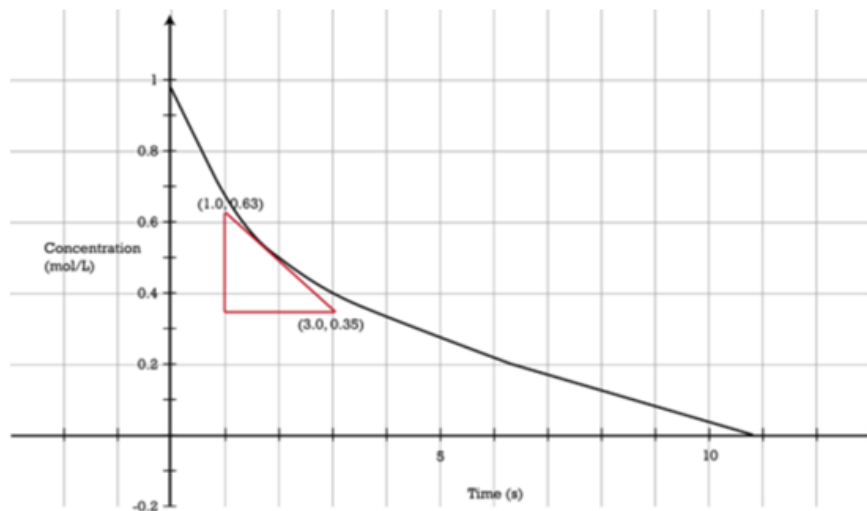
Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use a rate law to describe the concentration dependence of a reaction rate.

- The symbol \propto stands for
 - infinity
 - is less than
 - is proportional to
 - is greater than
- In a single-step reaction mechanism, all of the following are true except
 - as $[A]$ increases, rate increases
 - rate = $\frac{k}{[A]}$
 - as $[A]$ decreases, rate decreases
 - rate = $k[A]$
- In the reaction $A \rightarrow B$, the reaction is _____ - order.
 - second
 - third
 - zero
 - first
- In the reaction $A + B \rightarrow C$, if we reduce the amount of A by half,
 - the rate will be unchanged
 - the rate will decrease by half
 - the rate will be doubled
 - we cannot predict the rate with the available data
- In the reaction $A + B \rightarrow C$, the rate law is written as rate =
 - $k[A]$
 - $\frac{k[A]}{[B]}$
 - $k[A][B]$
 - $\frac{k[B]}{[A]}$
- The graphical plot of concentration versus time for a first-order reaction is a
 - curved line decreasing
 - a straight line decreasing
 - a curve line increasing
 - a straight line increasing
- True/False: The rate of a chemical reaction increases as the concentration of reactants increases.
- True/False: The rate of a reaction is independent of collision frequency.

9. True/False: The rate law for an equation can be determined by looking at the equation.
10. True/False: The rate constant value is affected by temperature.
11. True/False: In the reaction $A + B \rightarrow C$, if we double $[B]$, we do not affect the rate of the reaction.
12. True/False: In real reactions, we cannot assume a single-step mechanism.
13. Define the following terms:
 1. rate law
 2. specific rate constant
 3. first-order reaction
14. A large value for the rate constant means that the reaction is relatively _____, while a small value means that the reaction is relatively _____.
15. Determine the reaction rate at 4 seconds after initiation of the reaction.



16.

Lesson Objective: Determine the rate law for a reaction by analysis of experimental data.

16. In the reaction $A + B \rightarrow C + D$, we determine the rate law by measuring the rate under one of the following conditions:
 1. hold $[A]$ constant, measure rate at varying $[B]$
 2. hold both constant
 3. hold $[B]$ constant and measure rate at varying $[A]$
 4. hold each constant while varying the concentration of the other reactant.
17. Initial rates are determined because
 1. the slope of the curve is easier to measure at the beginning
 2. the rate slows down as the reaction proceeds
 3. the rate might increase as the reaction takes place
 4. it is more convenient
18. True/False: In the reaction between nitrogen monoxide and hydrogen, the reaction order for NO is two.
19. True/False: In the same reaction, the reaction order for hydrogen is one
20. Determine the reaction order and the rate law for each of the following reactions:
21. a.
22. $C + D \rightarrow E$

TABLE 18.1:

Exp #	[C]	[D]	Rate (M/sec)
1	0.1	0.01	0.02
2	0.1	0.02	0.04
3	0.1	0.03	0.06
4	0.1	0.04	0.08
5	0.2	0.04	0.08
6	0.3	0.04	0.08

b. $C + D \rightarrow E$

TABLE 18.2:

Exp #	[C]	[D]	Rate (M/sec)
1	0.1	0.01	0.02
2	0.1	0.02	0.08
3	0.1	0.03	0.18
4	0.1	0.04	0.32
5	0.2	0.04	1.28
6	0.3	0.04	2.88

Lesson Objective: Calculate the specific rate constant from experimental data.

- If the rate of a reaction does not vary as we change the reactant concentration, the reaction is _____ for that reactant.
 - first order
 - zero order
 - third order
 - second order
- In the $NO + H_2$ reaction, if the rate had quadrupled when we doubled the hydrogen concentration, the reaction would have been _____ for hydrogen.
 - first order
 - zero order
 - third order
 - second order
- True/False: In the $NO + H_2$ reaction, the rate tripled when the hydrogen gas concentration was doubled.
- True/False: The units for the specific rate constant vary with the order of the reaction.
- Calculate the specific rate constants for problem 20 reactions a and b above
- Since the rate of a reaction generally depends upon collision frequency, it stands to reason that as the concentration of A increases, the reaction rate increases. Likewise, as the concentration of A decreases, the reaction rate decreases. How does this observation related to collision theory?
- How does knowledge of a rate law help in the manufacture of a specific chemical?

18.3 Reaction Mechanisms

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe how the elementary steps of a chemical reaction combine to produce the overall reaction mechanism.

- The overall reaction specifies all of the following except
 - the starting materials
 - specific steps of the reaction
 - final products
 - catalysts
- True/False: Most reactions occur in a single step.
- Define the following terms:
 - elementary step
 - reaction mechanism
- What is the reaction intermediate in the following process?
- $Cl + O_3 \rightarrow ClO + O_2$
- $ClO + O \rightarrow Cl + O_2$
-
- net result:
- $O_3 + O \rightarrow 2O_2$

Lesson Objective: Be able to identify intermediates and catalysts when they appear in a reaction mechanism.

- In the reaction between nitrogen monoxide and oxygen, the reaction intermediate is
 - NO
 - NO_2
 - N_2O_2
 - O_2
- The reaction intermediate in the above reaction was
 - identified experimentally
 - added to balance the equation
 - included as a hypothesis
 - added as a reasonable guess
- True/False: A reaction intermediate is shown as part of the overall equation.
- True/False: The presence of a reaction intermediate allows us to understand better the entire reaction process.
- Define intermediate
- An intermediate is always formed in an _____ step of the mechanism and then _____ in a _____-
_ step.

Lesson Objective: Write a rate law for an elementary step, and identify the molecularity of the reaction.

11. Termolecular reaction steps are rare for all the following reasons except
 1. they require simultaneous collision of three molecules
 2. the ΔH for these systems is very high
 3. all molecules must be correctly oriented
 4. all three collisions must be of sufficient energy
12. True/False: A bimolecular reaction can only be second order.
13. Define molecularity
14. How does the presence of a reaction intermediate in the reaction between nitrogen monoxide and oxygen affect the rate of the reaction?
15. A proposed mechanism for the conversion of ozone is
16. (1)
17. $O_3 \rightarrow O_2 + O$
18. (2)
19. $O_3 + O \rightarrow 2O_2$
 1. Identify the molecularity of each step.
 2. Write a rate law for each step
20. Identify the molecularity of the following reactions:
 1. $2NO + O_2 \rightarrow 2NO_2$
 2. $N_2O_4 \rightarrow 2NO_2$
 3. $CO + NO_3 \rightarrow NO_2 + CO_2$
 4. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

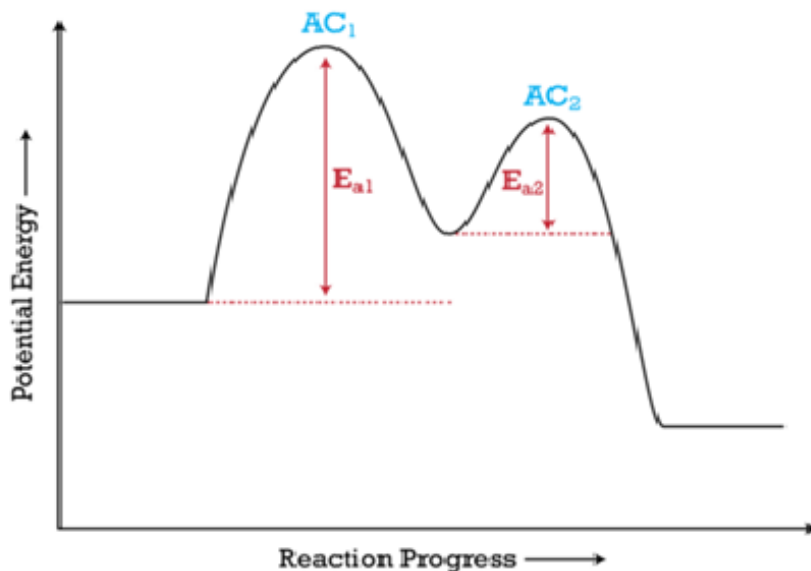
Lesson Objective: Identify the rate-determining step of a reaction.

17. In the decomposition of hydrogen peroxide, iodide ion functions as a
 1. catalyst
 2. reactant
 3. product
 4. intermediate
18. True/False: The decomposition of hydrogen peroxide is first-order with regard to iodide
19. Define rate-determining step
20. For the reaction $H_2 + 2ICl \rightarrow I_2 + 2HCl$ (all in the gas phase), the experimentally determined rate law is $rate = k[H_2][ICl]$. A proposed two-step mechanism is
21. step 1:
22. $H_2 + ICl \rightarrow HI + HCl$ (*slow*)
23. step 2:
24. $HI + ICl \rightarrow I_2 + HCl$ (*fast*)
25. Is this a reasonable mechanism? Explain your answer. Why is step one considered to be the rate-determining step?

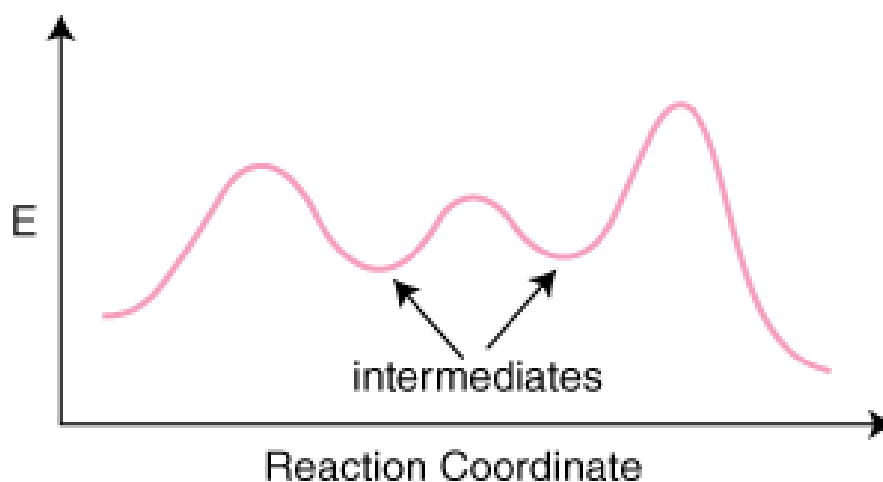
Lesson Objective: Analyze the potential energy diagram for a multi-step reaction.

21. All of the following statements about reaction intermediates are true except
 1. they can be detected
 2. they can be characterized
 3. they exist between activated complexes
 4. they are not included in rate law discussions

22. True/False: The overall enthalpy change is affected by the activated complexes.
 23. Indicate the actual ΔH for the reaction below.



24.
 25. How many steps would you expect to be in any valid reaction mechanism for this potential energy diagram? Explain your answer.



26.
 27. IO^- is present as an intermediate in the reaction. The iodide ion catalyst also appears in the mechanism. It is a reactant in the first elementary step and a product in the second step. By definition, catalysts are not used up in a reaction, so if they are consumed by one step in a mechanism, they must be regenerated by a subsequent step.
 28. How does this information affect our definition of a catalyst?
 29. If a reaction mechanism is proposed that is based on an intermediate that has not been detected, is this a useful mechanism? Explain your answer.

CHAPTER

19**Equilibrium Worksheets****Chapter Outline**

19.1 THE NATURE OF EQUILIBRIUM

19.2 LE CHÂTELIER'S PRINCIPLE

19.3 SOLUBILITY EQUILIBRIUM

19.1 The Nature of Equilibrium

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the nature of a reversible reaction.

- In a reversible reaction
 - all the reactants convert to products
 - all the products convert to reactants
 - the reaction proceeds in both directions simultaneously
 - the reaction goes in one direction and then reverses in the other direction
- Reversible reactions are indicated by
 - $\rightarrow \leftarrow$
 - $\leftarrow \rightarrow$
 - \rightleftharpoons
 - \rightleftharpoons
- True/False: All reactions are reversible.
- True/False: Many chemical reactions do not proceed to completion.
- Define reversible reaction.

Lesson Objective: Define chemical equilibrium.

- The forward reaction rate _____ as the reverse reaction begins
 - increases
 - decreases
 - remains unchanged
 - varies with the rate of the reverse reaction
- Chemical equilibrium can be attained under all but one of the following conditions
 - beginning only with reactants
 - beginning only with products
 - beginning with a mix of reactants and products
 - removing products as the reaction proceeds
- All of the following are equilibrium situations except
 - phase equilibrium
 - dynamic equilibrium
 - solution equilibrium
 - reaction equilibrium
- True/False: Equilibrium is reached when all concentrations have become constant.
- True/False: At equilibrium, the concentrations of all components are equal.

11. True/False: The equilibrium position depends on the initial concentrations of materials.
12. Define chemical equilibrium.
13. The equilibrium position is a property of the particular _____ reaction and does not depend upon the _____ concentrations of the _____ and _____.
14. List the conditions and properties of a system at equilibrium.

Lesson Objective: Write chemical equilibrium expressions from balanced chemical equations.

15. The concentrations of each substance in an equilibrium expression is measured in
 1. moles/L
 2. moles/sec
 3. molal
 4. L/mol
16. The value of the equilibrium constant is affected by
 1. volume
 2. reactant concentration
 3. temperature
 4. product concentration
17. The equilibrium constant symbol is
 1. k_{eq}
 2. K_{eq}
 3. K_{const}
 4. k_{const}
18. For the reaction: $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$, the equilibrium constant would be
 1. $\frac{[CaO][CO_2]}{[CaCO_3]}$
 2. $\frac{[CaCO_3]}{[CaO][CO_2]}$
 3. $[CO_2]$
 4. $\frac{[CaO]}{[CaCO_3]}$
19. True/False: A large equilibrium constant means that products are favored over reactants.
20. True/False: The units for the equilibrium constant vary depending on the reaction.
21. True/False: Balanced equations are needed before writing equilibrium constants.
22. At equilibrium, the concentrations of the substances are constant.
23. Define equilibrium constant.
24. Write equilibrium constants for the following reactions:
 1. $CO(g) + 2H_2(g) \rightleftharpoons CH_3OH(g)$
 2. $CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$
 3. $2NaHCO_3(s) \rightleftharpoons Na_2CO_3(s) + H_2O(g) + CO_2(g)$
 4. $CaH_2(s) + 2H_2O(g) \rightleftharpoons Ca(OH)_2(s) + 2H_2(g)$

Lesson Objective: Calculate the equilibrium constant for a reaction and perform calculations involving that constant

25. a. At the equilibrium point in the decomposition of phosphorus pentachloride to chlorine and phosphorus trichloride, the following concentrations are obtained: 0.010 mol/L PCl_5 , 0.15 mol/L PCl_3 and 0.37 mol/L Cl_2 . Determine the K_{eq} for the reaction.
 - b.



- c. At a certain temperature, the equilibrium mixture of PCl_5 , PCl_3 and Cl_2 has the following concentrations:
 d. $[\text{PCl}_3] = 0.035 \text{ M}$, $[\text{PCl}_5] = 0.017 \text{ M}$, $[\text{Cl}_2] = 0.074 \text{ M}$
 e. Calculate
 f. K_{eq}
 g. for the reaction
 h. $\text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g}) \rightleftharpoons \text{PCl}_5(\text{g})$
26. Given the following reaction and the concentrations of 0.5 M for H_2S and 0.2 M for HI, what is the value for K_{eq} ?
27. $\text{H}_2\text{S}(\text{g}) + \text{I}_2(\text{s}) \rightleftharpoons 2\text{HI}(\text{g}) + \text{S}(\text{s})$
28. Calculate the equilibrium constant (K_{eq}) for the following reaction:
29. $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$
30. when the equilibrium concentrations at
31. 25°C
32. were found to be:
- 33.

$$[\text{H}_2] = 0.0505 \text{ M}$$

I_2

$$= 0.0498 \text{ M}$$

HI

$$= 0.389 \text{ M}$$

34. The equilibrium expression only shows those substances whose concentrations are variable during the reaction. A pure solid or a pure liquid does not have a concentration that will vary during a reaction. Therefore, an equilibrium expression omits pure solids and liquids and only shows the concentrations of gases and aqueous solutions.
35. Would you expect an equilibrium constant to have a different value if we were able to incorporate information about pure solids or liquids?
36. Suggest one way we could incorporate information about changing amounts of solids during an equilibrium reaction.

19.2 Le Châtelier's Principle

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Explain the factors that stress a system at equilibrium and use Le Châtelier's principle to predict how the system will respond to each stress.

- A chemical system in equilibrium can be disrupted by changes in all of the following except
 - temperature
 - time
 - concentration
 - pressure
- Chemical equilibrium was studied by _____.
 - the French chemist Pierre Le Châtelier
 - the Belgian chemist Henri Le Châtelier
 - the English chemist Henri Le Châtelier
 - the French chemist Henri Le Châtelier
- Addition of a catalyst to a reaction in equilibrium.
 - increases the rate of the forward reaction
 - decreases the rate of the forward reaction
 - has no effect on the equilibrium
 - increases the rate of the reverse reaction
- In a gas-phase reaction $2A + 3B \rightleftharpoons 4C + D$, an increase in pressure will.
 - favor the forward reaction
 - favor the reverse reaction
 - have no effect on the equilibrium
 - decrease the rates of both reactions
- In the equilibrium system $A \rightleftharpoons B$, an increase in $[A]$ will.
 - cause an increase in the forward reaction
 - cause an increase in the reverse reaction
 - cause a decrease in the forward reaction
 - have no effect on concentrations
- In the equilibrium reaction $CaCO_3(s) \rightleftharpoons CaO(s) + O_2(g)$, a decrease in pressure produces.
 - an increase in $CaCO_3$ formation
 - an increase in the rate of the forward reaction
 - a decrease in CaO formation
 - an increase in the rate of the reverse reaction
- In the equilibrium reaction $A + B \rightleftharpoons C + D + 87.5 \text{ kJ}$, 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 C + D
 4. no effect on K_{eq}
8. True/False: After an equilibrium has been disturbed, the equilibrium position can be restored to its original position.
 9. True/False: A change in pressure will always affect equilibrium.
 10. True/False: If more N_2 is added to the Haber-Bosch process, the ammonia concentration will increase.
 11. True/False: Continued removal of ammonia in the Haber-Bosch process will force the reaction to completion.
 12. True/False: The forward Haber-Bosch reaction is endothermic.
 13. True/False: A change in pressure on a liquid or solid does not stress the system.
 14. True/False: An increase in pressure causes gas molecules to be forced closer together.
 15. State Le Châtelier's Principle
 16. When the pressure is increased by _____ the available volume, the reaction that produces _____ total moles of gas becomes _____.
 17. Catalysts have equal effects on the _____ and _____ rates, so for a system at equilibrium, these two rates remain _____.
 18. Predict how each of the following changes will move the direction of the reaction for the reaction $CO_2 + C + \text{heat} \rightleftharpoons 2CO$
 1. increase temperature
 2. increase CO concentration
 3. decrease pressure

Lesson Objective: Explain how a reaction is able to go to completion.

19. A reaction will go to completion under all of the following situations except
 1. formation of a precipitate
 2. formation of an ionized compound
 3. formation of a gas
 4. formation of a nonionized compound
20. A reaction goes to completion when all of the following occur except
 1. a product is removed as soon as it is formed
 2. the reverse reaction slows down
 3. the reverse reaction cannot establish itself
 4. equilibrium does not occur
21. _____ The formation of liquid water causes a reaction to go to completion.
22. _____ Formation of silver chloride is a reversible process.

Lesson Objective: Analyze how a change in the concentration of a reaction component can change the equilibrium position but not the value of the equilibrium constant.

23. In the equilibrium system $A \rightleftharpoons B$, an increase in [A] 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}
24. True/False: A shift in equilibrium position is not the same as a change in K_{eq} .
25. Given the equation:
26. $H_2 + I_2 \rightleftharpoons 2HI$

27. Calculate all three equilibrium concentrations when the initial concentrations of
28. H_2
29. and
30. $I_2 = 0.200 M$
31. each and
32. $K_c = 64.0$
33. .
34. Given the equation:
35. $PCl_3 + Cl_2 \rightleftharpoons PCl_5$
36. Calculate all three equilibrium concentrations when
37. $K_c = 16.0$
38. and the initial concentration of
39. $PCl_5 = 1.00 M$
40. .
41. Given the equation:
42. $COCl_2 \rightleftharpoons CO + Cl_2$
43. Calculate all three equilibrium concentrations when
44. $K_c = 0.680$
45. with
46. $[CO]_{init} = 0.500$
47. and
48. $[Cl_2]_{init} = 1.00 M$
49. .
50. Equilibrium represents a balance between the reactants and the products of a chemical reaction. Changes to the conditions of the system can disturb that equilibrium. When this occurs, the system reacts in such a way as to restore the equilibrium. However, the position of equilibrium will be changed following the disturbance. In other words, the response of the system involves a change in the amounts of the reactants and products. Some will increase and some will decrease until equilibrium is reestablished.
51. In the Haber-Bosch process for the synthesis of ammonia, the reaction takes place in a closed system. If additional nitrogen is added, that produces an increase in the concentration of material, which is expected to favor the forward reaction to form ammonia. How will this increase in nitrogen gas affect the pressure of the system and the equilibrium position?
52. In the reaction $CaCO_3(s) \rightleftharpoons CaO(s) + O_2(g)$, if the amount of O_2 is increased, does that decrease the decomposition of calcium carbonate or decrease the formation of calcium oxide?

TABLE 19.1:

	$[H_2]$	$[I_2]$	$[HI]$
Initial	0.200	0.200	0
Change	-x	-x	+2x
Equilibrium	0.200 - x	0.200 - x	2x

TABLE 19.2:

	$[PCl_3]$	$[Cl_2]$	$[PCl_5]$
Initial	0	0	1.00
Change	+ x	+ x	-x
Equilibrium	x	x	1.00 - x

TABLE 19.3:

	$[COCl_2]$	$[CO]$	$[Cl_2]$
Initial	0	0.500	1.00
Change	+ x	- x	- x
Equilibrium	x	0.500 - x	1.00 - x

19.3 Solubility Equilibrium

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives.

Lesson Objective: Write solubility product constant expressions for nearly insoluble ionic compounds.

- _____ salts tend to be very soluble
 - hydroxides
 - alkali metals
 - transition metals
 - halides
- The solubility product constant for $\text{Fe}(\text{OH})_2$ is written
 - $[\text{Fe}][\text{OH}]$
 - $\frac{[\text{Fe}]}{[\text{OH}]}$
 - $[\text{Fe}][\text{OH}]^2$
 - $\frac{[\text{Fe}]}{[\text{OH}]^2}$
- AgCl can still be considered a
 - weak electrolyte
 - strong electrolyte
 - nonelectrolyte
 - covalent compound
- True/False: NaCl is highly soluble in water.
- True/False: PbCl_2 is less soluble in water than PbCO_3 .
- True/False: The solid portion of an equilibrium equation appears in the K_{sp} expression.
- Define solubility product constant.
- These “mostly insoluble” compounds are still considered to be _____ electrolytes because, essentially, any portion of the compound that _____ will also dissociate into _____.
- Write the equilibrium expression and the K_{sp} expressions for the following compounds:
 - CaCO_3
 - $\text{Mg}(\text{OH})_2$
 - SrF_2
 - ZnCr_2O_7

Lesson Objective: Calculate the value of K_{sp} for a compound from its solubility and determine the solubility of a compound with a known K_{sp} .

- Solubility is normally expressed as
 - grams/mL
 - grams/L

3. mg/mL
 4. Kg/L
11. K_{sp} calculations uses concentration units of
1. moles/L
 2. moles/mL
 3. grams/L
 4. moles \times L
12. True/False: Molar mass is used to convert between solubility and molar solubility.
13. True/False: Solubility of a material can be used to calculate K_{sp} .
14. Define molar solubility.
15. The molar solubility of a given compound represents the _____ molarity _____ that is possible for that compound.
16. The aqueous solubility for PbI_2 is 5.4 grams/L at 25°C. Calculate the K_{sp} of PbI_2 at 25°C.
17. The K_{sp} at 25°C is 8.5×10^{-17} for AgI. What is the molar solubility?
18. Calculate the K_{sp} for Bi_2S_3 when the solubility of the compound is 1.7×10^{-14} g/L.
19. Calculate the molar solubility of $BaCrO_4$ given the K_{sp} of 1.17×10^{-10} .
20. Determine the molar concentration of the Ba^{2+} ion in a solution of $BaSO_3$ given a K_{sp} of 7.94×10^{-7} .

Lesson Objective: Use the K_{sp} of a compound to predict whether a precipitate will form when two solutions are mixed together.

21. When two solution are mixed that could form a precipitate, equilibrium is attained 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 less than the ion product
22. The _____ and _____ of each solution that is mixed together must be used to calculate the values of ions that might potentially form a precipitate.
23. Does a ppt. of AgCl form when 1 mL of 0.1 mol/L $AgNO_3$ is added to a beaker containing 1 L of tap water with a Cl^- ion concentration of 1.0×10^{-5} mol/L?
24. 50 mL of 0.0010 M NaCl is mixed with 50 mL of 0.00020 M $AgNO_3$. Will a precipitate form?
25. Equal volumes of 0.025 M potassium bromide (KBr) and 0.005 M lead(II) nitrate ($Pb(NO_3)_2$) solutions were mixed. The $PbBr_2 K_{sp} = 7.9 \times 10^{-5}$. Will a precipitate form?

Lesson Objective: Describe the common ion effect and its relationship to solubility.

26. True/False: The common ion effect is an application of LeChâtelier's principle.
27. Define the following terms:
1. common ion
 2. common ion effect
28. The K_{sp} for a saturated solution of $PbCl_2$ is 1.7×10^{-5} . What is the $[Pb^{2+}]$? Enough NaCl is added to this solution to produce a $[Cl^-]$ of 0.1 M with essentially no change in volume. What will the new $[Pb^{2+}]$ be?
29. As a result, the ion product $[Ca^{2+}][SO_4^{2-}]$ would increase to a value that is greater than the K_{sp} . According to Le Châtelier's principle, the equilibrium above would shift to the left in order to relieve the stress of the added calcium ion. Additional calcium sulfate would precipitate out of the solution until the ion product, once again, becomes equal to K_{sp} .
30. Why would added calcium ion be considered a stress on the system?
31. Addition of aluminum sulfate is often used in water purification systems. This salt is fairly soluble, but the presence of aluminum ions has been associated with a number of health problems. How could the aluminum ion content be lowered in order to decrease the levels of this material in water?

CHAPTER

20**Entropy and Free Energy
Worksheets****Chapter Outline**

20.1 ENTROPY

20.2 SPONTANEOUS REACTIONS AND FREE ENERGY

20.3 FREE ENERGY AND EQUILIBRIUM

20.1 Entropy

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Identify the two driving forces behind all chemical reactions and physical processes.

- One of the following is not a characteristic of exothermic reactions
 - reactants have high energy compared to products
 - products are more stable than reactants
 - energy is released into the surroundings
 - products have lower stability than reactants
- One of the following is not a characteristic of an endothermic reaction
 - energy is absorbed during the reaction
 - reaction is energetically favorable
 - products have larger quantity of energy than reactants
 - products are less stable than reactants
- True/False: Most naturally occurring reactions are exothermic.
- True/False: The energy of the system increases during an exothermic reaction.
- Low energy can be thought of as providing a _____ degree of _____ to a chemical system.

Lesson Objective: Describe entropy, and be able to predict whether the entropy change for a reaction is increasing or decreasing.

- One of the following does not show an increase in entropy
 - ice becomes liquid water
 - solid reactant forms liquid product
 - CO₂ condenses to form dry ice
 - iodine crystals form a vapor
- A decrease in entropy can be seen when
 - a sugar cube is crushed
 - sugar is crystallized out of solution
 - sugar is dissolved in water
 - sugar cubes are spilled on the table
- One of the following does not illustrate an increase in entropy
 - increase in temperature
 - decrease in temperature
 - increase in kinetic energy
 - increase in motion of particles
- A decrease in entropy is seen when

- NaCl is dissolved in water
 - CaCO_3 forms CaO and CO_2
 - Hydrogen and oxygen form water
 - Water evaporates
- True/False: The solid state of water is highly ordered.
 - True/False: The entropy of a gas is less than the entropy of the corresponding solid.
 - True/False: Chemical reactions tend to proceed in a way as to decrease the entropy of the system.
 - True/False: Work must be done to decrease entropy.
 - Define entropy
 - The change from the solid to liquid state of any substance corresponds to an _____ in the _____ of the system.
 - In the following reactions, indicate whether entropy is increasing or decreasing:
 - $2\text{Al}(s) + 3\text{Br}_2(g) \rightarrow 2\text{AlBr}_3(s)$
 - $2\text{CO}(g) + \text{O}_2(g) \rightarrow 2\text{CO}_2(g)$
 - $2\text{HgO}(s) \rightarrow 2\text{Hg}(l) + \text{O}_2(g)$
 - $2\text{Cl}_2\text{O}_5(g) \rightarrow 2\text{Cl}_2(g) + 5\text{O}_2(g)$
 - $\text{Ca}(s) + 2\text{H}_2\text{O}(l) \rightarrow \text{Ca}(\text{OH})_2(aq) + \text{H}_2(g)$
 - $3\text{Li}_2\text{CO}_3(aq) + 2\text{Al}(\text{NO}_3)_3(aq) \rightarrow 6\text{LiNO}_3(aq) + \text{Al}_2(\text{CO}_3)_3(s)$

Lesson Objective: Calculate the standard entropy change for a reaction from the standard entropies of all substances in the reaction.

- The symbol for entropy is
 - E
 - S
 - H
 - K
- The units for entropy are
 - $\text{J}/\text{K} \cdot \text{mol}$
 - $\text{K}/\text{J} \cdot \text{mol}$
 - J/mol
 - $\text{J} \cdot \text{mol}/\text{K}$
- Standard entropy change can be calculated using
 - $\Delta S^\circ = \sum S^\circ(\text{products}) - \sum S^\circ(\text{reactants})$
 - $\Delta S^\circ = \sum nS^\circ(\text{products}) - \sum nS^\circ(\text{reactants})$
 - $\Delta S^\circ = \sum nS^\circ(\text{reactants}) - \sum nS^\circ(\text{products})$
 - $\Delta S^\circ = \sum S^\circ(\text{reactants}) - \sum S^\circ(\text{products})$
- True/False: All molecular motion ceases at 0 degrees K.
- True/False: The standard entropy of a substance is given by S° .
- True/False: The entropy change for the vaporization of water is negative.
- As the temperature of the substance _____, its entropy increases because of an increase in molecular _____.
- At what temperature are standard entropy values determined?
- Is the standard entropy for the vapor state of water higher or lower than the liquid state? Explain your answer.
- Calculate ΔS° for the following reactions:
 - $\text{CO}(g) + 2\text{H}_2(g) \rightarrow \text{CH}_3\text{OH}(l)$
 - S° values: $\text{CO} = 197.9 \text{ J}/\text{K} \cdot \text{mol}$ $\text{H}_2 = 131.0 \text{ J}/\text{K} \cdot \text{mol}$ $\text{CH}_3\text{OH} = \frac{126}{8} \text{ J}/\text{K} \cdot \text{mol}$
 - $\text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g)$

d. S° values: $\text{CaCO}_3 = 92.9 \text{ J/K} \cdot \text{mol}$ $\text{CaO} = 38.2 \text{ J/K} \cdot \text{mol}$ $\text{CO}_2 = 213.6 \text{ J/K} \cdot \text{mol}$

27. All molecular motion ceases at absolute zero (0 K). Therefore, the entropy of a pure crystalline substance at absolute zero is defined to be equal to zero.
28. How does this definition involve motion?
29. No molecular motion is one definition of absolute zero and of zero entropy. What assumption is being made about the structure of the crystal?

20.2 Spontaneous Reactions and Free Energy

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the meaning of a spontaneous reaction in terms of enthalpy and entropy changes.

- All of the following are characteristics of spontaneous reactions except
 - decrease in enthalpy
 - increase in entropy
 - decrease in randomness
 - release of energy
- A combustion reaction is considered to be spontaneous because
 - entropy is lowered
 - the energy of the system increases
 - reaction is exothermic
 - randomness of the system decreases
- All of the following are characteristics of non-spontaneous reactions except
 - entropy increases
 - reaction is endothermic
 - reactants are favored over products
 - randomness decreases
- The release of CO_2 when carbonic acid is dissolved in water is spontaneous because
 - the process is endothermic
 - the equilibrium favors the forward reaction
 - non-spontaneous reactions do not occur
 - the entropy of the system is increased
- Situations associated with NO formation involve all but one of the following
 - high temperatures
 - lightning strikes
 - negative entropy change
 - highly endothermic
- True/False: The formation of NO from nitrogen and oxygen in the atmosphere is non-spontaneous at normal temperatures and pressures.
- True/False: A spontaneous reaction is one that occurs rapidly.
- True/False: Changes in enthalpy and entropy are the driving forces behind all chemical reactions.
- True/False: Entropy decreases during most combustion reactions.
- True/False: Formation of carbonic acid after CO_2 is added to water is a non-spontaneous reaction
- Define the following terms
 - spontaneous reaction

2. non spontaneous reaction
12. Write the expression for Gibbs free energy.
13. A spontaneous reaction is one in which product _____ favored, even if the reaction is extremely _____ .
14. Just because a reaction is not _____ does not mean that it does not occur at all. Rather, it means that the _____ will be favored over the _____ at equilibrium, even though some products may indeed form.
15. Why would a decrease in entropy be required for a non-spontaneous reaction to occur?

Lesson Objective: Define free energy and calculate the change in free energy for a reaction using known values of the changes in enthalpy and entropy.

16. Units for ΔG are
 1. $J/K \cdot mol$
 2. J/mol
 3. kJ/mol
 4. $kJ/K \cdot mol$
17. If ΔH is positive and ΔS is negative, then ΔG is
 1. always negative
 2. never negative
 3. negative at higher temperatures
 4. negative at lower temperatures
18. 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
19. Free energy is energy that is
 1. stored for future use
 2. lost as a product of the reaction
 3. available to do work
 4. used to drive the reaction
20. True/False: Spontaneous reactions release free energy as they proceed.
21. True/False: The symbol for free energy is in honor of the American scientist James Gibbs.
22. True/False: ΔG must be negative for a spontaneous reaction.
23. True/False: If ΔH is negative, the reaction is endothermic.
24. When ΔH is _____ and ΔS is positive, the sign of ΔG will always be _____, and the reaction will be _____ at all temperatures.
25. Why is a reaction always spontaneous (negative ΔG) when the reaction is exothermic and entropy increases?
26. Using the values provided below, calculate ΔG and predict whether or not the reaction will be spontaneous.

TABLE 20.1:

	ΔH (kJ/mol)	ΔS (J/K · mol)	T (K)
a	-315	-405	320
b	+95	- 120	285
c	-101	+15.7	279
d	+115	+183	525

Lesson Objective: Determine the spontaneity of a reaction based on the value of its change in free energy at various temperatures.

27. When the $T\Delta S$ term becomes larger than the ΔH term, the ΔG value will be
 1. negative if ΔS is negative
 2. negative if ΔS is positive
 3. positive if ΔS is positive
 4. variable value
28. True/False: When a reaction is exothermic, the ΔH term needs to be larger than the $T\Delta S$ term for the reaction to be spontaneous.
29. $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
30. $\Delta H^\circ = -92.22 \text{ kJ}$
31. $\Delta S^\circ = -198.75 \text{ J/K}$
32. Is the above reaction spontaneous at 298 K?
33. Assume that the
34. ΔH
35. and
36. ΔS
37. values are still valid at higher temperatures, will the reaction still be spontaneous at 775 K?
38. When ΔH is negative and ΔS is positive, the sign of ΔG will always be negative, and the reaction will be spontaneous at all temperatures. This corresponds to both driving forces being in favor of product formation. When ΔH is positive and ΔS is negative, the sign of ΔG will always be positive, and the reaction can never be spontaneous. This corresponds to both driving forces working against product formation.
39. What is required for a non-spontaneous reaction to occur?
40. A combustion reaction is considered to be spontaneous, but firewood does not burst into flame until an external source of heat is provided. Why is burning wood considered an spontaneous reaction?

20.3 Free Energy and Equilibrium

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Determine the temperature at which a reversible reaction will achieve equilibrium by using the Gibbs free energy equation.

- Assumptions made in calculations dealing with quicklime formation include all of the following except
 - ΔH and ΔS values are independent of temperature
 - no products are formed below 835°C
 - ΔG switches from positive to negative at 835°C
 - The CO_2 formed remains in the reaction vessel
- To determine the temperature at which ΔG changes sign, we calculate
 - the equation for $\Delta H = \text{zero}$
 - the equation using different T values
 - the equation for $\Delta S = \text{zero}$
 - the equation for $\Delta H = \text{zero}$
- In the quicklime manufacturing process
 - no CO_2 is produced below 835°C
 - measurable CO_2 is detected at 300°C
 - measurable CO_2 is detected above 700°C
 - The CO_2 pressure begins to exceed 1 atm at 700°C
- At the temperature at which a change of state occurs, all of the following are true except
 - $\Delta G = \text{zero}$
 - the two states are at equilibrium
 - $\Delta S_{\text{fus}} = \text{zero}$
 - T = temperature at which transition occurs
- True/False: The ΔH for the quicklime reaction is positive at 25°C .
- True/False: The entropy for the quicklime reaction decreases as the reaction proceeds.
- True/False: The entropy change during the melting of ice can be calculated using ΔH_{vap} .
- True/False: The ΔG value = zero at the equilibrium point in the ice-water system.
- For a given reaction $\Delta H^{\circ} = 35.6 \text{ kJ/mol}$ and $\Delta S^{\circ} = 113 \text{ J/K} \cdot \text{mol}$. Is the reaction spontaneous at 280 K? At what temperature will the reaction be spontaneous?
- The melting point of Hg is 234 K and the $\Delta H_{\text{fus}} = 2.30 \text{ kJ/mol}$. What is the ΔS_{fus} ?
- Metallic sodium boils at 1158 K and the ΔH_{vap} is 98 kJ/mol. What is the ΔS_{vap} ?
- For a reaction with a ΔH of 14.2 kJ/mol and a ΔS of $46.8 \text{ J/K} \cdot \text{mol}$, at what temperature will the reaction be spontaneous?

Lesson Objective: Describe the relationship between standard free energy change (ΔG°) and the equilibrium constant (K_{eq}) for reversible reactions.

13. The relationships between Gibbs free energy and the equilibrium constant is written
1. $\Delta G^\circ = RT \ln(K_{eq})$
 2. $\Delta G^\circ = -RT \ln(K_{eq})$
 3. $\Delta G^\circ = -RT \log(K_{eq})$
 4. $\Delta G^\circ = -RT \log(K_{eq})$
14. The variable R is the
1. ideal constant
 2. equilibrium ratio
 3. ideal gas constant
 4. gas constant
15. True/False: When K_{eq} is large, ΔG should be negative.
16. True/False: When $K_{eq} = 1$, then ΔG is positive.
17. Fill in the following table:

TABLE 20.2:

K_{eq}	$\ln(K_{eq})$	ΔG°	Description
>1		negative	
1	0		Reactants and products are equally favored.
<1	negative	positive	

Lesson Objective: Convert between K_{eq} and ΔG° for a reaction at a given temperature.

18. For the equilibrium reaction $A + B \rightleftharpoons AB$ at 298 K
19. $\Delta H^\circ = -55.8 \text{ kJ/mol}$
20. $\Delta S^\circ = +80.8 \text{ J/K} \cdot \text{mol}$
21. Calculate
22. ΔG
23. and
24. K_{eq}
25. for this process
26. The K_{eq} for the reaction $2SO_3(g) \rightleftharpoons 2SO_2(g) + O_2(g)$ is 1.4×10^{-25} . Calculate the ΔG° for this reaction at 298 K.
27. $AgCl(s) + H_2O \rightleftharpoons Ag^+(aq) + Cl^-(aq)$ $\Delta G = 55.6 \text{ kJ}$.
28. Calculate the
29. K_{eq}
30. for the reaction at 298 K
31. For the reaction $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$, the ΔG at 298 K = -32.96 kJ
32. Calculate the
33. K_{eq}
34. .
35. When this reaction is performed, the amount of products can be detected by monitoring the pressure of the CO_2 gas that is produced. Above about 700°C , measurable amounts of CO_2 are produced. The pressure of CO_2 at equilibrium gradually increases with increasing temperature. Above 835°C , the pressure of CO_2 at equilibrium begins to exceed 1 atm, the standard-state pressure. This is an indication that the products of the reaction are now favored above that temperature.
36. Why is this calculation helpful in the production of quicklime?
37. What other information might be useful in designing a system for production of quicklime from calcium carbonate?

TABLE 20.3:

K_{eq}	$\ln(K_{eq})$	ΔG°	Description
>1	positive	negative	Products are favored at equilibrium.
1	0	0	Reactants and products are equally favored.
<1	negative	positive	Reactants are favored at equilibrium.

CHAPTER 21**Acids and Bases
Worksheets****Chapter Outline**

- 21.1 ACID-BASE DEFINITIONS**
 - 21.2 THE PH CONCEPT**
 - 21.3 ACID AND BASE STRENGTH**
 - 21.4 ACID-BASE NEUTRALIZATION**
 - 21.5 SALT SOLUTIONS**
-

21.1 Acid-Base Definitions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the properties of acids and bases.

- All of the following are acids except
 - citrus fruits
 - carbonated soda
 - drain cleaner
 - vinegar
- All of the following are properties of acids except
 - sour taste
 - aqueous solutions are electrolytes
 - turn litmus blue
 - react with some metals to produce H_2
- All of the following are properties of bases except
 - turn litmus red
 - bitter taste
 - do not react with metals
 - aqueous solutions are electrolytes
- True/False: Oranges contain ascorbic acid
- True/False: Sodium hydrogen carbonate is an antacid.
- True/False: All acids are strong electrolytes.
- Fill in the table below

TABLE 21.1:

Property	Acids	Bases
color with phenolphthalein		
color with litmus		
taste		
reactivity with metals		

- Fill in the table below by naming the acid or base contained in the material

TABLE 21.2:

Material	Acid or Base Present in Material
drain cleaner	
carbonated soda	
vinegar	
antacid	
cleaners	
orange	

Lesson Objective: Define an acid and a base according to the Arrhenius theory.

9. An Arrhenius acid is a compound with
 1. reactive hydrogen atoms
 2. ionizable hydrogen atoms
 3. covalently bound hydrogen atoms
 4. hydrogen atoms attached to carbon
10. An Arrhenius base is a compound with
 1. ionizable OH group
 2. covalently bound OH group
 3. reactive OH group
 4. OH group attached to carbon
11. True/False: $\text{CH}_3\text{CH}_2\text{OH}$ is an Arrhenius base.
12. True/False: All Arrhenius bases are completely ionizable.
13. Define the following terms:
 1. monoprotic acid
 2. polyprotic acid
14. Fill in the blank spaces in the table below:

TABLE 21.3:

Compound	Name
CH_3COOH	
$\text{Mg}(\text{OH})_2$	
	hypochlorous acid
	phosphoric acid
	calcium hydroxide
KOH	
	sulfuric acid

Lesson Objective: Define an acid and a base according to the Brønsted-Lowry theory. Be able to identify the conjugate acid-base pairs in a Brønsted-Lowry acid-base reaction.

15. A hydrogen ion is often referred to as a
 1. positron
 2. H atom
 3. proton
 4. portion

16. All of the following are true about ammonia except
1. contains hydroxide group
 2. reacts with acid
 3. turns litmus blue
 4. turns phenolphthalein pink
17. A Brønsted-Lowry acid-base reaction involves
1. formation of hydroxide ion
 2. transfer of proton
 3. utilization of water
 4. formation of free hydrogen ion
18. True/False: A hydrogen atom is also called a proton.
19. True/False: Water can act as a Brønsted-Lowry acid.
20. True/False: The hydroxide ion is a Brønsted-Lowry acid.
21. Define the following terms:
1. Brønsted-Lowry acid
 2. Brønsted-Lowry base
 3. amphoteric
 4. conjugate acid
 5. conjugate base
22. Identify the acid and base and the conjugate acid and base in each of the following reactions:
1. $\text{HClO}_4(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{ClO}_4^-(\text{aq})$
 2. $\text{HF}(\text{aq}) + \text{HSO}_3^-(\text{aq}) \rightleftharpoons \text{F}^-(\text{aq}) + \text{H}_2\text{SO}_3(\text{aq})$
 3. $\text{HNO}_2(\text{aq}) + \text{HS}^-(\text{aq}) \rightleftharpoons \text{NO}_2^-(\text{aq}) + \text{H}_2\text{S}(\text{aq})$

Lesson Objective: Define an acid and a base according to the Lewis theory.

23. A Lewis acid
1. donates a pair of electrons to form a covalent bond
 2. releases a proton
 3. accepts a proton
 4. accepts a pair of electrons to form a covalent bond
24. A Lewis base
1. donates a pair of electrons to form a covalent bond
 2. releases a proton
 3. accepts a proton
 4. accepts a pair of electrons to form a covalent bond
25. True/False: Ammonia can serve as a Lewis base.
26. True/False: A Lewis base forms an ionized compound
27. Identify the acid and the base in each Lewis acid-base reaction.
1. $(\text{CH}_3)_2\text{O} + \text{BF}_3 \rightarrow (\text{CH}_3)_2\text{O} : \text{BF}_3$
 2. $\text{H}_2\text{O} + \text{SO}_3 \rightarrow \text{H}_2\text{SO}_4$
 3. $\text{BeCl}_2 + 2\text{Cl}^- \rightarrow \text{BeCl}_4^{2-}$
 4. $\text{BH}_3 + (\text{CH}_3)_2\text{S} \rightarrow \text{H}_3\text{B} : \text{S}(\text{CH}_3)_2$
28. A **Lewis acid** is a substance that accepts a pair of electrons to form a covalent bond. A **Lewis base** is a substance that donates a pair of electrons to form a covalent bond. In a Lewis acid-base reaction, a new covalent bond is formed, but both of the electrons in that bond were originally present on the base.
29. Is the bond formed truly covalent?

30. Why did scientists develop three different classifications for acids and bases?

TABLE 21.4:

Property	Acids	Bases
color with phenolphthalein	colorless	pink
color with litmus	red	blue
taste	sour	bitter
reactivity with metals	good	non-reactive

TABLE 21.5:

Material	Acid or Base Present in Material
drain cleaner	sodium hydroxide
carbonated soda	phosphoric acid
vinegar	acetic acid
antacid	magnesium hydroxide, sodium hydrogen carbonate
cleaners	ammonia
orange	citric acid, ascorbic acid

TABLE 21.6:

Compound	Name
CH ₃ COOH	acetic acid
Mg(OH) ₂	magnesium hydroxide
HClO	hypochlorous acid
H ₃ PO ₄	phosphoric acid
Ca(OH) ₂	calcium hydroxide
KOH	potassium hydroxide
H ₂ SO ₄	sulfuric acid

TABLE 21.7:

Acid	Base	Conjugate Acid	Conjugate Base
HClO ₄	H ₂ O	H ₃ O ⁺	ClO ₄ ⁻
HF	HSO ₃ ⁻	H ₂ SO ₃	F ⁻
HNO ₂	HNO ₂	H ₂ S	NO ₂ ⁻

21.2 The pH Concept

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe the equilibrium process of water ionizing into hydrogen and hydroxide ions, and explain the ion-product constant.

- When two molecules of water collide, a _____ ion is formed.
 - hydrogen
 - hydronium
 - hydrate
 - hydroscopic
- The equilibrium constant for water is
 - 1×10^{-12}
 - 1×10^{14}
 - 1×10^{12}
 - 1×10^{-14}
- True/False: The acidity of water is important for organisms that live in the water.
- True/False: Water is a strong electrolyte.
- Define the following terms:
 - self-ionization of water
 - ion-product of water

Lesson Objective: Calculate the concentration of either the hydrogen ion or the hydroxide ion in water when the other value is known.

- The K_w for water is
 - $[H^+] + [OH^-]$
 - $\frac{[H^+]}{[OH^-]}$
 - $[H^+][OH^-]$
 - $\frac{[OH^-]}{[H^+]}$
- For a neutral solution, the $[H^+]$ is
 - $1 \times 10^{-5} M$
 - $1 \times 10^{-7} M$
 - $1 \times 10^{-6} M$
 - $1 \times 10^{-8} M$
- True/False: The K_w quoted in the text is not influenced by temperature.
- True/False: When HCl is added to water, the equilibrium shifts to the $H_2O(l)$ side.
- Define the following terms:

1. acidic solution
 2. basic solution
11. HBr ionizes completely in water. Calculate both $[H^+]$ and $[OH^-]$ for a solution containing $1.7 \times 10^{-4} M$ HBr.
 12. KOH ionizes completely in water. What are the values for $[H^+]$ and $[OH^-]$ in a $3.7 \times 10^{-2} M$ solution of KOH?
 13. State whether the following solutions are acidic or basic:
 1. $[H^+] = 2.7 \times 10^{-3} M$
 2. $[OH^-] = 2.4 \times 10^{-6} M$

Lesson Objective: Calculate the pH of an aqueous solution of an acid or a base

14. All of the following have pH values below seven except
 1. vinegar
 2. tomato juice
 3. eggs
 4. coffee
15. All the following have pH values greater than seven except
 1. soap
 2. milk
 3. blood
 4. detergent
16. True/False: A 3.0 M solution of HCl has a negative pH
17. True/False: The letters pH stand for hydrogen ion purity.
18. Define pH
19. Fill in the following table:

TABLE 21.8:

Material	Approximate $[H^+]$	Approximate pH
lemon juice		
household bleach		
orange juice		
soap		

20. Calculate the pH of a solution in which the $[H^+]$ is $1.7 \times 10^{-4} M$.

Lesson Objective: Calculate the hydrogen-ion concentration of a solution with a known pH.

21. A solution with a pH value of 6 has a hydrogen ion concentration of
 1. 1×10^6
 2. 1×10^{-3}
 3. 1×10^{-6}
 4. 1×10^3
22. Soda pop has an approximate pH of 3. Its $[H^+]$ is approximately
 1. 1×10^{-4}
 2. 2.5×10^{-2}
 3. 1×10^{-3}

4. 9×10^{-3}
23. True/False: The hydrogen ion concentration of tomato juice is $10^{-5} M$.
24. True/False: Detergent has an approximate $[H^+]$ of $1 \times 10^{-10} M$.
25. Calculate the hydrogen ion concentrations of the following solutions:
1. $pH = 5.6$
 2. $pH = 1.9$
 3. $pH = 8.7$
 4. $pH = 10.5$

Lesson Objective: Calculate the pOH of acids and bases and know the relationships between pH, pOH, hydrogen-ion concentration, and hydroxide-ion concentration.

26. The hydroxide ion concentration is related to the hydrogen ion concentration by
1. $pH = pOH + 14$
 2. $pH + pOH = 14$
 3. $pOH = 14 + pH$
 4. $\frac{pH}{pOH} = 14$
27. A basic solution has a pOH of
1. greater than seven
 2. less than seven
 3. seven
 4. less than five
28. True/False: The $[OH^-]$ can be calculated for a solution if the pH is known.
29. True/False: The pOH value is a negative number
30. Define pOH.
31. Calculate the pOH of the following solutions:
1. $pH = 3.9$
 2. $1 \times 10^{-5} M HCl$
32. The **self-ionization** of water is the process in which water ionizes to hydronium ions and hydroxide ions. As with other aqueous acid-base reactions, the process is often simplified to show the ionization of just one water molecule into a hydrogen ion and a hydroxide ion.
33. $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$
34. Either equation is adequate, though the first is more accurate, since hydrogen ions in aqueous solution will always be attached to water molecules. Further discussion of acids and acid ionizations in this book will primarily show hydrogen ions in aqueous solution as H
35. +
36. , but keep in mind that this is just a commonly used abbreviation for the more accurate hydronium structure. Is it incorrect to talk about [H
37. +
38.] instead of [H
39.]
40. O
41. +
42.]?
43. Why would the equilibrium constant for water omit the water itself as a factor in the equilibrium?

TABLE 21.9:

Material	Approximate $[H^+]$	Approximate pH
lemon juice	10^{-2} M	2
household bleach	10^{-13} M	13
orange juice	10^{-4} M	4
soap	10^{-11} M	11

21.3 Acid and Base Strength

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Explain the difference between a strong acid or base and a weak acid or base.

- When HCl dissolves in water,
 - the molecule reacts with water
 - an equilibrium forms with water
 - 100% ionization occurs
 - partial ionization occurs
- When acetic acid dissolves in water,
 - the molecule reacts with water
 - an equilibrium forms with water
 - 100% ionization occurs
 - partial ionization occurs
- Weak acids ionize to form
 - protons and anions
 - protons and a conjugate base
 - protons and a conjugate acid
 - a conjugate acid and a conjugate base
- All of the acids below are strong acids except
 - nitric acid
 - hydrochloric acid
 - phosphoric acid
 - sulfuric acid
- True/False: Some weak acids are found in food and beverages.
- True/False: A double arrow is used to indicate ionization of strong acids.
- True/False: HCO_3^- is the conjugate base of acetic acid
- True/False: HCN is a weak acid.
- Define the following terms:
 - strong acid
 - weak acid
- Fill in the blanks

TABLE 21.10:

Acid	Conjugate Base
Strong Acids	

TABLE 21.10: (continued)

Acid	Conjugate Base
HCl (hydrochloric acid)	
	HSO ₄ ⁻ (hydrogen sulfate ion)
	NO ₃ ⁻ (nitrate ion)
Weak Acids	
H ₃ PO ₄ (phosphoric acid)	
	CH ₃ COO ⁻ (acetate ion)
	HCO ₃ ⁻ (hydrogen carbonate ion)
HCN (hydrocyanic acid)	

Lesson Objective: Write equilibrium expressions for the ionizations of weak acids and weak bases. Explain how the value of K_a or K_b relates to the strength of the acid or base.

- The acid ionization represents all of the following except
 - the extent of ionization
 - the strength of the acid
 - an indication of undissociated acid
 - the number of anions present
- The acid ionization equation is $K_a =$
 - $\frac{[H^+][A^-]}{[HA]}$
 - $\frac{[H^+]}{[HA]}$
 - $\frac{[H^+][A^-]}{[HA]}$
 - $\frac{[HA]}{[H^+][A^-]}$
- Based on their first acid ionization constant, the weakest acid is
 - oxalic acid
 - benzoic acid
 - acetic acid
 - nitrous acid
- Acetate ion is a stronger base than
 - methylamine
 - urea
 - ammonia
 - pyridine
- True/False: Water is omitted from the expression for the acid ionization constant.
- True/False: The K_a value is very large for a strong acid.
- True/False: Each successive ionization for a polyprotic acid is larger than the previous one.
- True/False: Calcium hydroxide is very water-soluble.
- Define the following terms:
 - acid ionization constant
 - strong base
 - weak base
 - base ionization constant
- Write K_a or K_b expressions for the following equilibria:
 - $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

2. $\text{HOCl} \rightleftharpoons \text{H}^+ + \text{OCl}^-$
3. $\text{F}^- \rightleftharpoons \text{HF} + \text{OH}^-$.

Lesson Objective: Calculate an unknown K_a or K_b from the solution concentration and the pH.

21. The K_a can be determined using a
 1. balance
 2. pH meter
 3. pH titrator
 4. litmus paper
22. True/False: The contribution to $[\text{H}^+]$ from dissociation of water can be ignored when determining K_a .
23. A 0.120 M solution of a generic weak acid (HA) has a pH of 3.26. Determine the K_a .
24. A 0.128 M solution of uric acid ($\text{HC}_5\text{H}_3\text{N}_4\text{O}_3$) has a pH of 2.39. Calculate the K_a of uric acid.
25. Calculate the value of K_b for methylamine, CH_3NH_2 given that the pH of a 0.10 mol/L solution is 11.8

Lesson Objective: Calculate the expected pH of a solution containing a given concentration of a weak acid or base that has a known K_a or K_b value.

26. In calculating the expected pH of a solution of a weak acid, a simplifying assumption that can be made is
 1. concentration of water is constant
 2. most of the weak acid will be ionized
 3. very little of the weak acid will be ionized
 4. the concentration of the weak acid changes significantly
27. True/False: The pH of a solution of a weak acid can be calculated if the ionization constant is known.
28. What is the pH of a 0.100 M solution of acetic acid? $K_a = 1.77 \times 10^{-5}$
29. Some acids, such as sulfuric acid and nitric acid, are strong acids. Working with strong acids requires special care and protective clothing and eyewear. Other acids, such as acetic acid and citric acid, are encountered in many foods and beverages. These weak acids are considerably less dangerous. In this lesson, you will learn about the differences between strong and weak acids and bases.
30. Why is it important to know if an acid is strong or weak?
31. The second K_a for carbonic acid is 4.8×10^{-11} . Can this dissociation really affect the pH of a solution?

TABLE 21.11:

Acid	Conjugate Base
Strong Acids	
HCl (hydrochloric acid)	Cl^- (chloride ion)
H_2SO_4 (sulfuric acid)	HSO_4^- (hydrogen sulfate ion)
HNO_3 (nitric acid)	NO_3^- (nitrate ion)
Weak Acids	
H_3PO_4 (phosphoric acid)	H_2PO_4^- (dihydrogen phosphate ion)
CH_3COOH (acetic acid)	CH_3COO^- (acetate ion)
H_2CO_3 (carbonic acid)	HCO_3^- (hydrogen carbonate ion)
HCN (hydrocyanic acid)	CN^- (cyanide ion)

21.4 Acid-Base Neutralization

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Write balanced molecular and net ionic equations for acid-base neutralization reactions.

- In the neutralization reaction between HCl and NaOH, the Na^+ and Cl^- ions are _____ ions
 - neutral
 - spectator
 - sideline
 - unnecessary
- The neutralization of a weak acid with a strong base produces a solution whose pH is
 - neutral
 - slightly acidic
 - slightly basic
 - strongly basic
- The neutralization of a weak acid with a weak base can produce all of the following except
 - slightly acidic
 - slightly basic
 - neutral
 - cannot predict
- True/False: Net ionic equations of weak acids include the weak acid as a molecule.
- True/False: The net ionic equation for the neutralization of a strong acid by a strong base is $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$.
- True/False: In neutralization reactions between strong acids and strong bases, neither the acid or the base is completely ionized.
- Define the following terms:
 - neutralization reaction
 - salt
- Write complete balanced equations and net ionic equations for the following:
 - hydrochloric acid + potassium hydroxide \rightarrow
 - phosphoric acid + lithium hydroxide \rightarrow

Lesson Objective: Know the steps of an acid-base titration experiment.

- The end-point of a titration is when
 - the first color momentarily appears
 - the color appears and disappears after swirling
 - the faint color stays after swirling
 - the color gets darker

10. The first step in the titration of an acid is
 1. fill the buret with known base
 2. measure out a known volume of acid
 3. slowly add base
 4. add indicator to the acid
11. True/False: The flask is constantly swirled while base is being added.
12. True/False: A graduated cylinder is used to slowly add the base to the acid-indicator mixture.
13. Define the following terms:
 1. equivalence point
 2. titration
 3. indicator
 4. standard solution

Lesson Objective: Calculate the concentration of an acid or a base from data obtained in a titration.

14. In a titration of phosphoric acid with NaOH, the ratio of acid to base for calculations is
 1. $\frac{1 \text{ mol phosphoric acid}}{1 \text{ mol NaOH}}$
 2. $\frac{2 \text{ mol phosphoric acid}}{1 \text{ mol NaOH}}$
 3. $\frac{1 \text{ mol phosphoric acid}}{2 \text{ mol NaOH}}$
 4. $\frac{1 \text{ mol phosphoric acid}}{3 \text{ mol NaOH}}$
15. True/False: $M_A \times V_A = M_B \times V_B$ can be used as is for all titration calculations
16. If it takes 54 mL of 0.1 M NaOH to neutralize 125 mL of an HCl solution, what is the HCl concentration?
17. If it takes 25 mL of 0.05 M HCl to neutralize 345 mL of NaOH solution, what is the NaOH concentration?
18. If it takes 50 mL of 0.5 M KOH solution to neutralize 125 mL of H₂SO₄ solution, what is the sulfuric acid concentration?

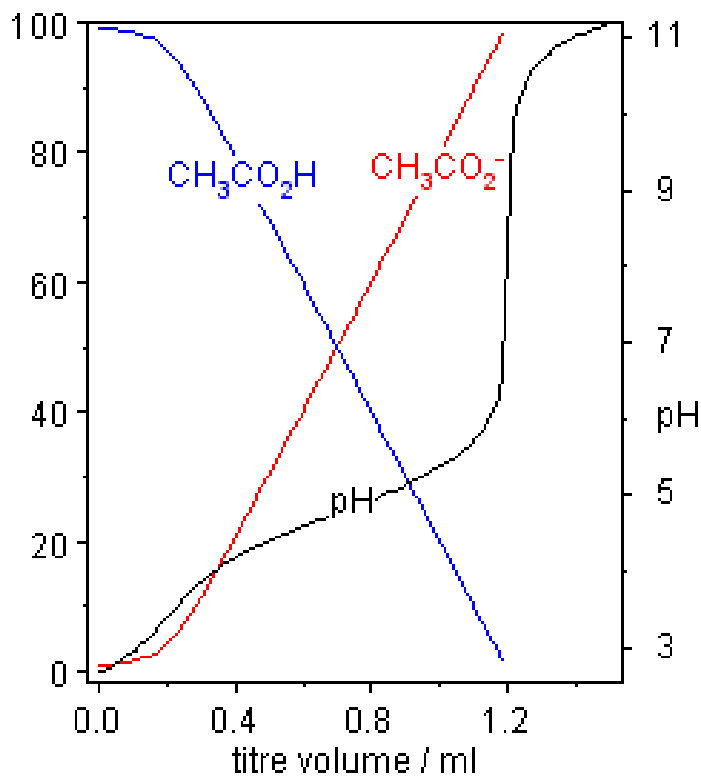
Lesson Objective: Describe the appearance of a titration curve.

19. At the beginning of a titration curve, the pH
 1. changes very slowly
 2. changes rapidly
 3. shows no change initially
 4. fluctuates depending on the acid
20. All of the following are true about the equivalence point except
 1. pH is seven for strong acid-strong base titrations
 2. pH is greater than seven when titrating a weak acid
 3. pH is greater than seven when titrating a weak base
 4. pH is less than seven when titrating a weak base
21. True/False: The titration curves for HCl and H₂SO₄ are identical.
22. True/False: A titration curve plots pH vs molarity of acid or base.
23. Define titration curve.
24. How does the titration curve of a strong acid being titrated by a strong base differ from the curve of a strong base being titrated by a strong acid?

Lesson Objective: Explain the chemistry of acid-base indicators.

25. An indicator ionizes with in a ____ pH unit range

1. 2
 2. 3
 3. 1
 4. 1.5
26. The best indicator for a titration with an equivalence point of 5.5 would be
1. bromocresol green
 2. methyl red
 3. bromocresol purple
 4. bromthymol blue
27. True/False: Thymol Blue indicator is yellow below about pH 8.7
28. True/False: Thymolphthalein is blue below pH 9.
29. An indicator is a weak _____ that _____ within a known pH range.
30. Write the equilibrium equation for methyl orange (use MO as the chemical formula).
31. Methyl red and bromocresol purple both show color changes in the pH 5-6 range. Which would be the better indicator for a material with an equivalence point of pH 6.0?
32. While neutralization reactions are all alike in many ways, it is useful to look at those involving only strong acids and bases separately from those involving a weak acid, a weak base, or both.
33. Why is it useful to look at these reactions separately?
34. Explain the data shown on the graph below.



35.

21.5 Salt Solutions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Predict whether a certain salt solution is acidic, basic, or neutral.

- When LiNO_3 is dissolved in water, the resulting solution will be
 - strongly basic
 - slightly basic
 - neutral
 - slightly acidic
- A solution of NaF in water will have a pH that is
 - strongly basic
 - slightly basic
 - neutral
 - slightly acidic
- A solution of a salt from a strong acid and a weak base will be
 - acidic
 - basic
 - neutral
 - depends on the base
- A solution of a salt from a weak acid and weak base will be
 - acidic
 - basic
 - too difficult to analyze
 - neutral
- The ammonium ion can react with water by
 - being a proton donor
 - being a proton acceptor
 - forming ammonium hydroxide
 - neutralizing the hydroxide ion
- True/False: All ions can react with water and change the pH of the solution.
- True/False: Anions of weak acids will contribute to OH^- formation.
- True/False: The fluoride ion acts as a Brønsted-Lowry base.
- True/False: The chloride ion is the conjugate base of HCl .
- True/False: Ammonium chloride acts as a base in water.
- Define salt hydrolysis.
- Predict whether the solutions made from the following salts will be neutral, acidic, or basic:
 - NaCN

2. K_3PO_4
3. methylamine chloride
13. The ions produced from the _____ of a salt are sometimes capable of undergoing reactions with _____ that change the _____ of the solution.
14. Salts that are derived from the _____ of a weak acid (HF) by a _____ base (NaOH) will always produce salt solutions that are _____.

Lesson Objective: Write balanced equations for hydrolysis reactions.

15. The ammonium ion can react with water by acting as all of the following except
 1. Arrhenius acid
 2. proton donor
 3. Brønsted-Lowry acid
 4. Brønsted-Lowry base
16. Fluoride in water acts as a
 1. proton donor
 2. Brønsted-Lowry acid
 3. Brønsted-Lowry base
 4. Arrhenius acid
17. The ClO^- anion will produce a salt solution that is
 1. slightly basic
 2. slightly acid
 3. neutral
 4. unable to predict
18. True/False: The production of hydronium ions causes the resulting solution to be acidic.
19. True/False: A solution of sodium acetate will be slightly acidic.
20. True/False: Knowledge of K_a or K_b is necessary to calculate the pH of a salt solution.
21. Write balanced equations for the following ions in water solution:
 1. $CH_3NH_3^+$
 2. CH_3COO^- (acetate anion)
 3. ClO^-

Lesson Objective: Calculate the pH of a salt solution when the K_a or the K_b of the hydrolyzing ion is known.

22. Calculations of pH from salt concentrations involves all of the following except
 1. ICE table
 2. knowledge of K_a or K_b
 3. complete equation for process
 4. temperature of system
23. True/False: The chloride ion can accept a proton from water.
24. The K_b of the acetate ion ($HCOO^-$) is 5.56×10^{-11} . What is the pH of 0.35 M solution of sodium formate ($NaHCOO$)?
25. What is the pH of a 0.0510 molar solution of salt NaCN (the K_b for CN^- is 1.62×10^{-5})?

Lesson Objective: Use equations to show how buffers resist changes in pH by using up additional hydrogen or hydroxide ions.

26. A buffer can consist of any of the following except

1. weak acid and salt of its conjugate base
 2. weak acid and salt of its conjugate base
 3. strong acid and salt of its conjugate base
 4. acetic acid and the acetate anion
27. True/False: The most important blood buffer is carbonic acid/hydrogen carbonate.
28. Define the following terms
1. buffer
 2. buffer capacity
29. Write equations to show how the carbonic acid/hydrogen carbonate buffer system neutralizes an increase in $[H^+]$.
30. Buffers are especially critical in biological systems. For example, the pH of blood must be maintained within a fairly narrow range of about 7.35 to 7.45. There are several buffers at work in blood, but the most important one involves carbonic acid (H_2CO_3) and the hydrogen carbonate ion (HCO_3^-). The ratio between the two is carefully maintained through metabolic processes so that the pH of blood is nearly constant.
31. What would be the effect on the buffer system of a decrease in hydrogen ion concentration?
32. H_2CO_3 is an unstable compound when dissolved in water, and readily forms H_2O and CO_2 . How would the amount of dissolved carbon dioxide help balance an increase in hydrogen ion concentration?

TABLE 21.12:

Concentration	$HCOO^-$	$HCOOH$	OH^-
I	0.35 M	0	0
C	$-x$	$+x$	$+x$
E	$0.35 - x$	x	x

TABLE 21.13:

Concentration	CN^-	HCN	OH^-
I	0.051 M	0	0
C	$-x$	$+x$	$+x$
E	$0.051 - x$	x	x

CHAPTER **22**

Oxidation Reduction Reactions Worksheets

Chapter Outline

22.1 THE NATURE OF OXIDATION AND REDUCTION

22.2 OXIDATION NUMBERS

22.3 BALANCING REDOX REACTIONS

22.1 The Nature of Oxidation and Reduction

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Define oxidation and reduction in terms of a gain or loss of oxygen atoms.

- All of the following statements about oxidation are true except
 - combustion is an oxidation process
 - oxidation involves the addition of oxygen
 - rusting of iron is an oxidation process
 - treating iron ore with carbon is an oxidative process
- All of the following statements about reduction are true except
 - bleaches reduce stains by removal of oxygen atoms
 - reduction is the removal of oxygen from a substance
 - iron oxide can be reduced to pure iron
 - every reduction process requires a simultaneous oxidation process
- True/False: Sodium hypochlorite oxidizes stains to make their removal easier.
- True/False: Reduction involves oxygen transfer to a metal.
- Since oxidation was originally defined as the _____ of oxygen, _____ was therefore the removal of oxygen from a substance.
- Many elements simply _____ with oxygen to form the _____ of that element.

Lesson Objective: Define oxidation and reduction in terms of a gain or loss of electrons.

- Oxidation involves
 - a shift of protons
 - a loss of electrons
 - shifting electrons from oxygen to another atom
 - shifting electrons from another atom to oxygen
- True/False: Some oxidation reactions do not involve oxygen.
- Define the following terms
 - oxidation
 - reduction
 - oxidation-reduction reaction

Lesson Objective: Describe oxidation and reduction as it occurs in both ionic and molecular reactions.

- The production of ZnS involves
 - transfer of one electron from Zn to S

2. transfer of two electrons from Zn to S
 3. transfer of one electron from S to Zn
 4. transfer of two electrons from S to Zn
11. All of the following are true about redox reactions except
1. electron loss is oxidation
 2. electron gain is reduction
 3. electron loss is reduction
 4. electrons that are lost must be gained by another particle
12. In the formation of water from O_2 and H_2 all of the following are true except
1. hydrogen atoms are oxidized
 2. oxygen atoms are reduced
 3. in the -O-H bond, the electrons are more attracted to O
 4. H atoms in water have a greater electron density than they did in H_2
13. True/False: H in the water molecule is more electron-rich than O is.
14. True/False: S is the oxidizing agent in the formation of ZnS from Zn and S.
15. True/False: Net ionic equations help identify oxidation and reduction processes.
16. Define the following terms
1. half-reaction
 2. oxidizing agent
 3. reducing agent

Lesson Objective: Identify the oxidizing and reducing agents in a redox reaction.

17. An oxidizing agent
1. reduces another atom
 2. oxidizes another atom
 3. is oxidized by another atom
 4. transfers oxygen from one atom to another
18. A reducing agent
1. reduces another atom
 2. oxidizes another atom
 3. is reduced by another atom
 4. transfers hydrogen from one atom to another
19. True/False: The half reaction $Mg \rightarrow Mg^{2+}$ represents a reduction process
20. True/False: Addition of H to the $O = O$ double bond is a reduction process
21. In the course of a chemical reaction between a metal and a _____, _____ are transferred from the metal atoms to the nonmetal atoms
22. For each of the following reactions, indicate which component is oxidized and which is reduced, then identify the oxidizing agent and the reducing agent:
1. $Zn + 2HCl \rightarrow Zn^{2+} + H_2 + 2Cl^-$
 2. $3CuS + 8HNO_3 \rightarrow 3CuSO_4 + 8NO + 4H_2O$
 3. $4Ag(s) + 2H_2S(g) + O_2(g) \rightarrow 2Ag_2S(s) + 2H_2O(g)$
 4. $SF_4(g) + F_2(g) \rightarrow SF_6(g)$

Lesson Objective: Describe corrosion and some of the ways to prevent it or slow it down

23. All of the following are iron oxidation states that are components of rust except

1. Fe°
 2. Fe^+
 3. Fe^{2+}
 4. Fe^{3+}
24. The green color on the Statue of Liberty is due to
1. CuO
 2. Cu_2O
 3. CuCO_3
 4. CuCl_2
25. True/False: Corrosion costs the U.S economy over \$100 billion per year.
26. True/False: Covering a metal surface with oil or paint is the best way to prevent corrosion.
27. The following reaction illustrates the process of corrosion of zinc metal to form “white rust”
28. $2\text{Zn} + 2\text{H}_2\text{O} + \text{O}_2 \rightarrow \text{Zn}(\text{OH})_2$
29. Which material is oxidized and which is reduced?
30. Oxygen is an element that has been known for centuries. In its pure elemental form, oxygen is highly reactive, and it readily makes compounds with most other elements. It is also the most abundant element by mass in the Earth’s crust. The class of reactions called oxidation and reduction were originally defined with respect to the element oxygen.
31. Why would the first definitions of oxidation and reduction be related to oxygen?
32. Other definitions of reduction include the addition of hydrogen and the gain of electrons. Give one example of each of these examples of reduction.

TABLE 22.1:

Reaction	Oxidized	Reduced	Oxidizing Agent	Reducing Agent
a	Zn to Zn^{2+}	H^+ to H	H^+	Zn
b	S^{2-} to S^{+6}	N^{+5} to N^{+2}	N^{+5}	S^{-2}
c	Ag to Ag^+	O to O^{2-}	O	Ag
d	S^{+4} to S^{+6}	F to F^{2-}	F	S^{+4}

22.2 Oxidation Numbers

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use the oxidation number rules to determine the oxidation number of an atom of any element in a pure substance.

- Oxygen can have all of the following oxidations numbers except
 - +1
 - +2
 - 3
 - 2
- The only halogen that does not have a variable oxidation number is
 - Cl
 - F
 - I
 - Br
- The oxidation number for H in NaBH_4 is
 - +1
 - +2
 - 1
 - 2
- True/False: The oxidation number of an atom in an element is zero.
- True/False: The oxidation number for S in compounds depends on the structure of the compound.
- True/False: In a polyatomic ion, the sum of oxidation numbers of the atoms equals zero.
- Define oxidation number.
- Indicate the oxidation numbers of the following atoms in compounds:

TABLE 22.2:

Atom	Compound	Oxidation Number
Cl	$\text{Fe}(\text{ClO}_2)_3$	
Mn	KMnO_4	
N	N_2O_3	
S	Al_2S_3	

Lesson Objective: Define oxidation and reduction in terms of a change in oxidation number.

- The oxidation process involves

1. loss of oxygen
 2. gain of hydrogen
 3. complete loss of electrons
 4. decrease in oxidation number
10. The reduction process involves
1. loss of oxygen
 2. increase in oxidation number
 3. complete loss of electrons
 4. gain of hydrogen
11. In a reaction, KNO_2 is converted to KNO_3 . The oxidation-reduction change was
1. $\text{K}^+ \rightarrow \text{K}^{2+}$
 2. $\text{O}^- \rightarrow \text{O}^{2-}$
 3. $\text{N}^{+3} \rightarrow \text{N}^{+4}$
 4. $\text{N}^{+3} \rightarrow \text{N}^{+5}$
12. True/False: A shift of electrons towards an atom in a covalent bond is an oxidation of that atom.
13. True/False: The change in Cl from KClO_3 to KCl represents a reduction of Cl.
14. True/False: The change in N from HNO_3 to NaNO_3 represents an oxidation of N.
15. State whether the indicated atom is oxidized or reduced as a result of the reaction.

TABLE 22.3:

Atom	Reactant	Product	Oxidized or Reduced?
S	S	SO_2	
Mn	KMnO_4	Mn_2O_3	
V	V_2O_3	VO	
Sn	SnCl_2	SnCl_4	

Lesson Objective: Identify which atoms are being oxidized and which atoms are being reduced in a redox reaction.

16. For the reaction below, all of the following statements are correct except one
17. $\text{Mn}(\text{NO}_3)_2 + \text{NaBiO}_3 + \text{HNO}_3 \rightarrow \text{HMnO}_4 + \text{Bi}(\text{NO}_3)_3 + \text{NaNO}_3 + \text{H}_2\text{O}$
1. Mn is oxidized
 2. Na is reduced
 3. Bi is reduced
 4. N is unchanged
18. True/False: In the reaction $\text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2$, Mg is oxidized
19. For each of the following reactions, list the atom(s) oxidized and the atom(s) reduced:
1. $\text{KMnO}_4 + \text{KNO}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{MnSO}_4 + \text{H}_2\text{O} + \text{KNO}_3 + \text{K}_2\text{SO}_4$
 2. $4\text{NH}_3 + 5\text{O}_2 \rightarrow 4\text{NO} + 6\text{H}_2\text{O}$
 3. $4\text{NaOH} + \text{Ca}(\text{OH})_2 + \text{C} + 4\text{ClO}_2 \rightarrow 4\text{NaClO}_2 + \text{CaCO}_3 + 3\text{H}_2\text{O}$

Lesson Objective: Know which types of reactions are usually redox reactions and which types are usually not. Use oxidation number changes to decide whether a given reaction involves oxidation and reduction.

19. A redox reaction must involve a change in oxidation number for at least ____ of the atoms in the reaction
1. one
 2. two

3. three
 4. four
20. Single replacement reactions are redox reactions because of all of the following except
 1. a free element is transformed to an ionized material
 2. an ionized material is transformed to a free element
 3. a cation becomes an anion
 4. a material with an oxidation number of zero experiences a change in oxidation number
 21. Combustion reactions are redox reactions for all of the following reasons except
 1. oxygen acts as a reducing agent
 2. oxygen is reduced
 3. carbon has oxygen added to form CO_2
 4. elemental oxygen is incorporated into a compound
 22. True/False: Double-displacement reactions are not redox reactions.
 23. True/False: The reaction of sodium metal with water to form NaOH and hydrogen gas is a redox reaction.
 24. True/False: The neutralization of HCl by NaOH is a redox reaction.
 25. Indicate which reactions shown below are redox reactions:
 1. $\text{Pb} + \text{FeSO}_4 \rightarrow \text{PbSO}_4 + \text{Fe}$
 2. $\text{Na}_3\text{PO}_4 + 3\text{KOH} \rightarrow 3\text{NaOH} + \text{K}_3\text{PO}_4$
 3. $\text{MgCl}_2 + \text{Li}_2\text{CO}_3 \rightarrow \text{MgCO}_3 + 2\text{LiCl}$
 4. $4\text{C}_5\text{H}_9\text{O} + 27\text{O}_2 \rightarrow 20\text{CO}_2 + 18\text{H}_2\text{O}$
 5. $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
 26. Overall, the oxidation number of an atom in a molecule is the charge that the atom would have if all polar covalent and ionic bonds resulted in a complete transfer of electrons from the less electronegative atom to the more electronegative one.
 27. Why is this explanation necessary?
 28. Since redox reactions involve the transfer of electrons, could this phenomenon be used in electrical circuits?

TABLE 22.4:

Atom	Compound	Oxidation Number
Cl	$\text{Fe}(\text{ClO}_2)_3$	+3
Mn	KMnO_4	+7
N	N_2O_3	+3
S	Al_2S_3	-2

TABLE 22.5:

Atom	Reactant	Product	Oxidized or Reduced?
S	S	SO_2	oxidized
Mn	KMnO_4	Mn_2O_3	reduced
V	V_2O_3	VO	reduced
Sn	SnCl_2	SnCl_4	oxidized

22.3 Balancing Redox Reactions

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Balance a redox equation using the oxidation-number-change method.

- One approach to balancing redox equations is called the
 - oxidation number method
 - oxidation number change method
 - oxidation change method
 - oxidation valance method
- The first step in this approach is to
 - balance the equation
 - identify what is oxidized and what is reduced
 - assign oxidation numbers
 - check to see if the equation is balance in terms of both atoms and charges
- The second step is to
 - balance the equation
 - identify what is oxidized and what is reduced
 - assign oxidation numbers
 - check to see if the equation is balance in terms of both atoms and charges
- Step three in the process is to
 - write half-reactions for the changes
 - use coefficients to balance the equation
 - use lines to connect atoms undergoing oxidation or reduction
 - check to see if the equation is balanced
- Step five is to
 - balance the equation
 - identify what is oxidized and what is reduced
 - assign oxidation numbers
 - check to see if the equation is balanced in terms of both atoms and charges
- True/False: Balancing solely by inspection is always an easy way to balance redox equations.
- True/False: Monitoring how the oxidation numbers change is good practice.
- True/False: Oxidation numbers are assigned to each atom in the equation.
- True/False: Coefficients are needed only for those materials directly involved in the redox process.
- True/False: The total increase in oxidation number must equal the total decrease in oxidation number.
- Balance the following equations using the oxidation number change method:
 - $HNO_3 + H_3AsO_3 \rightarrow NO + H_3AsO_4 + H_2O$
 - $HNO_3 + H_3AsO_3 \rightarrow NO + H_3AsO_4 + H_2O$

3. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$
4. $NO_2 + H_2 \rightarrow NH_3 + H_2O$
5. $Al + MnO_2 \rightarrow Al_2O_3 + Mn$

Lesson Objective: Balance a redox equation by breaking it down into oxidation and reduction half-reactions and using the half-reaction method.

12. The first step in balancing equations using the half-reaction method is to
 1. write the balanced ionic equation
 2. write the half-reactions
 3. write the unbalanced ionic equation
 4. add electrons to each half
13. The second step is to
 1. write the half-reactions
 2. add water where needed
 3. add acid where needed
 4. add electrons
14. The third step is to
 1. add water molecules where needed
 2. balance each half-reaction
 3. add acid where needed
 4. add water where needed
15. The fifth step is to
 1. add electrons where needed
 2. add water where needed
 3. multiply by a coefficient
 4. add the two half-reactions together
16. The sixth step is to
 1. add electrons where needed
 2. add water where needed
 3. multiply by a coefficient
 4. add the two half-reactions together
17. True/False: Reactions in a basic medium require addition of hydroxide ions.
18. True/False: Oxidation of Fe^{2+} by $Cr_2O_7^{2-}$ readily occurs in basic solution.
19. True/False: The half-reaction method is more useful for reactions involving ions in water.
20. True/False: Half-reactions treat oxidation and reduction as two sequential processes.
21. True/False: Electrons are added to each half-reaction to balance the charges.
22. Write half-reactions for the following redox processes:
 1. $Fe^{2+} + MnO_4^- \rightarrow Fe^{3+} + Mn^{2+}$
 2. $3CuS + 8HNO_3 \rightarrow 3CuSO_4 + 8NO + 4H_2O$
 3. $4Ag + 2H_2S + O_2 \rightarrow 2Ag_2S + 2H_2O$
23. Balance the following equations using the half-reaction method:
 1. $Cu + HNO_3 \rightarrow Cu^{2+} + NO$
 2. $ClO_3^- + SO_2 \rightarrow SO_4^{2-} + Cl^-$
 3. $MnO_4^- + H_2S \rightarrow Mn^{2+} + S_8$

CHAPTER 23**Electrochemistry
Worksheets****Chapter Outline**

23.1 ELECTROCHEMICAL CELLS**23.2 CELL POTENTIALS****23.3 ELECTROLYSIS**

23.1 Electrochemical Cells

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Use the activity series to identify elements that are more easily oxidized than others, and write oxidation and reduction half-reactions.

- All of the following are applications of electrochemistry except
 - providing power to devices
 - removing electrons for storage
 - plate metals
 - nerve transmission
- Nickel metal will oxidize when in a solution with _____ cations
 - magnesium
 - barium
 - copper
 - potassium
- Aluminum metal will not oxidize when in a solution of _____ cations.
 - magnesium
 - mercury
 - iron
 - lead
- The reaction involving zinc metal immersed in a solution of copper ions is called a ____ reaction.
 - indirect
 - oscillating
 - direct
 - solution
- True/False: Some chemical reactions can involve production of electricity.
- True/False: Oxidation and reduction reactions occur sequentially.
- True/False: Lithium will spontaneously oxidize in a solution of silver nitrate.
- True/False: Iron will spontaneously oxidize in a solution of sodium chloride.
- Define the following terms:
 - electrochemistry
 - electrochemical cell
- The electrons that are transferred in the reaction go _____ from the Zn atoms on the _____ of the strip to the Cu^{2+} ions that are in the _____ adjacent to the zinc strip.
- In each of the following combinations, indicate whether or not a reaction will occur.
 - $\text{NaCl}(aq) + \text{Ni}(s)$
 - $\text{Ca}(s) + \text{ZnCl}_2(aq)$

- $CuCl_2(aq) + Ag(s)$
- $K(s) + CaCl_2(aq)$

Lesson Objective: Describe the parts of a voltaic cell and explain how redox reactions are used to generate an electric current.

- Volta constructed a cell that generated
 - alternating current
 - fluctuating current
 - intermittent current
 - direct current
- The cell notation for the zinc-copper cell is
 - $Cu(s)|Cu^{2+}(1M)||Zn^{2+}(1M)|Zn(s)$
 - $Zn(s)|Zn^{2+}(1M)||Cu^{2+}(1M)|Cu(s)$
 - $Cu^{2+}(1M)|Cu(s)||Zn^{2+}(1M)|Zn(s)$
 - $Zn^{2+}(1M)|Zn(s)||Cu^{2+}(1M)|Cu(s)$
- All of the following are indicated in a voltaic cell except
 - salt bridge
 - anode
 - phase boundaries
 - salt bridge composition
- True/False: A non-porous plug is used in the salt bridge.
- True/False: A metal wire connects the two electrodes.
- True/False: The anode is conventionally written on the right side of the notation.
- Define the following terms
 - voltaic cell
 - half-cell
 - electrode
 - salt bridge
- Write the shorthand notation for a galvanic cell that uses the reaction
- $Fe(s) + Sn^{2+}(aq) \rightarrow Fe^{2+}(aq) + Sn(s)$
- Assume any solutions to be 1.0 M

Lesson Objective: Describe the general features of a dry cell, a lead storage battery, and a fuel cell.

- One of the following is not an example of a voltaic cell
 - dry cell
 - fuel cell
 - solar cell
 - lead storage battery
- One of the following is not a component of the dry cell paste
 - MnO_2
 - ZnO_2
 - NH_4Cl
 - $ZnCl_2$
- The cathode in a lead storage battery is
 - PbO_2

2. PbO
 3. Pb
 4. PbSO₄
23. True/False: The anode for a dry cell battery is the Zn container.
 24. True/False: The dry cell battery uses a salt bridge.
 25. True/False: In a fuel cell, hydrogen gas is oxidized to water.
 26. Define the following terms
 1. battery
 2. fuel cell
 27. Write the overall reactions for
 - a. a zinc dry cell battery
 - b. an alkaline battery
 28. Write half-reactions for the recharging of a lead storage battery.
 29. The inert electrolyte in the salt bridge is often potassium chloride (KCl) or sodium nitrate (NaNO₃).
 30. Why would these two salts be considered to be inert electrolytes?
 31. Alessandro Volta tested the strength of the early electrochemical cells he developed by measuring the amount of shock they produced. Discuss the pros and cons of this experimental approach.

23.2 Cell Potentials

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Describe how an electrical potential is generated in an electrochemical cell.

1. The electrical potential for a cell comes about as a result of competition for
 1. protons
 2. electrons
 3. hydrogen ions
 4. ions
2. Difference in electrical potential between two _____ produces voltage.
 1. ions
 2. electrodes
 3. half-cells
 4. anodes
3. True/False: The electrical potential of an isolated half-cell can be measured.
4. True/False: Spontaneous redox reactions produce electrical energy.
5. Define the following terms:
 1. electrical potential
 2. reduction potential
 3. cell potential

Lesson Objective: Describe the standard hydrogen electrode and how it is used to determine the standard cell potentials of other half-cells.

6. Determination of standard cell potential requires an electrochemical cell with aqueous components present at _____ concentration.
 1. 0.1 M
 2. 0.5 M
 3. 2.0 M
 4. 1.0 M
7. The standard potential for the hydrogen electrode is
 1. 0.00 V
 2. 0.10 V
 3. 0.01 mv
 4. 0.10 mv
8. A negative 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
9. True/False: The standard hydrogen electrode is immersed in 1.0 M H^+ .
10. True/False: The metal for the electrode is made of gold.
11. True/False: Hydrogen gas is formed in all applications of the hydrogen electrode.
12. Define the following terms:
 1. standard cell potential
 2. standard hydrogen electrode
13. Why is Pt used as the metal for the standard hydrogen electrode?
14. How can the SHE serve both as a cathode and an anode?

Lesson Objective: Calculate the standard cell potentials from a table of standard reduction potentials.

15. The half-cell with the higher reduction potential will
 1. undergo reduction
 2. undergo oxidation
 3. cause the other cell to reduce
 4. lose electrons
16. Co^{2+} will undergo reduction when paired with half-cells made of all of the following except
 1. Cd^{2+}
 2. Sn^{2+}
 3. Ca^{2+}
 4. Cs^+
17. The cell potential can be calculated using the formula
 1. $E_{ox} - E_{red}$
 2. $E_{ox} \div E_{red}$
 3. $E_{red} - E_{ox}$
 4. $E_{red} \div E_{ox}$
18. True/False: The cell potential must be negative for the redox reaction to be spontaneous.
19. True/False: In the voltaic cell using $Zn|Zn^{2+}$ and $Fe|Fe^{2+}$, Fe^{2+} will be reduced.
20. True/False: Lithium ions reduce easily.
21. For the following cell combinations calculate the standard cell potential and write the overall cell reaction:
 1. $Co|Co^{2+}$ and $Pb|Pb^{2+}$
 2. $Cu|Cu^+$ and $Be|Be^{2+}$
 3. $Mg|Mg^{2+}$ and $Sn|Sn^{4+}$

Lesson Objective: Predict the behavior of oxidizing and reducing agents based on their position in the table of standard reduction potentials.

22. A strong oxidizing agent is one that can
 1. pull electrons away from other species
 2. donate electrons to other species
 3. transfer electrons to other species
 4. convert metallic ions to metals
23. Cr will reduce all of the following except
 1. Ag^+

2. Br_2
 3. Mn^{2+}
 4. Pb^{2+}
24. True/False: Fluorine is the strongest oxidizing agent.
 25. True/False: Metals will reduce metals listed below them in the standard reduction potential table.
 26. Which of the following metals will react with HCl? Hg, Sn, Cr, Mg
 27. Rank the following oxidizing agents with 1 as strongest and 4 as weakest: Cd^{2+} , Al^{3+} , Au^{3+} , Ni^{2+} .
 28. The electrical potential of a cell results from a competition for electrons. In the zinc-copper voltaic cell described in the previous lesson, the copper(II) ions were reduced to copper metal. That is because the Cu^{2+} ions have a greater attraction for electrons than the Zn^{2+} ions in the other half-cell. Instead, the zinc metal is oxidized. A **reduction potential** *measures the tendency of a given half-reaction to occur as a reduction in an electrochemical cell*. In a given voltaic cell, the half-cell that has the greater reduction potential is the one in which reduction will occur.
 29. Is there any advantage to using the term reduction potential or would oxidation potential be just as useful?
 30. Lead in water supplies is a serious problem in many parts of the world. Your local chemistry genius is proposing to bubble fluorine gas into water supplies to precipitate the lead. Do you think this is a good idea. Explain your answer.

23.3 Electrolysis

Worksheet

Name _____ Class _____ Date _____

Answer each of the questions below to show your achievement of the lesson objectives

Lesson Objective: Distinguish between voltaic and electrolytic cells.

- All of the following are true statements except
 - voltaic cells use spontaneous reactions to generate electric current
 - nonspontaneous reactions are driven by an external electrical current
 - products are favored over reactants in a nonspontaneous reaction
 - half cells are involved in both voltaic and electrolytic processes
- In an electrolytic process involving Zn and Cu
 - the Zn electrode is the anode
 - the Cu electrode is the anode
 - Zn is oxidized
 - Cu is reduced
- True/False: Electrolysis is used to plate silver on jewelry.
- True/False: A voltaic cell uses an electric current to initiate a spontaneous reaction.
- Define the following terms:
 - electrolysis
 - electrolytic cell

Lesson Objective: Describe a Down's cell, and identify the products of the electrolysis of molten sodium chloride.

- In the Down's cell
 - molten NaCl reacts at the cathode to form metallic sodium
 - molten NaCl reacts at the anode to form metallic sodium
 - chlorine ions are reduced to form chlorine gas
 - chlorine ions are converted to chlorine gas at the cathode
- The overall Down's cell requires over _____ volts for the reaction to occur
 - 2.5
 - 3.4
 - 1.8
 - 4.1
- True/False: The Down's cell is a minor producer of sodium metal.
- True/False: Chlorine is used as a disinfectant.
- Write the half-reactions and the overall reaction for the Down's process.
- Would you expect the melted Na to be more or less dense than melted NaCl? Explain your answer.

12. Solid NaCl is a poor conductor of electricity. Why would melted NaCl conduct electricity well?

Lesson Objective: Describe the reactions that occur during the electrolysis of water.

13. All of the following are components of the apparatus for electrolysis of water except
1. sulfuric acid
 2. platinum electrode
 3. power supply
 4. gold electrode
14. Oxygen is formed at the
1. anode
 2. cathode
 3. electrode
 4. electrical interface
15. True/False: Pure water will effectively conduct an electric current
16. True/False: Water is simultaneously oxidized and reduced.
17. Write the half-reactions and the overall reaction for the electrolysis of water.
18. Describe the Hoffman apparatus for the electrolysis of water. Include a diagram.

Lesson Objective: Identify the products that would be generated during the electrolysis of an aqueous solution of sodium chloride.

19. One of the following is not a product of brine electrolysis
1. Cl₂
 2. Na
 3. H₂
 4. O₂
20. The reduction potential for water is
1. -0.70 V
 2. -0.80 V
 3. -0.83 V
 4. -0.89 V
21. True/False: Chlorine gas is produced at the cathode
22. True/False: NaOH can be obtained by evaporating the water.
23. Write the half-reactions and the overall reaction for the electrolysis of brine.
24. Could you use a standard apparatus for electrolysis of water to carry out the electrolysis of brine? Explain your answer.

Lesson Objective: Describe the process of electroplating.

25. All of the following are true statements about electroplating except
1. silver is both oxidized and reduced
 2. a silver strip serves as the anode
 3. silver nitrate solution is used in the process
 4. silver metal is oxidized to provide Ag⁺ for plating
26. In electroplating, silver metal is deposited on a
1. stainless steel spoon

2. silver fork
 3. platinum rod
 4. aluminum foil
27. True/False: The spoon serves as the cathode in the process.
 28. True/False: The Ag^+ concentration gradually changes.
 29. Define electroplating.
 30. If the silver nitrate solution provides the ions for electroplating, why is the silver strip needed?
 31. A nonspontaneous reaction is one in which the reactants are favored over the products under a given set of reaction conditions. However, if a chemical system is supplied with energy from an external source, it is possible to drive a reaction in the nonspontaneous direction.
 32. Why would a nonspontaneous reaction be needed?
 33. How would you electroplate copper (hint: look at the table in lesson 23-2)?