Glencoe Science

Student Edition

California Standards Practice for Chemistry

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Glencoe

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To the Student

Welcome to the Student Edition of *California Standards Practice for Chemistry.*

Task Regimen

Task	At-Home Assignment	In-Class Assignment
Task 1	Using an answer key from the teacher, locate and review any questions you missed. Place a question mark beside any question you do not understand, and bring it to class for discussion.	The teacher administers the test in a realistic test-taking environment.
Task 2	For each question you missed, find the pages in the textbook that cover the material and explain what specific information was needed to answer the question correctly. If you cannot find any helpful information in the textbook, write out three questions you have about the test question.	Work in a group to discuss any confusing questions and content areas. Then work through the confusing questions together.
Task 3	For every incorrect question, go through each answer choice and explain why it is correct or incorrect. Include any tips or hints you noticed that helped you eliminate choices. Place a question mark beside any question you do not understand, and bring it to class for discussion.	Your teacher will lead a discussion for each question. Share your ideas and observations with the class. Keep notes of the discussion to help you review.
Task 4	Your teacher will provide you with a list of questions to practice. For each question, make observations and write down all of the information given in the test in the form of a graphic, a passage, or otherwise. Write the information directly on the test.	Work in a group to discuss each question. Make sure to note the location in the textbook where helpful information was found.

Test-Taking Tips: Student

Before the Test

- Be sure to get plenty of sleep the week before the test. A healthy amount of sleep is eight to nine hours every night.
- On the night before the test, try to do something relaxing but stimulating, such as playing a board game, exercising, or reading an enjoyable book. Cramming the night before the test can often hamper your memory and make you tired.
- On the morning of the test, eat a healthy breakfast with fresh foods that are high in protein and carbohydrates.
- On the morning of the test, clear your mind of any outside distractions so that you will be better able to focus on the test. If breaks are given during the test, use that time to relax and clear your mind.

During the Test

- Listen to and read all directions.
- Be sure you understand the question before reading the answer choices. Then, make sure to read and consider **every** answer choice.
- Remember to carefully consider all the information presented in the test's graphics.
- If the test is timed, be sure to pace yourself.
- Always choose an answer. By eliminating as many incorrect choices as possible, you will have a good chance of guessing correctly and obtaining more points.

California Standards for Chemistry

Atomic and Molecular Structure

1. **The periodic table displays the elements in increasing atomic number and shows how periodicity of the physical and chemical properties of the elements relates to atomic structure. As a basis for understanding this concept:**
   a. Students know how to relate the position of an element in the periodic table to its atomic number and atomic mass.
   b. Students know how to use the periodic table to identify metals, semimetals, nonmetals, and halogens.
   c. Students know how to use the periodic table to identify alkali metals, alkaline earth metals and transition metals, trends in ionization energy, electronegativity, and the relative sizes of ions and atoms.
   d. Students know how to use the periodic table to determine the number of electrons available for bonding.
   e. Students know the nucleus of the atom is much smaller than the atom yet contains most of its mass.
   f.* Students know how to use the periodic table to identify the lanthanide, actinide, and transactinide elements and know that the transuranium elements were synthesized and identified in laboratory experiments through the use of nuclear accelerators.
   g.* Students know how to relate the position of an element in the periodic table to its quantum electron configuration and to its reactivity with other elements in the table.
   h.* Students know the experimental basis for Thomson's discovery of the electron, Rutherford's nuclear atom, Millikan's oil drop experiment, and Einstein's explanation of the photoelectric effect.
   i.* Students know the experimental basis for the development of the quantum theory of atomic structure and the historical importance of the Bohr model of the atom.
   j.* Students know that spectral lines are the result of transitions of electrons between energy levels and that these lines correspond to photons with a frequency related to the energy spacing between levels by using Planck's relationship ($E=hv$).

Chemical Bonds

2. Biological, chemical, and physical properties of matter result from the ability of atoms to form bonds from electrostatic forces between electrons and protons and between atoms and molecules. As a basis for understanding this concept:

a. Students know atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.

b. Students know chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, Cl_2, and many large biological molecules are covalent.

c. Students know salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction.

d. Students know the atoms and molecules in liquids move in a random pattern relative to one another because the intermolecular forces are too weak to hold the atoms or molecules in a solid form.

e. Students know how to draw Lewis dot structures.

f.* Students know how to predict the shape of simple molecules and their polarity from Lewis dot structures.

g.* Students know how electronegativity and ionization energy relate to bond formation.

h.* Students know how to identify solids and liquids held together by van der Waals forces or hydrogen bonding and relate these forces to volatility and boiling/ melting point temperatures.

Conservation of Matter and Stoichiometry

3. The conservation of atoms in chemical reactions leads to the principle of conservation of matter and the ability to calculate the mass of products and reactants. As a basis for understanding this concept:

a. Students know how to describe chemical reactions by writing balanced equations.

b. Students know the quantity one mole is set by defining one mole of carbon 12 atoms to have a mass of exactly 12 grams.

c. Students know one mole equals 6.02×10^{23} particles (atoms or molecules).

d. Students know how to determine the molar mass of a molecule from its chemical formula and a table of atomic masses and how to convert the mass of a molecular substance to moles, number of particles, or volume of gas at standard temperature and pressure.

e. Students know how to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

f.* Students know how to calculate percent yield in a chemical reaction.

g.* Students know how to identify reactions that involve oxidation and reduction and how to balance oxidation-reduction reactions.

Gases and Their Properties

4. The kinetic molecular theory describes the motion of atoms and molecules and explains the properties of gases. As a basis for understanding this concept:

a. Students know the random motion of molecules and their collisions with a surface create the observable pressure on that surface.
b. Students know the random motion of molecules explains the diffusion of gases.
c. Students know how to apply the gas laws to relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases.
d. Students know the values and meanings of standard temperature and pressure (STP).
e. Students know how to convert between the Celsius and Kelvin temperature scales. f. Students know there is no temperature lower than 0 Kelvin.
g.* Students know the kinetic theory of gases relates the absolute temperature of a gas to the average kinetic energy of its molecules or atoms.
h.* Students know how to solve problems by using the ideal gas law in the form $PV=nRT$.
i.* Students know how to apply Dalton's law of partial pressures to describe the composition of gases and Graham's law to predict diffusion of gases.

Acids and Bases

5. Acids, bases, and salts are three classes of compounds that form ions in water solutions. As a basis for understanding this concept:

a. Students know the observable properties of acids, bases, and salt solutions.
b. Students know acids are hydrogen-ion-donating and bases are hydrogen-ion-accepting substances.
c. Students know strong acids and bases fully dissociate and weak acids and bases partially dissociate.
d. Students know how to use the pH scale to characterize acid and base solutions.
e.* Students know the Arrhenius, Brønsted-Lowry, and Lewis acid–base definitions.
f.* Students know how to calculate pH from the hydrogen-ion concentration.
g.* Students know buffers stabilize pH in acid–base reactions.

Solutions

6. **Solutions are homogeneous mixtures of two or more substances. As a basis for understanding this concept:**
   a. Students know the definitions of solute and solvent.
   b. Students know how to describe the dissolving process at the molecular level by using the concept of random molecular motion.
   c. Students know temperature, pressure, and surface area affect the dissolving process.
   d. Students know how to calculate the concentration of a solute in terms of grams per liter, molarity, parts per million, and percent composition.
   e.* Students know the relationship between the molality of a solute in a solution and the solution's depressed freezing point or elevated boiling point.
   f.* Students know how molecules in a solution are separated or purified by the methods of chromatography and distillation.

Chemical Thermodynamics

7. **Energy is exchanged or transformed in all chemical reactions and physical changes of matter. As a basis for understanding this concept:**
   a. Students know how to describe temperature and heat flow in terms of the motion of molecules (or atoms).
   b. Students know chemical processes can either release (exothermic) or absorb (endothermic) thermal energy.
   c. Students know energy is released when a material condenses or freezes and is absorbed when a material evaporates or melts.
   d. Students know how to solve problems involving heat flow and temperature changes, using known values of specific heat and latent heat of phase change.
   e.* Students know how to apply Hess's law to calculate enthalpy change in a reaction.
   f.* Students know how to use the Gibbs free energy equation to determine whether a reaction would be spontaneous.

Reaction Rates

8. Chemical reaction rates depend on factors that influence the frequency of collision of reactant molecules. As a basis for understanding this concept:

a. Students know the rate of reaction is the decrease in concentration of reactants or the increase in concentration of products with time.

b. Students know how reaction rates depend on such factors as concentration, temperature, and pressure.

c. Students know the role a catalyst plays in increasing the reaction rate.

d.* Students know the definition and role of activation energy in a chemical reaction.

Chemical Equilibrium

9. Chemical equilibrium is a dynamic process at the molecular level. As a basis for understanding this concept:

a. Students know how to use Le Chatelier's principle to predict the effect of changes in concentration, temperature, and pressure.

b. Students know equilibrium is established when forward and reverse reaction rates are equal.

c.* Students know how to write and calculate an equilibrium constant expression for a reaction.

Organic Chemistry and Biochemistry

10. The bonding characteristics of carbon allow the formation of many different organic molecules of varied sizes, shapes, and chemical properties and provide the biochemicalbasis of life. As a basis for understanding this concept:

a. Students know large molecules (polymers), such as proteins, nucleic acids, and starch, are formed by repetitive combinations of simple subunits.

b. Students know the bonding characteristics of carbon that result in the formation of a large variety of structures ranging from simple hydrocarbons to complex polymers and biological molecules.

c. Students know amino acids are the building blocks of proteins.

d.* Students know the system for naming the ten simplest linear hydrocarbons and isomers that contain single bonds, simple hydrocarbons with double and triple bonds, and simple molecules that contain a benzene ring.

e.* Students know how to identify the functional groups that form the basis of alcohols, ketones, ethers, amines, esters, aldehydes, and organic acids.

f.* Students know the R-group structure of amino acids and know how they combine to form the polypeptide backbone structure of proteins.

Nuclear Processes

11. **Nuclear processes are those in which an atomic nucleus changes, including radioactive decay of naturally occurring and human-made isotopes, nuclear fission, and nuclear fusion. As a basis for understanding this concept:**
    a. Students know protons and neutrons in the nucleus are held together by nuclear forces that overcome the electromagnetic repulsion between the protons.
    b. Students know the energy release per gram of material is much larger in nuclear fusion or fission reactions than in chemical reactions. The change in mass (calculated by $E=mc^2$) is small but significant in nuclear reactions.
    c. Students know some naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.
    d. Students know the three most common forms of radioactive decay (alpha, beta, and gamma) and know how the nucleus changes in each type of decay.
    e. Students know alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.
    f.* Students know how to calculate the amount of a radioactive substance remaining after an integral number of half-lives have passed.
    g.* Students know protons and neutrons have substructures and consist of particles called quarks.

Name ____

Student Recording Chart

Directions: Circle each question from the Diagnostic Test that you answered *incorrectly*. If there are one or two circles marked for an indicator, write ***Yes*** in the ***Need Practice?*** box. Then complete the practice pages for that indicator.

Indicator	**1. a.**	**1. b.**	**1. c.**	**1. d.**	**1. e.**	**1. f.**	**1. g.**	**1. h.**	**1. i.**	**1. j.**
Test Questions	19	52	38	60	11	51	39	58	70	13
Need Practice?										
Practice Pages	9	9	10	10	11	11	12	12	13	13

Indicator	**2. a.**	**2. b.**	**2. c.**	**2. d.**	**2. e.**	**2. f.**	**2. g.**	**2. h.**
Test Questions	18	26	47	16	1	59	63	43
Need Practice?								
Practice Pages	14	14	15	15	16	16	17	17

Indicator	**3. a.**	**3. b.**	**3. c.**	**3. d.**	**3. e.**	**3. f.**	**3. g.**
Test Questions	45	48	32	62	66	37	31
Need Practice?							
Practice Pages	18	18	19	19	20	20	21

Indicator	**4. a.**	**4. b.**	**4. c.**	**4. d.**	**4. e.**	**4. f.**	**4. g.**	**4. h.**	**4. i.**
Test Questions	7	29	6	55	25	64	69	50	57
Need Practice?									
Practice Pages	22	22	23	23	24	24	25	25	26

Name ____

Student Recording Chart *(continued)*

Indicator	5. a.	5. b.	5. c.	5. d.	5. e.	5. f.	5. g.
Test Questions	17	68	53	28	71	44	20
Need Practice?							
Practice Pages	27	27	28	28	29	29	30

Indicator	6. a.	6. b.	6. c.	6. d.	6. e.	6. f.
Test Questions	54	15	56	30	41	46
Need Practice?						
Practice Pages	31	31	32	32	33	33

Indicator	7. a.	7. b.	7. c.	7. d.	7. e.	7. f.
Test Questions	65	27	12	9	33	14
Need Practice?						
Practice Pages	34	34	35	35	36	36

Indicator	8. a.	8. b.	8. c.	8. d.
Test Questions	24	67	22	2
Need Practice?				
Practice Pages	37	37	38	38

Name ____

Student Recording Chart *(continued)*

Indicator	9. a.	9. b.	9. c.
Test Questions	23	49	73
Need Practice?			
Practice Pages	39	39	40

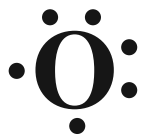
Indicator	10. a.	10. b.	10. c.	10. d.	10. e.	10. f.
Test Questions	4	36	5	42	10	21
Need Practice?						
Practice Pages	41	41	42	42	43	43

Indicator	11. a.	11. b.	11. c.	11. d.	11. e.	11. f.	11. g.
Test Questions	3	72	8	35	61	40	34
Need Practice?							
Practice Pages	44	44	45	45	46	46	47

Diagnostic Test

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

1. The Lewis dot structure for oxygen is shown below. What is the electron configuration for oxygen? **2. e.**



A. $1s^2 2s^2$
B. $1s^2 2s^2 2p^2$
C. $1s^2 2s^2 2p^4$
D. $1s^2 2s^2 2p^6 3s^2$

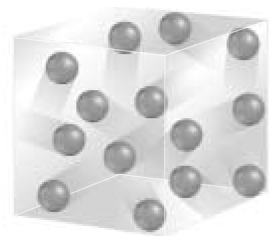
2. What is activation energy? **8. d.**
A. electrical energy required to initiate a chemical reaction
B. kinetic energy required to initiate a chemical reaction
C. potential energy required to initiate a chemical reaction
D. nuclear energy required to initiate a chemical reaction

3. Which is one of the forces that holds an atom together? **11. a.**
A. magnetic force between protons and neutrons
B. electrical force between electrons and protons
C. gravitational force between protons and neutrons
D. gravitational force between the nucleus and electrons

4. Starch is composed of many molecules of **10. a.**
A. aldehydes.
B. amino acids.
C. ketones.
D. monosaccharides.

5. Proteins are made from strings of **10. c.**
A. amino acids.
B. carboxyl groups.
C. esters.
D. nucleic acids.

6. If the size of this container is reduced by half, the pressure will **4. c.**



A. decrease.
B. increase.
C. remain the same.
D. remain the same but cause the temperature to increase.

7. The pressure on the inside walls of a sealed jar is caused by the **4. a.**
A. attraction of the gas molecules inside it on the wall of the jar.
B. collision of the gas molecules inside it against the wall of the jar.
C. temperature of the gas molecules inside it on the wall of the jar.
D. volume of the gas molecules inside it against the wall of the jar.

8. Which is most likely to be radioactive? **11. c.**
A. isotope formed in a nuclear reaction
B. isotope of carbon
C. metal formed by ores
D. salt formed from potassium

9. How much heat in joules does it take to raise 100 g of water from 25°C to 35°C? (The specific heat of water is 4.180 J/g•°C.) **7. d.**

$$q = mc\Delta T$$

A. 41.8 J
B. 418 J
C. 4180 J
D. 41,800 J

Go on

10. An alcohol has a functional group of **10. e.**
- **A.** –CHO.
- **B.** –CO.
- **C.** –H_2O.
- **D.** –OH.

11. Where is most of the mass of an atom located? **1. e.**
- **A.** electrons
- **B.** nucleus
- **C.** neutrons
- **D.** protons

12. When a material freezes, energy **7. c.**
- **A.** is absorbed by the material.
- **B.** is released by the material.
- **C.** is absorbed by a catalyst.
- **D.** remains unchanged.

13. What causes spectral lines of an element? **1. j.**
- **A.** release of protons from the nucleus
- **B.** varying number of neutrons in isotopes
- **C.** transitions of electrons between energy levels
- **D.** discrepancy between atomic number and atomic mass

14. A reaction will be spontaneous if it has a **7. f.**
- **A.** positive change in molarity.
- **B.** negative change in ionic charge.
- **C.** positive change in enthalpy.
- **D.** negative change in free energy.

15. In a solution, molecules are distributed uniformly by **6. b.**
- **A.** dissociation.
- **B.** ionic attraction.
- **C.** osmosis.
- **D.** random motion.

16. Which substance has weak intermolecular forces? **2.d**
- **A.** aluminum
- **B.** salt (NaCl)
- **C.** tin
- **D.** water

17. An acidic solution usually tastes **5. a.**
- **A.** bitter.
- **B.** salty.
- **C.** sour.
- **D.** sweet.

18. A substance's malleability is directly related to the kinds of bonds within and between the substance's molecules. Which chemical bond makes a substance the most malleable? **2. a.**
- **A.** covalent
- **B.** ionic
- **C.** metallic
- **D.** nucleic

19. Refer to the periodic table on page 56. As one moves down the elements in the first column of the periodic table, the **1. a.**
- **A.** atomic number of the elements increases.
- **B.** reactivity of the elements increases.
- **C.** atomic mass of the elements decreases.
- **D.** number of electrons available for bonding decreases.

20. A buffer helps to **5. g.**
- **A.** dissociate an acid.
- **B.** dissociate a base.
- **C.** neutralize a salt.
- **D.** stabilize pH.

21. Which part of an amino acid distinguishes it from other amino acids? **10. f.**
- **A.** CH
- **B.** COO^-
- **C.** H
- **D.** R group

22. Which can increase reaction rates? **8. c.**
- **A.** acid
- **B.** base
- **C.** catalyst
- **D.** salt

Go on

23. According to Le Châtelier's principle, what will happen if the reactants are increased in a reaction at equilibrium? **9. a.**
- **A.** There will be a net reaction to the left and no equilibrium.
- **B.** There will be a net reaction to the left until a new equilibrium occurs.
- **C.** There will be a net reaction to the right and no equilibrium.
- **D.** There will be a net reaction to the right until a new equilibrium occurs.

24. The rate of a reaction can be described using **8. a.**
- **A.** the increase in the concentration of the reactants with time.
- **B.** the increase in the concentration of the products with time.
- **C.** evidence of product formation.
- **D.** the change of enthalpy of the reaction.

25. What is the boiling point of water in kelvins? **4. e.**
- **A.** 32 K
- **B.** 100 K
- **C.** 212 K
- **D.** 373 K

26. What kind of bond exists between the two nitrogen bonds in N_2? **2. b.**
- **A.** covalent
- **B.** ionic
- **C.** metallic
- **D.** nuclear

27. Which is always true of an endothermic reaction? **7. b.**
- **A.** Chemical bonds are broken.
- **B.** Energy is released.
- **C.** Light is created.
- **D.** Energy is absorbed.

28. Substance X has a pH of 3, and substance Y has a pH of 6. This means that **5. d.**
- **A.** Substance X has 1000 times the concentration of hydrogen ions that Substance Y has, making it a stronger acid.
- **B.** Substance X has twice the concentration of hydrogen ions that Substance Y has, making it a weaker acid.
- **C.** Substance Y has 1000 times the concentration of hydrogen ions that Substance X has, making it a stronger base.
- **D.** Substance Y has twice the concentration of hydrogen ions that Substance X has, making it a weaker base.

29. A gas diffuses because the molecules of a gas **4. b.**
- **A.** have little motion.
- **B.** have random motion.
- **C.** vibrate slowly.
- **D.** vibrate quickly.

30. What is the molarity of 80 g of NaOH (40 g/mol) in 500 mL of solution? **6. d.**
- **A.** 2*M* solution
- **B.** 4*M* solution
- **C.** 8*M* solution
- **D.** 40*M* solution

31. In an oxidation reaction, a substance **3. g.**
- **A.** gains electrons.
- **B.** loses electrons.
- **C.** gains oxygen.
- **D.** loses oxygen.

32. A collection of 6.02×10^{23} is called a **3. c.**
- **A.** dalton.
- **B.** kelvin.
- **C.** mole.
- **D.** yield.

Go on

33. How do you figure the enthalpy change for the reaction shown below? **7. e.**

$$2CO\ (g) + O_2\ (g) \rightarrow 2CO_2\ (g)$$

Standard enthalpies of formation:

CO (g) = –110.5 kJ/mol
O_2 (g) = 0 kJ/mol
CO_2 (g) = –393.5 kJ/mol

A. 2(–393.5 kJ/mol) – [2(–110.5 kJ/mol) + 0 kJ/mol]
B. 2(–393.5 kJ/mol) + [2(–110.5 kJ/mol) – 0 kJ/mol]
C. – [2(–110.5 kJ/mol) + 0 kJ/mol] – 2(–393.5 kJ/mol)
D. [2(–110.5 kJ/mol) – 0 kJ/mol] + 2(–393.5 kJ/mol)

34. Which comprises the substructure of a proton? **11. g.**

A. electron
B. isotope
C. neutron
D. quark

35. Which is not a common form of radioactive decay? **11. d.**

A. alpha
B. beta
C. delta
D. gamma

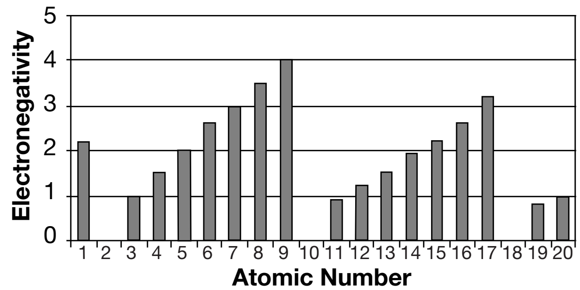
36. The most stable structure of glucose ($C_6H_{12}O_6$) contains a(n) **10. b.**

A. cyclic ring.
B. open chain.
C. double bond.
D. triple bond.

37. A scientist expects a yield of 100.0 g of a substance, based on the amount of reactants. She produces 98.0 g of substance. What is her percent yield? **3. f.**

A. 2.00%
B. 98.0%
C. 102%
D. 198%

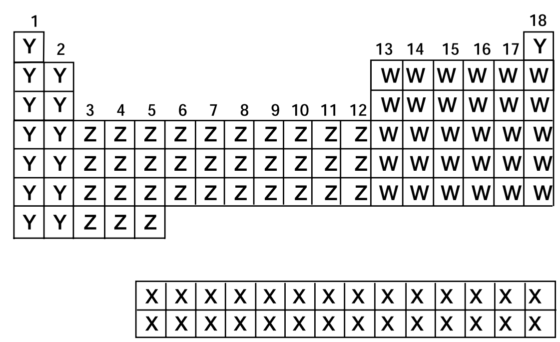
38. According to the graph, which element has the strongest attraction (highest electronegativity) for electrons? **1. c.**



A. aluminum (atomic number = 13)
B. boron (atomic number = 5)
C. oxygen (atomic number = 8)
D. sulfur (atomic number = 16)

39. The periodic table is organized into blocks representing the energy sublevel that is being filled with valence electrons. In the periodic table, which block represents the s sublevel? **1. g.**

Periodic Table



A. W
B. X
C. Y
D. Z

40. After three half-lives have passed, how much of the original radioactive material will remain? **11. f.**

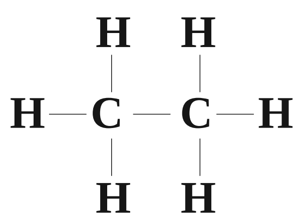
A. one-third
B. one-fourth
C. one-eighth
D. one-sixteenth

Go on

41. How will adding salt to water affect the freezing point of water? **6. e.**

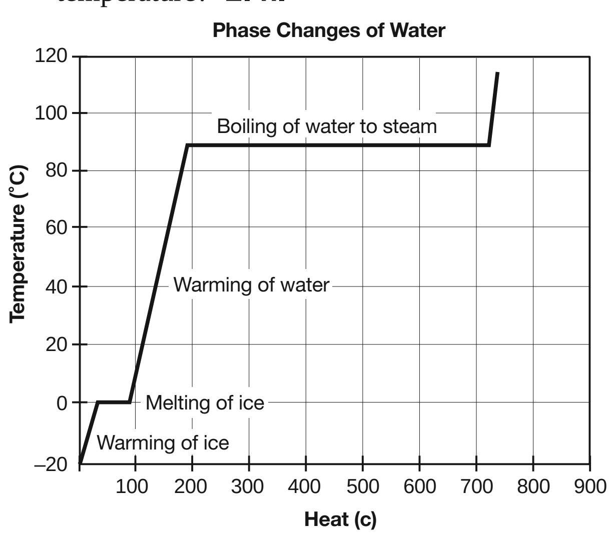
A. It will lower the freezing point.
B. It will raise the freezing point.
C. It will not affect the freezing point.
D. Water will not be able to freeze.

42. The compound shown below is known as **10. d.**



A. ethene.
B. hexene.
C. methane.
D. propane.

43. A student researched an experiment about phase changes in water and found the following graph. Which property of water contributes most to its boiling-point temperature? **2. h.**



A. cohesion
B. hydrogen bonds
C. pH
D. role in formation of other molecules

44. Coffee contains $1.0 \times 10^{-5}M$ of hydrogen ions. What is its pH? **5. f.**

A. 1.0
B. 4.0
C. 5.0
D. 9.0

45. Balance the following equation. In this equation, the question mark should be replaced by the number **3. a.**

$$2KClO_3 \rightarrow 2KCl + ?O_2$$

A. 1.
B. 2.
C. 3.
D. 6.

46. In distillation, a solute is separated from a solvent by **6. f.**

A. boiling off the solvent.
B. siphoning off the solvent.
C. using magnetic attraction to separate it from the solvent.
D. using molecular weight to separate it from the solvent.

47. Atoms in salt (NaCl) are arranged **2. c.**

A. as dissociated ions.
B. as polar molecules.
C. in a crystal lattice.
D. in random patterns.

48. How much mass does 1 mol of carbon-12 atoms have? **3. b.**

A. 12 g
B. 23 g
C. 24 g
D. 100 g

Go on

Diagnostic Test *(continued)*

49. How can someone tell when equilibrium in a reaction has occurred? **9. b.**

A. The forward reaction rate increases, and the reverse reaction rate decreases.
B. The forward reaction rate decreases, and the reverse reaction rate increases.
C. The forward and reverse reaction rates fluctuate back and forth.
D. The forward and reverse reaction rates are the same.

50. Use the ideal gas law below. What is the pressure produced by 0.5 mol O_2 in a 22.4-L container at 273 K? Use R = 0.0821 (L•atm)/(mol•K). **4. h.**

$$PV = nRT$$

A. 0.5 atm
B. 1.0 atm
C. 2.0 atm
D. 2.5 atm

51. Which instrument was important for the discovery of transuranium elements? **1. f.**

A. centrifuge
B. electron microscope
C. Geiger counter
D. nuclear accelerator

52. In the periodic table, nonmetals are located **1.b**

A. on the left side.
B. on the right side.
C. in the middle.
D. in the last two rows.

53. Which will fully dissociate in water? **5. c.**

A. strong acid
B. neutral solution
C. weak acid
D. weak base

54. Which is true of a solute? **6. a.**

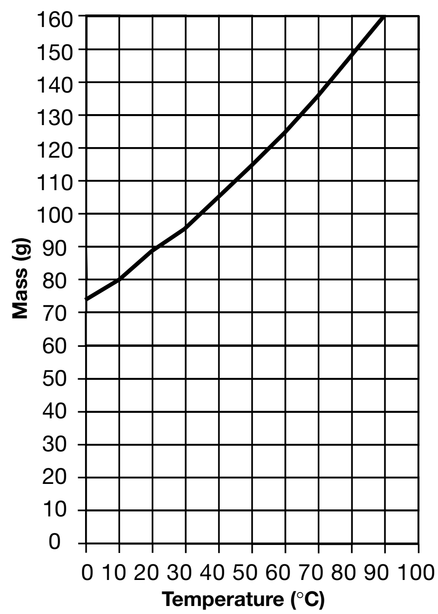
A. It contains a dissolved substance.
B. It does not react with the solvent.
C. It has a pH of less than 7.
D. It forms a salt.

55. Standard temperature and pressure (STP) occurs at **4. d.**

A. 0°C.
B. 0 K.
C. 32°F.
D. 100°F.

56. According to these data, what is the approximate number of grams of sodium nitrate that can be dissolved at a temperature of 50°C? **6. c.**

Mass of Sodium Nitrate that can be Dissolved in 100 mL of Water



A. 95 g
B. 105 g
C. 115 g
D. 125 g

57. A tank contains helium gas with a partial pressure of 2.5 atm, nitrogen gas with a partial pressure of 1.2 atm, and oxygen gas with a partial pressure of 5.7 atm. What is the total pressure in the tank? **4. i.**

A. 1.3 atm
B. 2.0 atm
C. 3.2 atm
D. 9.4 atm

Go on

58. Ernest Rutherford thought that most of the mass of an atom is contained in the nucleus. He based this theory on an experiment during which positively-charged particles **1. h.**

A. passed unchanged through gold foil.
B. changed direction as they passed through gold foil.
C. were stopped by gold foil.
D. bounced back after striking gold foil.

59. The Lewis dot structure for N_2 is shown below. What is the probable shape of N_2? **2. f.**

$$:N\equiv N:$$

A. linear
B. octahedral
C. tetrahedral
D. trigonal planar

60. Refer to the periodic table on page 56 of this booklet. How many electrons do the elements in the last column of the periodic table have available for bonding? **1. d.**

A. 0
B. 2
C. 8
D. 18

61. Alpha particles **11. e.**

A. have a charge of 1.
B. have a mass of 2.
C. are not deflected in an electric field.
D. can be stopped by a piece of paper.

62. How many particles does 16 g of sulfur (atomic weight = 32) have? **3. d.**

A. 3.01×10^{23}
B. 6.02×10^{23}
C. 1.20×10^{24}
D. 6.02×10^{24}

63. When a halogen and an alkali metal bond, a(n) **2. g.**

A. covalent bond is formed because the difference in electronegativity is small.
B. covalent bond is formed because the difference in electronegativity is great.
C. ionic bond is formed because the difference in electronegativity is small.
D. ionic bond is formed because the difference in electronegativity is great.

64. Which temperature represents absolute zero? **4. f.**

A. 0°C
B. 100°C
C. 0 K
D. 273 K

65. When molecules of a high temperature collide with molecules of a low temperature, **7. a.**

A. chemical energy is transferred.
B. kinetic energy is transferred.
C. nuclear energy is transferred.
D. potential energy is transferred.

66. The diagram shows a chemical equation representing a chemical reaction. The name and mass of each substance involved in the chemical reaction are also shown. What mass of water was produced in this reaction? **3. e.**

36.5 g		40.0 g		58.5 g		?
HCl	+	**NaOH**	→	**NaCl**	+	**H_2O**
hydrochloric acid		sodium hydroxide		sodium chloride		water

A. 16.0 g
B. 18.0 g
C. 20.0 g
D. 22.0 g

67. Which will increase reaction rates? **8. b.**

A. high concentration of products
B. low temperature
C. large surface area of reactants
D. low mass of products

Go on

68. Bases accept **5. b.**
- **A.** anhydrous compounds.
- **B.** ionic compounds.
- **C.** hydrogen ions.
- **D.** hydroxide ions.

69. As the temperature of a gas increases, the molecules of the gas **4. g.**
- **A.** increase in size.
- **B.** lose energy.
- **C.** move faster.
- **D.** repel each other.

70. Niels Bohr proposed that an atom contains **1. i.**
- **A.** a fixed number of neutrons.
- **B.** electrons in specific energy levels.
- **C.** protons scattered throughout the atom.
- **D.** electrons scattered throughout the atom.

71. A Brønsted-Lowry base is to a hydrogen-ion acceptor as a Brønsted-Lowry acid is to a(n) **5. e.**
- **A.** hydroxide-ion producer.
- **B.** hydroxide-ion donor.
- **C.** electron-pair donor.
- **D.** hydrogen-ion donor.

72. Which type of reaction has the largest release of energy per gram of material? **11. b.**
- **A.** breakdown of steak into protein
- **B.** fusion of hydrogen into helium
- **C.** formation of glycogen from glucose
- **D.** formation of salt from sodium and chlorine

73. Which is the correct way to express the equilibrium constant for the reaction below? **9. c.**

$$2NO_2\ (g) \rightleftharpoons 2NO\ (g) + O_2\ (g)$$

- **A.** $K_{eq} = [N_2O_4]$
- **B.** $K_{eq} = [NO]^2[O_2]^2$
- **C.** $K_{eq} = \dfrac{[NO][O_2]^2}{[NO_2]}$
- **D.** $K_{eq} = \dfrac{[NO]^2\,[O_2]}{[NO_2]^2}$

Stop

Standards Practice

Atomic and Molecular Structure

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

1. a. **Students know how to relate the position of an element in the periodic table to its atomic number and atomic mass.**

Use the periodic table on page 56 to answer questions 1–5.

1. The atomic number in the periodic table is highest in the
- **A.** upper left corner.
- **B.** upper right corner.
- **C.** lower left corner.
- **D.** lower right corner.

2. The atomic mass in the periodic table is highest in the
- **A.** upper left corner.
- **B.** upper right corner.
- **C.** lower left corner.
- **D.** lower right corner.

3. When moving across a period of the periodic table, the
- **A.** atomic masses of the elements decrease.
- **B.** atomic numbers of the elements increase.
- **C.** reactivity of the elements decreases.
- **D.** number of electrons available to be given up for bonding increases.

4. Which element has the highest atomic mass?
- **A.** copper (atomic number 29)
- **B.** gold (atomic number 79)
- **C.** platinum (atomic number 78)
- **D.** silver (atomic number 47)

5. As one moves across a period of the periodic table, the
- **A.** atomic masses of the elements increase.
- **B.** atomic numbers of the elements decrease.
- **C.** the electronegativity of the elements decreases.
- **D.** the number of valence electrons decreases.

1. b. **Students know how to use the periodic table to identify metals, semimetals, nonmetals, and halogens.**

Use the periodic table on page 56 to answer questions 6–9.

6. In the periodic table, metals are best described as located
- **A.** on the left side.
- **B.** on the right side.
- **D.** in the first two rows.
- **C.** in the last two columns.

7. In the periodic table, which column (group) contains the most nonmetals?
- **A.** 1
- **B.** 2
- **C.** 17
- **D.** 18

8. In the periodic table, which column (group) contains the most halogens?
- **A.** 1A
- **B.** 2A
- **C.** 7A
- **D.** 8A

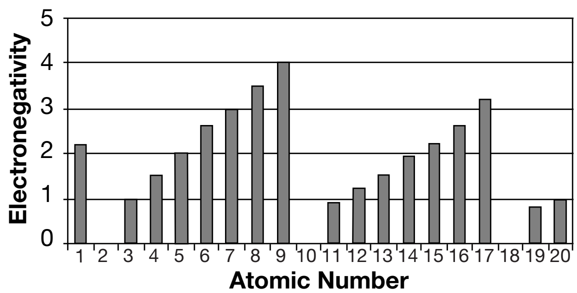
9. Most elements located against the diagonal line in the periodic table are
- **A.** halogens.
- **B.** metals.
- **C.** nonmetals.
- **D.** semimetals (metalloids).

Standards Practice

Atomic and Molecular Structure

1. c. **Students know how to use the periodic table to identify alkali metals, alkaline earth metals and transition metals, trends in ionization energy, electronegativity, and the relative sizes of ions and atoms.**

Use the graph below to answers questions 10 and 11.



10. The electronegativity of an element indicates the relative ability of its atoms to attract electrons in a chemical bond. According to the graph, as you move across a period in the periodic table, the atomic number

A. increases and the electronegativity increases.
B. increases and the electronegativity decreases.
C. decreases and the electronegativity increases.
D. decreases and the electronegativity decreases.

11. According to the graph, which of these elements has the strongest attraction for electrons?

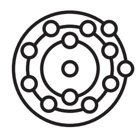
A. Boron (atomic number = 5)
B. Calcium (atomic number = 20)
C. Hydrogen (atomic number = 1)
D. Sulfur (atomic number = 16)

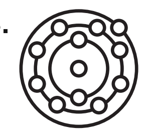
12. The ionization energy of an element indicates the energy it takes to remove an electron from the outermost level. As you move down a group in the periodic table, the ionization energy

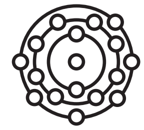
A. decreases because the electron is farther from the nucleus.
B. increases because the number of protons increases.
C. decreases because the number of neutrons increases.
D. increases because the size of the atom increases.

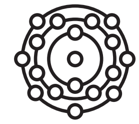
1. d. **Students know how to use the periodic table to determine the number of electrons available for bonding.**

13. Alkali metals belong to a group of elements whose atoms have only one electron in their outer energy level. According to this definition, which of these is an atom of an alkali metal?

A.


B.


C.


D.


14. How many electrons does the element found in the second row and second column of the periodic table have available for bonding?

A. 0
B. 1
C. 2
D. 4

15. How many electrons does the element found in row 2 and group 16 of the periodic table have in the outer shell?

A. 0
B. 6
C. 8
D. 12

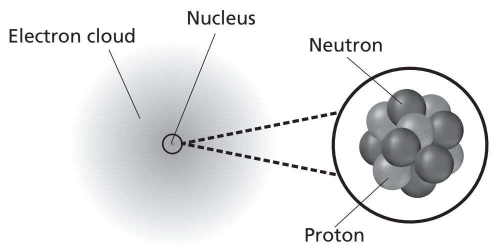
16. Why are the elements in column 17 of the periodic table particularly reactive?

A. They have only one electron in their outer electron shells, so they frequently form singly charged positive ions.
B. They have only two electrons in their outer shells, so they frequently bond with doubly charged negative ions.
C. They need only two electrons to complete their outer shells, so they easily bond with doubly charged positive ions.
D. They need only one electron to complete their outer shells, so they easily bond with many types of elements.

Standards Practice
Atomic and Molecular Structure

1. e. **Students know the nucleus of the atom is much smaller than the atom yet contains most of its mass.**

Use the illustration below to answer question 17.



17. How does the nucleus compare to the entire atom?
- **A.** It is slightly smaller than the atom.
- **B.** It is the only part of the atom that has a charge.
- **C.** It contains most of the atom's mass.
- **D.** It contains all of the atom's mass.

18. Which part of the atom has the least amount of mass?
- **A.** first s orbital
- **B.** neutrons
- **C.** nucleus
- **D.** protons

19. When a hydrogen atom loses its electron, how do its density and weight change?
- **A.** Its density and weight both decrease slightly.
- **B.** Its density decreases significantly, and its weight decreases slightly.
- **C.** Its density and weight both decrease significantly.
- **D.** Its density increases significantly, and its weight decreases slightly.

20. An atom
- **A.** is much larger than its nucleus.
- **B.** is much lighter than its nucleus.
- **C.** is much denser than its nucleus.
- **D.** has a higher positive charge than its nucleus.

1. f. **Students know how to use the periodic table to identify the lanthanide, actinide, and transactinide elements and know that the transuranium elements were synthesized and identified in laboratory experiments through the use of nuclear accelerators.**

Use the periodic table on page 56 to answer questions 21–24.

21. The lanthanide and actinide elements are located together on the periodic table because they utilize the
- **A.** s orbital.
- **B.** p orbital.
- **C.** d orbital.
- **D.** f orbital.

22. Where are the lanthanide and actinide elements located on the periodic table?
- **A.** the first row
- **B.** the first two rows
- **C.** two rows normally shown as an insert
- **D.** the middle columns

23. Transuranium elements
- **A.** have outer electrons only in the s orbital.
- **B.** have an atomic number greater than 92.
- **C.** are inert.
- **D.** are nonmetals.

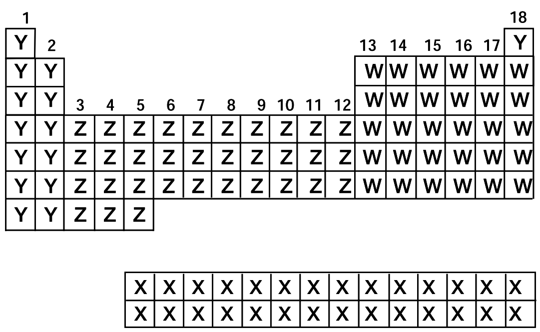
24. Which elements are not found naturally?
- **A.** actinide elements
- **B.** lanthanide elements
- **C.** transuranium elements
- **D.** uranium elements

Standards Practice
Atomic and Molecular Structure

1. g. **Students know how to relate the position of an element in the periodic table to its quantum electron configuration and to its reactivity with other elements in the table.**

Use the periodic table shown below to answer questions 25 and 26.

Periodic Table



25. The periodic table is organized into blocks representing the energy sublevel being filled with valence electrons. In the periodic table, which sequence lists the blocks in s-p-d-f order?

A. W, Y, X, Z
B. X, Y, Z, W
C. Y, W, Z, X
D. Y, Z, W, X

26. In the periodic table, which block represents the d sublevel?

A. W
B. X
C. Y
D. Z

27. An atom with the electron configuration $1s^22s^22p^63s^23p^64s^2$ is most likely

A. a metal that forms a positive ion.
B. a metal that forms a negative ion.
C. a nonmetal that forms a positive ion.
D. a nonmetal that forms a negative ion.

28. An atom with the electron configuration $1s^2\ 2s^2\ 2p^4$ is most likely

A. a metal that forms a positive ion.
B. a metal that forms a negative ion.
C. a nonmetal that forms a positive ion.
D. a nonmetal that forms a negative ion.

1. h. **Students know the experimental basis for Thomson's discovery of the electron, Rutherford's nuclear atom, Millikan's oil drop experiment, and Einstein's explanation of the photoelectric effect.**

29. By measuring the effects of magnetic and electrical fields on cathode rays, J.J. Thomson discovered that these particles

A. had no mass.
B. were heavier than hydrogen atoms.
C. were smaller than hydrogen atoms.
D. were easily broken apart.

30. Robert Millikan dropped negatively charged oil droplets inside and outside an electric field to see

A. if oil drops fell faster than electrons.
B. if oil drops fell faster than protons.
C. how their rates of falling would be affected by a positively charged plate.
D. if negatively charged particles were attracted to a positively charged plate.

31. Ernest Rutherford's gold foil experiment showed that some alpha particles beamed at a thin sheet of gold foil were severely deflected, contrary to his expectations. What caused the deflection?

A. an electrical field
B. magnetism
C. an inner nucleus
D. electron shells

32. Albert Einstein proposed that electromagnetic energy is contained in

A. electrons.
B. photons.
C. protons.
D. quarks.

Standards Practice

Atomic and Molecular Structure

1. i. **Students know the experimental basis for the development of the quantum theory of atomic structure and the historical importance of the Bohr model of the atom.**

33. The Bohr model of the atom contained
- **A.** neutrons and electrons in the nucleus.
- **B.** protons and electrons in the nucleus.
- **C.** electrons scattered throughout the nucleus.
- **D.** electrons contained in energy levels.

34. The Bohr model of the atom was also known as the
- **A.** balloon model.
- **B.** igneous rock model.
- **C.** planetary model.
- **D.** plum pudding model.

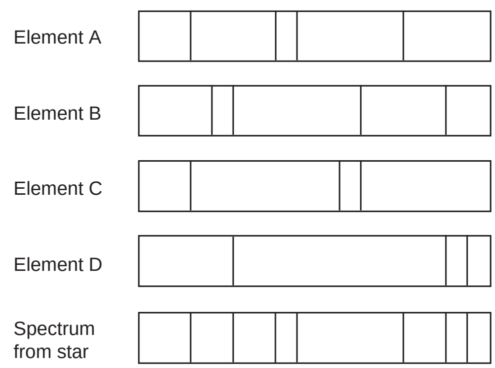
35. The Bohr model was probably most influenced by the work of
- **A.** Max Born.
- **B.** Robert Millikan.
- **C.** Max Planck.
- **D.** J.J. Thomson.

36. Niels Bohr's model of the atom helped to explain
- **A.** particles passing through foil.
- **B.** spectral lines.
- **C.** the formation of isotopes.
- **D.** the properties of hydrogen.

37. Niels Bohr proposed that atoms
- **A.** were put together like plum pudding.
- **B.** had a negatively charged nucleus.
- **C.** had protons scattered throughout the atom.
- **D.** contained electrons in specific energy levels.

1. j. **Students know that spectral lines are the result of transitions of electrons between energy levels and that these lines correspond to photons with a frequency related to the energy spacing between levels by using Planck's relationship ($E = hv$).**

38. The figure below shows the bright line spectra of four elements and a spectrum from a star. According to the spectrum, which elements are present in this star?



- **A.** elements A and B
- **B.** elements B and C
- **C.** elements A and C
- **D.** elements B, C, and D

39. Planck's constant (h) equals 6.626×10^{-34} J•s. According to Albert Einstein, $E_{photon} = hv$. What is the energy of a photon if it has a frequency (v) of 6.82×10^{14} s^{-1}?
- **A.** 4.52×10^{-19}
- **B.** 1.03×10^{-20}
- **C.** 4.52×10^{-20}
- **D.** 9.72×10^{-20}

40. Based on the equation $E = hv$, which is not a possible value for an energy change?
- **A.** hv
- **C.** $2hv$
- **D.** 2^2hv
- **B.** $6.626hv$

Standards Practice
Chemical Bonds

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

2. a. **Students know that atoms combine to form molecules by sharing electrons to form covalent or metallic bonds or by exchanging electrons to form ionic bonds.**

1. Potassium (K) and chlorine (Cl) form a(n)
A. covalent bond.
B. hydrogen bond.
C. ionic bond.
D. metallic bond.

2. When atoms combine to form a molecule by sharing electrons, what type of bonds are formed?
A. covalent
B. hydrogen
C. ionic
D. polar ionic

3. Which is the best way to express the relationship between hydrogen and fluorine when they combine?

A. H – F

B. δ^{+} H – Fδ^{-}

C. H : F

D. H:F̤̈:

4. A metallic bond is formed between
A. a metal atom and a hydrogen atom.
B. a metal atom and a nonmetallic atom.
C. a metal atom and a noble gas.
D. two metal atoms.

2. b. **Students know chemical bonds between atoms in molecules such as H_2, CH_4, NH_3, H_2CCH_2, N_2, Cl_2, and many large biological molecules are covalent.**

5. Which do not form covalent bonds?
A. diatomic molecules
B. large biological molecules
C. molecules containing carbon
D. salts

6. The bonds found in C_2H_4 are
A. covalent.
B. ionic.
C. metallic.
D. polar.

7. Which is a covalent compound?
A. $AlBr_3$
B. CO_2
C. KCl
D. NaF

8. Which is a covalent compound?
A. Mg_3N_2
B. NaCl
C. NaF
D. SiF_4

9. Which is not a covalent compound?
A. CCl_4
B. H_2
C. $MgCl_2$
D. SO_3

Standards Practice
Chemical Bonds

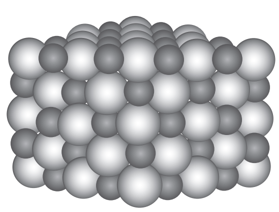
2. c. **Students know salt crystals, such as NaCl, are repeating patterns of positive and negative ions held together by electrostatic attraction.**

10. Salts are compounds that form a crystal lattice. Which types of bonds are responsible for this lattice formation?
- **A.** covalent
- **B.** hydrogen
- **C.** ionic
- **D.** polar

11. When atoms of sodium (Na) and chlorine (Cl) combine to form salt (NaCl), the Na+ ion is smaller than the Na atom, while the Cl– ion is larger than the Cl atom. Why?
- **A.** The Na and Cl atoms both lost electrons.
- **B.** The Na and Cl atoms both gained electrons.
- **C.** The Na atom lost an electron, while the Cl atom gained an electron.
- **D.** The Na atom gained an electron, while the Cl atom lost an electron.

12. The electrostatic attraction between atoms in a salt is
- **A.** strongest when the ions are small.
- **B.** weakest when one of the ions is hydrogen (atomic number 1).
- **C.** strongest when one of the ions is potassium (atomic number 19).
- **D.** strongest when one of the ions is iodine (atomic number 53).

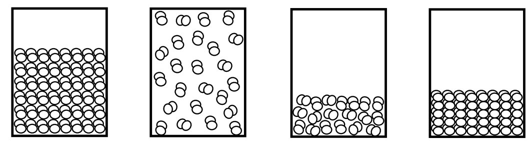
13. Why does sodium chloride form a lattice?



- **A.** Sodium as a positive ion and chlorine as a positive ion are held in this position.
- **B.** Sodium as a negative ion and chlorine as a negative ion are held in this position.
- **C.** Sodium as a positive ion and chlorine as a negative ion are held in this position.
- **D.** Sodium as a negative ion and chlorine as a positive ion are held in this position.

2. d. **Students know the atoms and molecules in liquids move in a random pattern relative to one another because the intermolecular forces are too weak to hold the atoms or molecules in a solid form.**

14. Which illustration most likely represents a liquid?



Q R S T

- **A.** Q
- **B.** R
- **C.** S
- **D.** T

15. Why do liquids take on the shapes of the containers that hold them?
- **A.** Their molecules are held together by strong intermolecular forces.
- **B.** Their molecules are organized in the form of a lattice.
- **C.** Their molecules move in a random pattern.
- **D.** Their molecules move quickly and collide frequently with each other.

16. At room temperature, which substance has the weakest intermolecular forces?
- **A.** oxygen
- **B.** salt
- **C.** steel
- **D.** uranium

17. At room temperature, which substance has the strongest intermolecular forces?
- **A.** water
- **B.** zinc
- **C.** 1*M* solution HCl
- **D.** 1*M* solution NaOH

Standards Practice

Chemical Bonds

2. e. **Students know how to draw Lewis dot structures.**

18. Why are Lewis dot structures used?
- **A.** to show what type of bonds are formed
- **B.** to show valence electrons
- **C.** to show electronegativity values
- **D.** to show intermolecular forces

19. Argon has an electron configuration of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. How many electrons are shown in its Lewis dot structure?
- **A.** 0
- **B.** 2
- **C.** 6
- **D.** 8

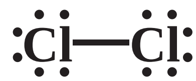
20. Fluorine has seven electrons in its Lewis dot structure. What is the electron configuration for fluorine?
- **A.** $1s^2 2s^2$
- **B.** $1s^2 2s^2 2p^3$
- **C.** $1s^2 2s^2 2p^4$
- **D.** $1s^2 2s^2 2p^5$

21. Magnesium has two electrons in its Lewis dot structure. What is the electron configuration for magnesium?
- **A.** $1s^2\ 2s^2\ 2p^3$
- **B.** $1s^2\ 2s^2\ 2p^6$
- **C.** $1s^2\ 2s^2\ 2p^6\ 3s^2$
- **D.** $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

22. The electron configuration for an atom of iron is [Ar] $3d^6 4s^2$. Which is the correct Lewis dot structure for iron?
- **A.** Fe•
- **B.** Fe:
- **C.** 
- **D.** 

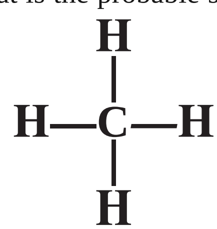
2. f. **Students know how to predict the shape of simple molecules and their polarity from Lewis dot structures.**

23. The Lewis dot structure for Cl_2 is shown below. What is the probable shape of Cl_2?



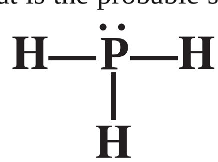
- **A.** linear
- **B.** pyramidal
- **C.** tetrahedral
- **D.** trigonal planar

24. The Lewis dot structure for CH_4 is shown below. What is the probable shape of CH_4?



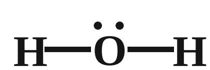
- **A.** bent
- **B.** linear
- **C.** pyramidal
- **D.** tetrahedral

25. The Lewis dot structure for PH_3 is shown below. What is the probable shape of PH_3?



- **A.** linear
- **B.** pyramidal
- **C.** tetrahedral
- **D.** trigonal planar

26. The Lewis dot structure for H_2O is shown below. What is the probable shape and polarity of H_2O?



- **A.** bent and nonpolar
- **B.** bent and polar
- **C.** linear and nonpolar
- **D.** linear and polar

Standards Practice
Chemical Bonds

2. g. **Students know how electronegativity and ionization energy relate to bond formation.**

27. The bond that holds two fluorine atoms together in an F_2 molecule is classified as nonpolar covalent because the difference in electronegativity is
- **A.** large.
- **B.** small.
- **C.** zero.
- **D.** indiscernible.

28. When two metal atoms bond,
- **A.** a covalent bond is formed because the difference in electronegativity is small.
- **B.** a covalent bond is formed because the difference in electronegativity is great.
- **C.** an ionic bond is formed because the difference in electronegativity is small.
- **D.** an ionic bond is formed because the difference in electronegativity is great.

29. What situation results in a polar covalent bond?
- **A.** two nonmetals bonding, with a small difference in electronegativity
- **B.** two nonmetals bonding, with a large difference in electronegativity
- **C.** two metals bonding, with a small difference in electronegativity
- **D.** a metal and a nonmetal bonding, with a large difference in electronegativity

30. The table below shows the lattice energy for some ionic compounds. Based on these data, which of these compounds would require the most energy to separate the bonded ions?

Compound	Lattice Energy (kJ/mol)
NaCl	-769
KBr	-671
LiF	-1030
MgO	-3795

- **A.** NaCl
- **B.** KBr
- **C.** LiF
- **D.** MgO

2. h. **Students know how to identify solids and liquids held together by van der Waals forces or hydrogen bonding and relate these forces to volatility and boiling/melting point temperatures.**

31. Which type of attraction is weakest?
- **A.** covalent bond
- **B.** hydrogen bond
- **C.** ionic bond
- **D.** van der Waals forces

32. Compared to other group 6A hydrides, the hydrogen bonds in water keep boiling-point temperature
- **A.** lower than expected because of the hydrogen atoms.
- **B.** lower than expected because of the strength of the bonds.
- **C.** higher than expected because of the strength of the bonds.
- **D.** higher than expected because of the greater length of the bonds.

33. How do van der Waals forces affect the boiling point of helium?
- **A.** They are the reason that helium is a liquid at room temperature.
- **B.** They cause helium to boil at a very low temperature.
- **C.** They keep the boiling-point temperature higher than expected because of the strength of the bonds.
- **D.** Van der Waals forces do not affect the boiling point of helium.

34. High pressure helps to disrupt hydrogen bonds. How would high pressure affect the temperature at which ice will melt?
- **A.** High pressure will lower the temperature at which ice will melt.
- **B.** High pressure will raise the temperature at which ice will melt.
- **C.** Hydrogen bonds do not affect the freezing point, so it will not change the temperature at which ice will melt.
- **D.** The freezing point will remain the same because the high pressure will negate the effect of the hydrogen bonds.

Standards Practice

Conservation of Matter and Stoichiometry

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

3. a. **Students know how to describe chemical reactions by writing balanced equations.**

1. Equations can be balanced because
A. energy in equals energy out.
B. matter is neither created nor destroyed.
C. atoms break down easily.
D. molecules are virtually inseparable.

2. Which is a correct balanced chemical equation?
A. $2Zn\ (s) + HCl\ (aq) \rightarrow 2ZnCl_2 + H_2\ (g)$
B. $Zn\ (s) + 2HCl\ (aq) \rightarrow ZnCl_2 + 2H_2\ (g)$
C. $Zn\ (s) + 2HCl\ (aq) \rightarrow ZnCl_2 + H_2\ (g)$
D. $2Zn\ (s) + 2HCl\ (aq) \rightarrow 2ZnCl_2 + H_2\ (g)$

3. Which is a correct balanced chemical equation?
A. $2Al + 3CuSO_4 \rightarrow 3Cu + Al_2(SO_4)^3$
B. $6Al + 3CuSO_4 \rightarrow 3Cu + 3Al2(SO_4)^3$
C. $Al + 2CuSO_4 \rightarrow 2Cu + Al2(SO_4)^3$
D. $3Al + 3CuSO_4 \rightarrow 3Cu + Al2(SO_4)^3$

4. Balance the following equation. In this equation, ? should be replaced by

$$Mg + ?AgNO_3 \rightarrow Mg(NO_3)^2 + 2Ag$$

A. 1.
B. 2.
C. 3.
D. 4.

5. Balance the following equation. In this equation, ? should be replaced by

$$4Fe + ?O_2 \rightarrow 2Fe_2O_3$$

A. 1.
B. 2.
C. 3.
D. 4.

6. Balance the following equation. In this equation, ? should be replaced by

$$C_3H_8 + 5O_2 \rightarrow 3CO_2 + ?H_2O$$

A. 1.
B. 2.
C. 3.
D. 4.

3. b. **Students know that the quantity one mole is set by defining one mole of carbon 12 atoms to have a mass of exactly 12 grams.**

7. Twelve grams of carbon equals
A. 0.1 mol.
B. 0.5 mol.
C. 1 mol.
D. 2 mol.

8. Why is setting a standard for the quantity of 1 mol important?
A. The quantity is needed to determine the volume of a solid.
B. The quantity is needed to determine ionic composition.
C. The quantity is needed to set proportions for the conservation of mass.
D. The quantity is needed to determine a molecule's energy constant.

9. If one mole of carbon-12 has a mass of 12 g, what should be the mass of 1 mol of the isotope carbon-13 (atomic number = 6; atomic mass = 13)?
A. 6 g
B. 7 g
C. 12 g
D. 13 g

10. If the quantity of 1 mol of carbon is 12, what can be said about the quantity of 1 mol of lithium?
A. 1 mol of lithium should also be 12 g.
B. 1 mol of lithium should be 3 g (atomic number).
C. 1 mol of lithium should be 4 g (number of neutrons).
D. 1 mol of lithium should be 7 g (atomic weight).

Standards Practice

Conservation of Matter and Stoichiometry

3. c. **Students know one mole equals 6.02×10^{23} particles (atoms or molecules).**

11. How many atoms does 1 mol of carbon-12 have?
- **A.** 6.0×10^{23} molecules
- **B.** 6.02×10^{23} molecules
- **C.** 9.01×10^{23} molecules
- **D.** 12.0×10^{23} molecules

12. Which could not be determined by knowing the number of particles of a substance that are present?
- **A.** the balanced equation for a reaction
- **B.** the ionization energy of a substance
- **C.** the mass of the substance (given its chemical formula)
- **D.** the number of moles present

13. What can be said for magnesium, which has an atomic number of 12 and an atomic weight of 24.3?
- **A.** One mole of magnesium will have half the number of atoms as one mole of carbon.
- **B.** One mole of magnesium will have the same number of atoms as one mole of carbon.
- **C.** One mole of magnesium will have twice the number of atoms as one mole of carbon.
- **D.** One mole of magnesium will have four times the number of atoms as one mole of carbon.

14. How many particles are present in 1 mol of the isotope carbon-13 (atomic number = 6; atomic mass = 13)?
- **A.** 3.01×10^{23} molecules
- **B.** 6×10^{23} molecules
- **C.** 6.02×10^{23} molecules
- **D.** 1.3×10^{24} molecules

15. How many molecules does 1 mole of NaOH have?
- **A.** 3.01×10^{23} molecules
- **B.** 6.02×10^{23} molecules
- **C.** 1.20×10^{24} molecules
- **D.** 6.02×10^{24} molecules

3. d. **Students know how to determine the molar mass of a molecule from its chemical formula and a table of atomic masses and how to convert the mass of a molecular substance to moles, number of particles, or volume of gas at standard temperature and pressure.**

Use the table below to answer questions 16–18.

Element	Atomic Weight (g)
Carbon (C)	12.0
Hydrogen (H)	1.0
Iron (Fe)	55.9
Nitrogen (N)	14.0
Oxygen (O)	16.0
Sodium (Na)	23.0

16. What is the weight of 1 mol of CH_3OH?
- **A.** 15 g
- **B.** 28 g
- **C.** 29 g
- **D.** 32 g

17. What is the weight of 1 mol of $NaNO_3$?
- **A.** 53 g
- **B.** 69 g
- **C.** 85 g
- **D.** 101g

18. What is the weight of 1 mole of Fe_2O_3?
- **A.** 103.9 g
- **B.** 127.8 g
- **C.** 143.8 g
- **D.** 159.8 g

19. If 1 mol of gas has a volume of 22.4 L at standard temperature and pressure (STP), how much volume would 0.5 mol of the same gas have?
- **A.** 0.5 L
- **B.** 11.2 L
- **C.** 22.4 L
- **D.** 44.8 L

20. How many molecules do 2 mol of HCl have?
- **A.** 3.01×10^{23}
- **B.** 6.02×10^{23}
- **C.** 1.20×10^{24}
- **D.** 6.02×10^{24}

Standards Practice

Conservation of Matter and Stoichiometry

3. e. Students know how to calculate the masses of reactants and products in a chemical reaction from the mass of one of the reactants or products and the relevant atomic masses.

21. You can use a balanced chemical equation and the atomic masses of products and reactants of the chemical equation to determine a missing mass because

A. compounds always combine in the same way.
B. a chemical equation will eventually reach equilibrium.
C. matter is neither created nor destroyed.
D. a given equation will only balance in one way.

Use the following equation to answer questions 22 and 23. Hydrogen has an atomic mass of 1, and oxygen has an atomic mass of 16.

$$2H_2 + O_2 \rightarrow 2H_2O$$

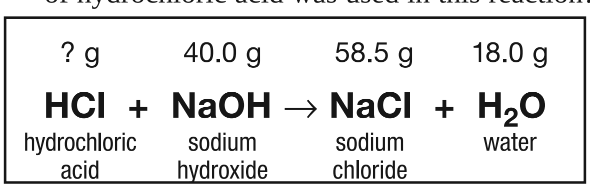
22. If 4 g of hydrogen reacts with an unlimited amount of oxygen, how many grams of water will be produced?

A. 18
B. 32
C. 36
D. 40

23. If 8 g of oxygen reacts with an unlimited amount of hydrogen, how many grams of water will be produced?

A. 9
B. 16
C. 18
D. 36

24. The diagram shows a chemical equation representing a chemical reaction. The name and mass of each substance involved in the chemical reaction are also shown. What mass of hydrochloric acid was used in this reaction?



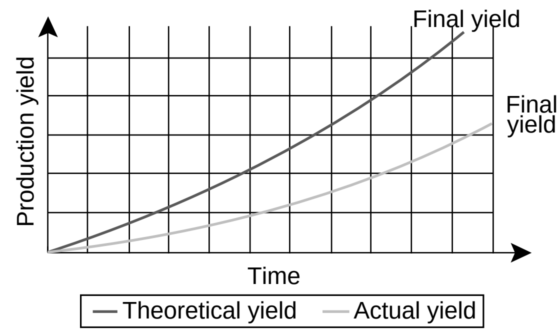
A. 24.0 g
B. 36.5 g
C. 48.0 g
D. 73.0 g

3. f. Students know how to calculate percent yield in a chemical reaction.

25. Which equation gives percent yield?

A. (actual yield ÷ theoretical yield) × 100
B. (theoretical yield ÷ actual yield) × 100
C. (actual yield × theoretical yield) × 100
D. (actual yield × theoretical yield) ÷ 100

Use the graph below to answer questions 26 and 27.



26. According to the graph, when a chemical reaction is first occurring,

A. there is little difference between actual yield and theoretical yield.
B. there is a great difference between actual yield and theoretical yield.
C. the actual yield and theoretical yield start at 100 percent.
D. the actual yield and theoretical yield approach 100 percent.

27. According to the graph, as more product is produced during a chemical reaction,

A. the actual yield will reach 100 percent.
B. the theoretical yield will reach 0 percent.
C. there will be a greater difference between actual yield and theoretical yield.
D. the actual yield and theoretical yield will become equal.

28. During a chemical reaction, the

A. actual yield is higher than the theoretical yield.
B. calculated percent yield of product is less than 100 percent.
C. theoretical yield and actual yield are equal.
D. percent yield of product is more than 100 percent.

Standards Practice

Conservation of Matter and Stoichiometry

3. g. **Students know how to identify reactions that involve oxidation and reduction and how to balance oxidation-reduction reactions.**

29. In a reduction reaction, a substance

A. gains electrons.
B. loses electrons.
C. gains oxygen.
D. loses oxygen.

30. In the oxidation reaction shown below, which letter represents the electrons?

$$A\ Pb\ B \rightarrow Pb^{2+} + C$$

A. A
B. B
C. C
D. No electrons are needed.

31. In the reduction reaction shown below, how many electrons are needed to balance the equation?

$$Zn^{2}+ + ?e^{-} \rightarrow Zn$$

A. 0
B. 1
C. 2
D. 4

32. In the oxidation reaction shown below, how many electrons are needed to balance the equation?

$$Al \rightarrow Al^{3}+ + ?e–$$

A. 1
B. 2
C. 3
D. 4

33. In the reaction shown below, which is reduced?

$$Fe + CuSO_4 \rightarrow FeSO_4 + Cu$$

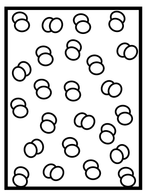
A. Cu
B. Fe
C. SO4
D. No element or compound is reduced.

Standards Practice
Gases and Their Properties

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

4. a. Students know the random motion of molecules and their collisions with a surface create the observable pressure on that surface.

1. What causes the pressure on the inside walls of this container?



A. charge of the gas molecules in the container
B. temperature of the gas molecules in the container
C. collision of the gas molecules on the container
D. weight of the gas molecules on the container

2. How do the properties of a gas differ from those of a liquid?
A. Gas molecules have a greater random motion than liquid molecules.
B. Gas molecules have less energy than liquid molecules.
C. Gas molecules have more mass than liquid molecules.
D. Gas molecules put more pressure on the walls of a container than liquid molecules.

3. What does the random motion of molecules and their collisions with a surface produce?
A. mass
B. phase change
C. pressure
D. weight

4. What causes a balloon to hold its shape?
A. random motion of gas molecules
B. collisions of gas molecules against a balloon's walls
C. weight of the gas molecules inside of the balloon
D. energy held by each gas molecule within the balloon

4. b. Students know the random motion of molecules explains the diffusion of gases.

5. Which causes the addition of a colored gas to change the color of all of the gas in a container?
A. diffusion
B. mass
C. pressure
D. temperature

6. Diffusion is the term used to describe the movement of one material through another. The diffusion of gases can be explained by
A. relative molar masses.
B. differences in volume.
C. evaporation.
D. random motion.

7. Which is an example of diffusion?
A. disappearance of a puddle in sunlight
B. smell of a rotten egg across a room
C. filled balloon that shrinks over time
D. balloon that shrinks when it becomes cold

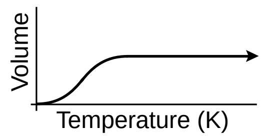
8. Which is not an example of diffusion?
A. boiling of water
B. smell of food cooking
C. colored gas moving throughout a room
D. poisonous gas leaking from an open container

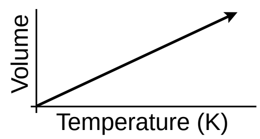
Standards Practice

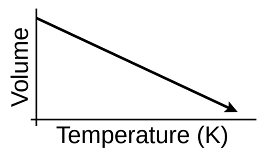
Gases and Their Properties

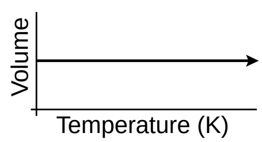
4. c. **Students know how to apply the gas laws to relations between the pressure, temperature, and volume of any amount of an ideal gas or any mixture of ideal gases.**

9. Charles's law explains the relationship between the temperature and volume of a gas. Which graph best represents this relationship?

A. 

B. 

C. 

D. 

10. How are pressure and volume of gas related?
- **A.** As pressure decreases, volume increases.
- **B.** As pressure increases, volume increases.
- **C.** As pressure decreases, volume decreases.
- **D.** Pressure and the volume are not related.

11. Which of these decreases as a given volume of gas increases?
- **A.** number of gas particles
- **B.** temperature
- **C.** pressure
- **D.** kinetic energy

12. There are two containers of two different gases at the same temperature and pressure. Each statement below can be assumed except
- **A.** when the temperature is increased, the volume of both containers will increase.
- **B.** when the pressure is increased, the volume of both containers will decrease.
- **C.** both containers contain the same number of gas particles.
- **D.** when the pressure is decreased, the temperature of both containers will increase.

4. d. **Students know the values and meanings of standard temperature and pressure (STP).**

13. Standard temperature and pressure (STP) helps scientists to
- **A.** compare gases.
- **B.** compare liquids.
- **C.** calculate ionic charges.
- **D.** calculate entropy.

14. Standard temperature and pressure (STP) occurs at
- **A.** 32°F.
- **B.** 100°F.
- **C.** 273°F.
- **D.** 373°F.

15. Standard temperature and pressure (STP) occurs at
- **A.** 273°C.
- **B.** 273 K.
- **C.** 0°F.
- **D.** 100°F.

16. Standard temperature and pressure (STP) occurs at
- **A.** 76 atm.
- **B.** 76 mm Hg.
- **C.** 760 atm.
- **D.** 760 mm Hg.

17. Standard temperature and pressure (STP) occurs at
- **A.** 14.7 atm.
- **B.** 14.7 mm Hg.
- **C.** 14.7 psi.
- **D.** 14.7 torr.

Standards Practice

Gases and Their Properties

4. e. **Students know how to convert between the Celsius and Kelvin temperature scales.**

18. How do the units in the Kelvin scale and the Celsius scale compare?
- **A.** The kelvin units are smaller than the Celsius units.
- **B.** The kelvin units are larger than the Celsius units.
- **C.** The units are equal for both scales.
- **D.** The scales are 100 units apart.

19. What is the freezing point of water in kelvins?
- **A.** 32 K
- **B.** 100 K
- **C.** 212 K
- **D.** 273 K

20. What is 20°C in kelvins?
- **A.** 253 K
- **B.** 273 K
- **C.** 293 K
- **D.** 373 K

21. What is 100°C in kelvins?
- **A.** 32 K
- **B.** 100 K
- **C.** 212 K
- **D.** 373 K

22. What is 0 K in Celsius?
- **A.** –373°C
- **B.** –273°C
- **C.** 100°C
- **D.** 212°C

23. What is 100 K in Celsius?
- **A.** –273°C
- **B.** –173°C
- **C.** 0°C
- **D.** 100°C

4. f. **Students know there is no temperature lower than 0 Kelvin.**

24. The coldest temperature possible is called
- **A.** absolute cold.
- **B.** absolute freeze.
- **C.** absolute nil.
- **D.** absolute zero.

25. The Kelvin scale
- **A.** has larger units than the Celsius scale.
- **B.** has units about half the size of the Fahrenheit scale.
- **C.** does not have negative numbers.
- **D.** is a theoretical scale only.

26. Which temperature is impossible?
- **A.** 20°C
- **B.** –20°F
- **C.** 20 K
- **D.** –20 K

27. Which temperature is impossible?
- **A.** –273°C
- **B.** –273°F
- **C.** –273 K
- **D.** 0 K

28. At absolute zero,
- **A.** the Fahrenheit scale ceases to exist.
- **B.** no further heat could be removed from a body.
- **C.** H_2O is in the form of a liquid.
- **D.** every substance must be in a gaseous phase.

Standards Practice
Gases and Their Properties

4. g. **Students know the kinetic theory of gases relates the absolute temperature of a gas to the average kinetic energy of its molecules or atoms.**

29. The kinetic molecular theory of gases explains the behavior of gases at the molecular level. All of these statements are part of this theory except
- **A.** gas molecules experience completely elastic collisions.
- **B.** all gas molecules have the same average kinetic energy at the same temperature.
- **C.** gas particles are in constant, random motion.
- **D.** gas molecules are incompressible.

30. Gas particles
- **A.** move faster as temperature increases.
- **B.** move faster as temperature decreases.
- **C.** move slower as temperature increases.
- **D.** do not show a correlation between movement and temperature.

31. At higher temperatures, gas molecules
- **A.** contract.
- **B.** slow down in movement.
- **C.** hit the walls of the container harder.
- **D.** hit the walls of the container softer.

32. At higher temperatures, gas molecules
- **A.** slow down.
- **B.** exert more pressure.
- **C.** have less energy.
- **D.** have more organization.

33. At absolute zero, gas molecules
- **A.** move slower than liquid molecules.
- **B.** are converted to liquid molecules.
- **C.** show very little movement.
- **D.** move in a straight line.

4. h. **Students know how to solve problems by using the ideal gas law in the form** *PV = nRT.*

Use the ideal gas law below to answer questions 34–38.

$$PV = n\mathrm{R}T$$

34. What happens to the volume of a gas when the pressure is increased by a factor of 4 (assuming all other factors remain the same)?
- **A.** It is reduced by a factor of 4.
- **B.** It is reduced by a factor of 2.
- **C.** It is increased by a factor of 2.
- **D.** It is increased by a factor of 4.

35. What happens to the volume of a gas when the temperature is increased by a factor of 4, if all other factors remain the same?
- **A.** It is reduced by a factor of 4.
- **B.** It is reduced by a factor of 2.
- **C.** It is increased by a factor of 2.
- **D.** It is increased by a factor of 4.

36. What is the pressure produced by 1.0 mol O_2 in a 22.4-L container at 273 K? Use R = 0.0821 (L•atm)/(mol•K).
- **A.** 0.5 atm
- **B.** 1.0 atm
- **C.** 2.0 atm
- **D.** 4.0 atm

37. What is the pressure produced by 1.0 mol O_2 in an 11.2-L container at 273 K? Use R = 0.0821 (L•atm)/(mol•K).
- **A.** 0.25 atm
- **B.** 0.5 atm
- **C.** 2.0 atm
- **D.** 4.0 atm

38. What is the pressure produced by 4.0 mol O_2 in a 22.4-L container at 273 K? Use R = 0.0821 (L• atm)/(mol•K).
- **A.** 0.25 atm
- **B.** 0.5 atm
- **C.** 2.0 atm
- **D.** 4.0 atm

Standards Practice

Gases and Their Properties

4. i. **Students know how to apply Dalton's law of partial pressures to describe the composition of gases and Graham's law to predict diffusion of gases.**

Use the table below to answer questions 39 and 40.

Composition of Air by Gas		
Gas	**Partial pressure (mmHg)**	**Percentage in air (%)**
Nitrogen (N_2)	594.0	78.1
Oxygen (O_2)	160.0	21.1
Carbon dioxide (CO_2)	0.3	0.04
Water vapor (H_2O)	5.7	0.75

39. In air, the partial pressures add up to

A. 160.0 mm Hg.
B. 594.0 mm Hg.
C. 600.0 mm Hg.
D. 760.0 mm Hg.

40. How does air illustrate Dalton's law of partial pressures?

A. Gases are necessary to make up Earth's atmosphere.
B. The amount of oxygen is 21.1 percent.
C. The pressures of gases in air add up to 1 atm.
D. Percentages of the gases in air add up to 100 percent.

41. According to Dalton's law of partial pressures, the addition of a gas from one tank to gas in another tank will cause the

A. pressures of the tanks to be added together.
B. pressure of the first tank to be subtracted from the pressure of the second tank.
C. pressures of the tanks to be multiplied together.
D. pressure of the first tank to be divided by the pressure of the second tank.

42. According to Graham's law, the distance traveled by a gas is inversely related to the square root of its molecular mass. If the molecular mass of Gas 1 was 4 times that of Gas 2, Gas 1 would travel

$$\frac{\text{Distance traveled by Gas 1}}{\text{Distance traveled by Gas 2}} = \frac{\sqrt{(\text{Molecular mass of Gas 1}}}{\sqrt{(\text{Molecular mass of Gas 2}}}$$

A. one-fourth the rate of Gas 2.
B. one-half the rate of Gas 2.
C. two times the rate of Gas 2.
D. four times the rate of Gas 2.

43. A tank containing 0.5 atm of oxygen gas is combined with a tank containing 0.5 atm of oxygen gas and a tank containing 2.5 atm of oxygen gas. What is the final pressure in the tank?

A. 0.5 atm
B. 1.5 atm
C. 3.5 atm
D. 4.5 atm

Standards Practice

Acids and Bases

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

5. a. **Students know the observable properties of acids, bases, and salt solutions.**

1. Which is a condition of an acid solution?
- **A.** It feels soapy.
- **B.** It is held together by ionic bonds.
- **C.** It tastes sour.
- **D.** It will neutralize a buffer.

2. A basic solution usually tastes
- **A.** bitter.
- **B.** salty.
- **C.** sour.
- **D.** sweet.

3. Salt solutions
- **A.** taste bitter.
- **B.** can conduct electricity.
- **C.** have many hydrogen ions.
- **D.** have many hydroxide ions.

4. Which is an acid?
- **A.** antacid
- **B.** apple juice
- **C.** drain cleaner
- **D.** oven cleaner

5. Which is a base?
- **A.** grapefruit juice
- **B.** grease remover
- **C.** soft drink
- **D.** vinegar

6. Which is a salt?
- **A.** HCl
- **B.** HNO_3
- **C.** KCl
- **D.** $NaOH$

5. b. **Students know acids are hydrogen-ion-donating and bases are hydrogen-ion-accepting substances.**

7. Acids
- **A.** donate hydrogen ions.
- **B.** donate hydroxide ions.
- **C.** accept anhydrous compounds.
- **D.** accept inert compounds.

8. Hydrogen ions are accepted by
- **A.** acids.
- **B.** bases.
- **C.** isotopes.
- **D.** salts.

9. Which is a hydrogen-ion donator?
- **A.** $HClO_2$
- **B.** KCl
- **C.** KOH
- **D.** $NaOH$

10. Which is a hydrogen-ion donator?
- **A.** CO_2
- **B.** HCl
- **C.** KOH
- **D.** $NaOH$

11. Which is a hydrogen-ion acceptor?
- **A.** CH_4
- **B.** H_2O
- **C.** H_2SO_4
- **D.** KOH

12. Which is a hydrogen-ion acceptor?
- **A.** CO
- **B.** HCl
- **C.** $HClO_2$
- **D.** $NaOH$

Standards Practice

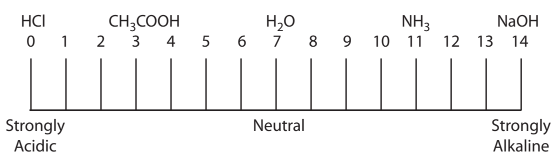
Acids and Bases

5. c. **Students know strong acids and bases fully dissociate and weak acids and bases partially dissociate.**

13. Which will fully dissociate in water?
- **A.** weak acid
- **B.** weak base
- **C.** strong base
- **D.** neutral solution

14. Strong acids or bases make the best electrolytes because they
- **A.** do not ionize in solution.
- **B.** react in an equilibrating manner.
- **C.** ionize completely in solution.
- **D.** have extremely small ionization constants.

15. Hydrochloric acid is a strong acid. How will it dissociate in water?



- **A.** It will not dissociate in water.
- **B.** It will partially dissociate in water.
- **C.** It will fully dissociate in water.
- **D.** It will not mix with water.

16. Ammonia is a weak base. How will it dissociate in water?
- **A.** It will not dissociate in water.
- **B.** It will partially dissociate in water.
- **C.** It will fully dissociate in water.
- **D.** It will not mix with water.

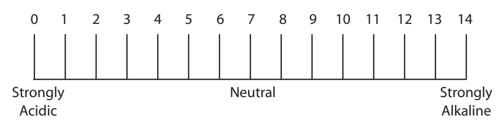
5. d. **Students know how to use the pH scale to characterize acid and base solutions.**

17. Which sequence shows the solutions listed from least acidic to most acidic?

Solution	pH
Baking soda	8.5
Gastric juice	1.5
Human blood	7.4
Vinegar	2.8

- **A.** gastric juice, vinegar, human blood, baking soda
- **B.** baking soda, human blood, vinegar, gastric juice
- **C.** baking soda, vinegar, human blood, gastric juice
- **D.** gastric juice, human blood, vinegar, baking soda

Use the pH scale below to answer questions 18–20.



18. Which of these decreases as the pH of a solution increases?
- **A.** basicity of a solution
- **B.** dissociation of water
- **C.** number of hydrogen ions
- **D.** percentage of hydroxide ions

19. Cherries have a pH of about 4.0. This makes cherries
- **A.** strong acids.
- **B.** strong bases.
- **C.** weak acids.
- **D.** weak bases.

20. Milk of magnesia has a pH of about 10.5. This makes milk of magnesia a
- **A.** strong acid.
- **B.** strong base.
- **C.** weak acid.
- **D.** weak base.

Standards Practice
Acids and Bases

5. e. **Students know the Arrhenius, Brønsted-Lowry, and Lewis acid-base definitions.**

21. Brønsted-Lowry acids and bases are defined using
- **A.** dissociation.
- **B.** electron pairs.
- **C.** hydrogen ions.
- **D.** hydroxide ions.

22. Lewis acids and bases are defined using
- **A.** dissociation.
- **B.** electron pairs.
- **C.** hydrogen ions.
- **D.** hydroxide ions.

23. An Arrhenius acid is a(n)
- **A.** hydroxide-ion producer.
- **B.** hydroxide-ion donor.
- **C.** ionic compound that dissociates into a metal and a hydrogen ion.
- **D.** substance that, in water, dissociates into hydroxide ions.

24. A Brønsted-Lowry acid is to a hydrogen-ion donor as a Brønsted-Lowry base is to a(n)
- **A.** hydrogen-ion acceptor.
- **B.** hydrogen-ion producer.
- **C.** hydroxide-ion donor.
- **D.** electron-pair donor.

25. A Lewis acid is a(n)
- **A.** hydrogen-ion acceptor.
- **B.** hydrogen-ion producer.
- **C.** hydrogen-ion donor.
- **D.** electron-pair acceptor.

5. f. **Students know how to calculate pH from the hydrogen-ion concentration.**

26. Grapefruit juice has a pOH of approximately 11.0. What is the pH of grapefruit juice?
- **A.** 3.0
- **B.** 5.0
- **C.** 9.0
- **D.** 11.0

27. Black coffee has a pOH of approximately 9.0. What is the pH of black coffee?
- **A.** 3.0
- **B.** 5.0
- **C.** 9.0
- **D.** 11.0

28. Which is the correct formula to determine pH?
- **A.** $pH = –\log[H_3O^+]$
- **B.** $pH = \log[H_3O^+]$
- **C.** $pH = 14 \times \log[H_3O^+]$
- **D.** $pH = \log[OH–]$

29. What is the pH of a substance with $1.0 \times 10^{–2}M$ $[H_3O^+]$?
- **A.** 2.0
- **B.** 4.0
- **C.** 8.0
- **D.** 12.0

30. What is the pH of a substance with $6.0 \times 10^{–8}M$ $[H_3O+]$?
- **A.** 6 + log 8
- **B.** 6 – log 8
- **C.** 8 + log 6
- **D.** 8 – log 6

Standards Practice

Acids and Bases

5. g. **Students know buffers stabilize pH in acid-base reactions.**

31. What is the purpose of a buffer in a solution?
- **A.** A buffer helps to neutralize a salt.
- **B.** A buffer helps to stabilize pH.
- **C.** A buffer helps to strengthen an acid.
- **D.** A buffer helps to strengthen a base.

32. Which makes a good buffer?
- **A.** strong acid and a strong base
- **B.** strong acid and a salt
- **C.** weak acid and a salt
- **D.** weak acid alone

33. Why must blood be a good buffer?
- **A.** Blood is naturally acidic.
- **B.** Blood is naturally basic.
- **C.** Oxygen carried in blood acts as a strong base.
- **D.** Humans take in acids and bases from foods.

34. A buffer must contain
- **A.** only hydrogen ions.
- **B.** only hydroxide ions.
- **C.** both hydrogen ions and hydroxide ions.
- **D.** neither hydrogen ions nor hydroxide ions.

35. Why is an acid-base conjugate pair necessary for a buffer?
- **A.** The acid and base should neutralize each other.
- **B.** The acid and base must not neutralize each other.
- **C.** The acid must be stronger than the base in order to work.
- **D.** The base must be stronger than the acid in order to work.

Standards Practice

Solutions

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

6. a. **Students know the definitions of *solute* and *solvent.***

1. A solution is a(n)
- **A.** compound.
- **B.** mixture.
- **C.** acid.
- **D.** base.

2. Which is true of a solute?
- **A.** It contains a dissolved substance.
- **B.** It is the substance present in the lesser amount.
- **C.** It donates electrons.
- **D.** It must be water.

3. Which is true of a solvent?
- **A.** It is the substance present in the greatest amount.
- **B.** It reacts with a dissolved substance.
- **C.** It must be an acid.
- **D.** It must be an base.

4. Which is not a solute?
- **A.** gas
- **B.** liquid
- **C.** solid
- **D.** quark

5. An example of a gas-in-gas solution is
- **A.** air.
- **B.** brass.
- **C.** a soft drink.
- **D.** water vapor.

6. One of nature's most common solvents is
- **A.** ammonia.
- **B.** plants.
- **C.** soil.
- **D.** water.

6. b. **Students know how to describe the dissolving process at the molecular level by using the concept of random molecular motion.**

7. The random motion of molecules helps molecules to
- **A.** return to their original position in a solution.
- **B.** bind to other molecules in a solution.
- **C.** move throughout a solution.
- **D.** separate into one layer in an oil-water solution.

8. Dissolving is caused at the molecular level by
- **A.** ionic forces.
- **B.** osmosis.
- **C.** random motion.
- **D.** weight.

9. Gases are easily distributed in a solution because
- **A.** their atomic weight is so light that attraction between particles is virtually impossible.
- **B.** their particles move quickly and are far apart, so they are not influenced by other gas particles.
- **C.** their temperature is so high that attraction between particles is virtually impossible.
- **D.** gases are the universal solvent.

10. Particles of liquids are distributed in a solution even though their
- **A.** random motion is not as great as that of gas particles.
- **B.** ionic attraction is not as high as the ionic attraction in gases.
- **C.** temperature is lower than the temperature normally needed to sustain a gas.
- **D.** molecules tend to be denser than those of gases.

Standards Practice
Solutions

6. c. **Students know temperature, pressure, and surface area affect the dissolving process.**

Use the table below to answer questions 11 and 12.

Mass of Sodium Nitrate that Can Be Dissolved in 100 mL of Water	
Temperature (°C)	**Mass (g)**
0	74
20	88
40	105
60	125
80	148

11. The table shows that the amount of sodium nitrate that can be dissolved in water
A. increases as the temperature increases.
B. increases as the surface area of molecules of sodium nitrate increases.
C. decreases as the molarity increases.
D. decreases as the pressure increases.

12. According to these data, approximately how many grams of sodium nitrate can be dissolved at 70°C?
A. 115 g
B. 125 g
C. 131 g
D. 137 g

13. Breaking a large solid into smaller pieces increases its rate of solvation in a solvent. This process accelerates the rate because
A. it makes the solid immiscible.
B. it creates an adiabatic environment.
C. greater surface area increases the likelihood of collisions.
D. greater surface area decreases the likelihood of collisions.

14. To increase the rate of solvation of carbon dioxide, a scientist might consider
A. decreasing the temperature and increasing the pressure.
B. increasing the temperature and decreasing the pressure.
C. decreasing the temperature and decreasing the pressure.
D. increasing the temperature and increasing the pressure.

6. d. **Students know how to calculate the concentration of a solute in terms of grams per liter, molarity, parts per million, and percent composition.**

15. Which is the correct formula for percent composition?
A. percent concentration of a solution = (amount of solute × amount of solution) × 100
B. percent concentration of a solution = (amount of solute ÷ amount of solution) × 100
C. percent concentration of a solution = (amount of solution – amount of solute) × 100
D. percent concentration of a solution = (amount of solution ÷ amount of solute) × 100

16. One part per million (ppm) is the same as
A. 1 mg/mL.
B. 1 mg/L.
C. 1 g/mL.
D. 1 g/L.

17. Suppose 16 g of solute is dissolved in 84 g of solvent. What is the mass percent concentration of the solution?
A. 16 percent
B. 42 percent
C. 84 percent
D. 100 percent

18. Suppose 8 mol of solute is dissolved in 2 L of solution. What is the molarity of the solution?
A. 2*M*
B. 4*M*
C. 8*M*
D. 16*M*

19. Suppose 10 g of solute is dissolved in 0.5 L of solution. What is the grams per liter of solvent in the solution?
A. 0.05 g/L
B. 5.0 g/L
C. 10 g/L
D. 20 g/L

Standards Practice
Solutions

6. e. **Students know the relationship between the molality of a solute in a solution and the solution's depressed freezing point or elevated boiling point.**

20. Which of these decreases as the amount of solute particles in a solution increases?
- **A.** boiling point
- **B.** osmotic pressure
- **C.** freezing point
- **D.** molality

21. How will adding salt to water affect the boiling point of water?
- **A.** It will lower the boiling point.
- **B.** It will raise the boiling point.
- **C.** It will not affect the boiling point.
- **D.** Water will not be able to boil.

22. As the molality of a solute in a solution decreases,
- **A.** the boiling point will decrease.
- **B.** the freezing point will decrease.
- **C.** the solution will not be able to boil.
- **D.** the solution will not be able to freeze.

23. The table shows the effects of various solutes in a given volume of water. Without knowing the actual values, which is the most likely reason that Na_2CO_3 will cause the greatest boiling point elevation?

Solute Effect on Water's Boiling Point		
Solute	**Quantity**	**Boiling Point Evaluation**
NaCl	1 mol	?
$KMnO_4$	1 mol	?
Na_2CO_3	1 mol	?
$C_6H_{12}O_6$	1 mol	?

- **A.** Na_2CO_3 is the only solute that exhibits the Tyndall effect.
- **B.** Na_2CO_3 produces the smallest number of moles in solution.
- **C.** Na_2CO_3 has the greatest heat of enthalpy.
- **D.** Na_2CO_3 produces the largest number of solute particles in solution.

6. f. **Students know how molecules in a solution are separated or purified by the methods of chromatography and distillation.**

24. Which uses electrostatic attraction to separate a solute from a solvent?
- **A.** chromatography
- **B.** distillation
- **C.** evaporation
- **D.** filtering

25. Chromatography works because
- **A.** certain molecules are too large to pass through a porous barrier.
- **B.** certain molecules are more attracted to a given material than other molecules.
- **C.** boiling a solution will cause one material to boil off before the other.
- **D.** heating a solution will increase the random motion of molecules.

26. Distillation is a separation technique that involves
- **A.** using a porous barrier to separate a solid from a liquid.
- **B.** separating dissolved substances based on their tendency to be drawn across a surface.
- **C.** the formation of pure, solid particles of a substance from a solution containing the dissolved substance.
- **D.** separating two or more liquids based on differences in their boiling points.

27. A quality-control technician is given a saline solution and asked to separate it into its components and to determine the masses of the salt and the water. Which design would be best for the technician to use to carry out this task?
- **A.** distillation apparatus and balance
- **B.** filtration setup and balance
- **C.** fine-wire mesh screen and balance
- **D.** gas chromatograph and balance

Standards Practice

Chemical Thermodynamics

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

7. a. **Students know how to describe temperature and heat flow in terms of the motion of molecules (or atoms).**

1. Heat energy is transferred through
- **A.** active transport.
- **B.** entropy.
- **C.** molecules.
- **D.** temperature.

2. When temperature drops,
- **A.** energy increases.
- **B.** molecules move slower.
- **C.** collisions occur more often.
- **D.** entropy increases.

3. Heating a substance increases
- **A.** atomic motion.
- **B.** molecular motion.
- **C.** both atomic and molecular motion.
- **D.** neither atomic nor molecular motion.

4. How are increased temperature of a substance and entropy related?
- **A.** Increased temperature will increase entropy.
- **B.** Increased temperature will decrease entropy.
- **C.** Increased temperature will stabilize entropy.
- **D.** Increased temperature will not affect entropy.

7. b. **Students know chemical processes can either release (exothermic) or absorb (endothermic) thermal energy.**

5. Which is an endothermic reaction?
- **A.** breaking a chemical bond
- **B.** combustion of wood
- **C.** evaporating water
- **D.** making ice

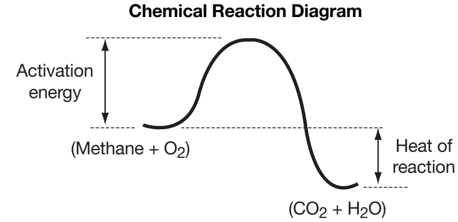
6. Which is an exothermic reaction?
- **A.** boiling an egg
- **B.** mixing water and a strong acid
- **C.** melting an ice cube
- **D.** evaporating water

7. In order for a reaction to be called exothermic,
- **A.** the enthalpy of the reactants must be less than that of the products.
- **B.** the sign of the change in enthalpy for the reactants must be positive.
- **C.** the enthalpy of the products must be less than that of the reactants.
- **D.** heat must flow from surroundings into the system.

8. Fusion, or melting, is an endothermic process because it
- **A.** requires heat to be transferred from system to surroundings and has a negative ΔH.
- **B.** requires heat to be transferred from the surroundings to system and has a positive ΔH.
- **C.** involves a decrease in entropy.
- **D.** involves a decrease in kinetic energy.

9. A student performed the following experiment. He drew and labeled the graph below based on his results. Which best describes his results?

Chemical Reaction Diagram



- **A.** The reaction is endothermic.
- **B.** The reaction is exothermic.
- **C.** The reaction requires no energy.
- **D.** The reaction shows no entropy.

Standards Practice

Chemical Thermodynamics

7. c. **Students know energy is released when a material condenses or freezes and is absorbed when a material evaporates or melts.**

10. Which of these phase changes does not involve the absorption of heat energy?
- **A.** boiling
- **B.** condensation
- **C.** melting
- **D.** vaporization

11. When a material boils, energy is
- **A.** absorbed by the material.
- **B.** released by the material.
- **C.** absorbed by a catalyst.
- **D.** not exchanged.

12. When water evaporates, energy is
- **A.** not exchanged.
- **B.** transferred to a catalyst.
- **C.** absorbed by the water.
- **D.** released by the water vapor.

13. When steam condenses, energy is
- **A.** not exchanged.
- **B.** transferred to a catalyst.
- **C.** absorbed by the water.
- **D.** released by the water vapor.

7. d. **Students know how to solve problems involving heat flow and temperature changes, using known values of specific heat and latent heat of phase change.**

14. The equation shows the change in enthalpy when 1 mol of liquid water vaporizes into water vapor. This is called the molar heat of vaporization. Given this information, which is the proper value for the molar heat of condensation?

$$H_2O(l) = H_2O(g) \Delta Hvap = +40.7 \text{ kJ}$$

- **A.** ΔHcond = –40.7 kJ
- **B.** ΔHcond = 0 kJ
- **C.** ΔHcond = –571.6 kJ
- **D.** ΔHcond = +571.6 kJ

15. How much heat in joules does it take to raise 10 g of water from 25°C to 35°C? [The specific heat of water is 4.180 J/g•˚C]

$$q = mc\Delta T$$

- **A.** 41.8 J
- **B.** 418 J
- **C.** 4180 J
- **D.** 41,800 J

16. How is the amount of heat needed to melt ice related to the amount of heat needed to freeze the same amount of water?
- **A.** The amount needed to melt ice is twice as high as the amount needed to freeze water.
- **B.** The amount needed to melt ice is three times as high as the amount needed to freeze water.
- **C.** The two amounts are the same but with the opposite sign (positive versus negative).
- **D.** The two values are not related to each other.

17. How much heat in joules is required to melt 3 mol of ice? (The heat of fusion is 6.01 kJ/mol.)
- **A.** 6.01 kJ
- **B.** 18.03 kJ
- **C.** –6.01 kJ
- **D.** –18.03 kJ

Standards Practice
Chemical Thermodynamics

7. e. **Students know how to apply Hess's law to calculate enthalpy change in a reaction.**

18. Which of these increases when the sign of ΔS (entropy change) is negative?
- **A.** disorder
- **B.** temperature
- **C.** order
- **D.** kinetic energy

19. Hess's law states that the heat of a reaction
- **A.** is directly related to the temperature of the reactants.
- **B.** is the same, whether the reaction is carried out in one step or several.
- **C.** is cumulative, with the number of steps increasing the heat by a specific constant.
- **D.** varies depending on laboratory conditions.

20. When calculating enthalpy, ΔH is
- **A.** ΔHproducts + ΔHreactants.
- **B.** ΔHproducts – ΔHreactants.
- **C.** ΔHproducts × ΔHreactants.
- **D.** ΔHproducts ÷ ΔHreactants.

21. When calculating enthalpy change in a reaction, what is the enthalpy of formation for an element in its most stable state under the experimental condition (for example, hydrogen gas at 72°C)?
- **A.** 0 kJ/mol
- **B.** 100 kJ/mol
- **C.** –100 kJ/mol
- **D.** 1000 kJ/mol

22. How do you figure the enthalpy change for the reaction shown below?

$$C_2H_4\ (g) + 3O_2\ (g) \rightarrow 2CO_2\ (g) + 2H_2O\ (l)$$

Enthalpies: C_2H_4 (g) = 52.26 kJ/mol
O_2 (g) = 0 kJ/mol
CO_2 (g) = –393.5 kJ/mol
H_2O (l) = –285.8 kJ/mol

- **A.** [2(–393.5 kJ/mol) + 2(–285.8 kJ/mol)] – [(52.26 kJ/mol) + 3(0 kJ/mol)]
- **B.** –[2(–393.5 kJ/mol) + 2(–285.8 kJ/mol)] – [(52.26 kJ/mol) + 3(0 kJ/mol)]
- **C.** [(52.26 kJ/mol) + 3(0 kJ/mol)] + [2(–393.5 kJ/mol) + 2(–285.8 kJ/mol)]
- **D.** [(52.26 kJ/mol) + 3(0 kJ/mol)] – [2(–393.5 kJ/mol) + 2(–285.8 kJ/mol)]

7. f. **Students know how to use the Gibbs free energy equation to determine whether a reaction would be spontaneous.**

23. Which of these is not a variable in the Gibbs free energy equation, which determines reaction spontaneity?
- **A.** endothermy
- **B.** enthalpy
- **C.** entropy
- **D.** temperature

24. Which of these is the Gibbs free energy equation?
- **A.** $\Delta G = \Delta H \times T\Delta S$
- **B.** $\Delta G = T\Delta S \div \Delta H$
- **C.** $\Delta G = \Delta H - T\Delta S$
- **D.** $\Delta G = T\Delta S - \Delta H$

25. If a reaction has a positive change in free energy, it is
- **A.** a spontaneous process.
- **B.** nonspontaneous process.
- **C.** at equilibrium.
- **D.** an exothermic reaction.

26. Which of these would always be called a spontaneous reaction?
- **A.** reaction with $-\Delta S$ and a $+\Delta H$
- **B.** reaction with $-\Delta H$ and a $-\Delta S$
- **C.** reaction with $+\Delta H$ and a $+\Delta S$
- **D.** reaction with $+\Delta S$ and a $-\Delta H$

27. Which of these combinations of factors must be true for a reaction to be nonspontaneous?
- **A.** ΔG is negative; ΔS is positive.
- **B.** ΔG is positive; ΔS is positive.
- **C.** ΔG is negative; ΔS is negative.
- **D.** ΔG is positive; ΔS is negative.

Standards Practice

Reaction Rates

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

8. a. **Students know the rate of reaction is the decrease in concentration of reactants or the increase in concentration of products with time.**

1. Which of these is required to calculate the rate of a reaction?
   - **A.** The change in enthalpy over time for the reaction.
   - **B.** The time it takes the reaction to go halfway to completion.
   - **C.** The change in concentration of either the product or reactant over time.
   - **D.** The change in temperature for the reaction over time.

2. The rate of a reaction can be described by calculating
   - **A.** evidence of reactant depletion.
   - **B.** evidence of product formation.
   - **C.** the increase in the concentration of the reactants with time.
   - **D.** the increase in the concentration of the products with time.

3. The rate of a reaction is measured by calculating change in
   - **A.** concentration.
   - **B.** pressure.
   - **C.** temperature.
   - **D.** volume.

4. Which indicates a reaction rate?
   - **A.** the activation energy
   - **B.** presence of a catalyst
   - **C.** decrease in reactant concentration after 1 h
   - **D.** evidence of product formation

5. The rate of the following reaction can be described using the

$$Fe + 2HCl \rightarrow FeCl_2 + H_2$$

   - **A.** concentration change of Cl_2.
   - **B.** concentration change of Fe.
   - **C.** presence of $FeCl_2$.
   - **D.** absence of O_2.

8. b. **Students know how reaction rates depend on such factors as concentration, temperature, and pressure.**

6. Which would not increase the rate of a reaction?
   - **A.** increase in the concentration of reactants
   - **B.** removal of a catalyst
   - **C.** increase in solvent volume
   - **D.** increase in temperature

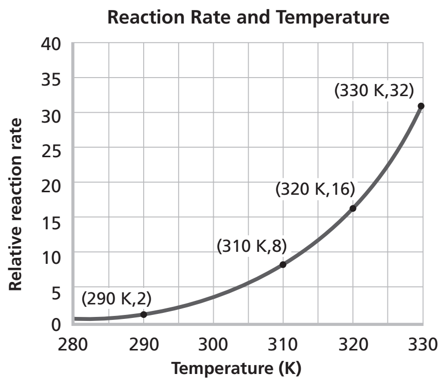
7. A change in temperature affects the reaction rate because as the temperature increases,
   - **A.** more collisions with the required energy occur.
   - **B.** molecules become more pliable.
   - **C.** pressure decreases.
   - **D.** molecules expand and take up more room.

8. The formula below shows the rate law for a certain reaction. Which gives the correct results when the concentration of NO is doubled?

$$\boxed{\text{Rate} = k[NO]^2[Cl_2]}$$

   - **A.** The reaction rate is unaffected.
   - **B.** The reaction rate doubles.
   - **C.** The reaction rate triples.
   - **D.** The reaction rate quadruples.

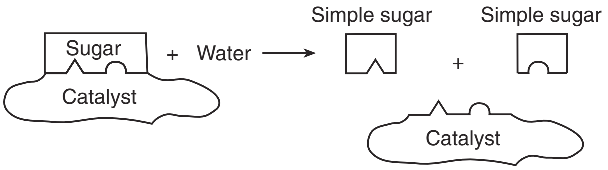
9. The reaction rate of a substance is shown below. What does this graph illustrate?



   - **A.** As the concentration increases, the reaction rate decreases.
   - **B.** As the temperature increases, the reaction rate increases.
   - **C.** As the pressure increases, the reaction rate decreases.
   - **D.** The reaction rate stays constant.

8. c. **Students know the role a catalyst plays in increasing the reaction rate.**

10. The diagram below illustrates a biochemical process that occurs in organisms. In the diagram, the substance labeled *catalyst*



A. lowers the reaction rate.
B. splits the sugar into its components.
C. raises the activation energy.
D. creates a larger molecule from two smaller ones.

11. How does a catalyst increase the rate of a chemical reaction?
A. by increasing the concentration of the reactants
B. by increasing the speed of the molecules
C. by lowering the energy of the products
D. by lowering the activation energy

12. A catalyst
A. lowers the amount of energy needed to initiate a reaction.
B. causes a reaction to slow down.
C. shifts the equilibrium constant of a chemical reaction.
D. causes a reaction to stop.

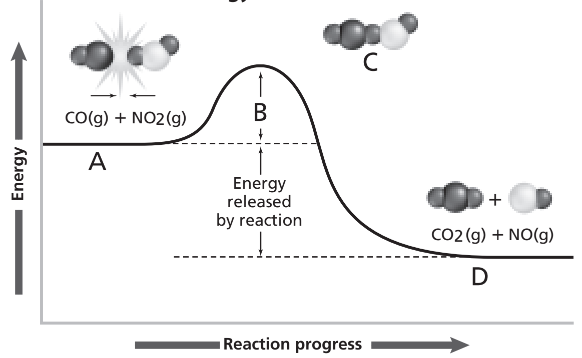
13. How does a catalyst usually emerge from a reaction?
A. in a decreased concentration after the reaction
B. as a different molecule than it was before the reaction
C. in a different phase than it was before the reaction
D. in the same form as it was before the reaction

8. d. **Students know the definition and role of activation energy in a chemical reaction.**

14. Activation energy is the energy
A. required to start a chemical reaction.
B. expended by uranium during one half-life.
C. given off by 1 g of water as it freezes.
D. needed to heat 1 g of water 1°C.

15. The energy of a reaction rate is shown below. Where should the label *Activation Energy* go?

Energy of a Reaction Rate



A. A
B. B
C. C
D. D

16. As activation energy decreases, the rate of a reaction
A. decreases.
B. increases.
C. stays the same.
D. stops.

17. Which decreases as the activation energy for a reaction increases?
A. number of inactivated complexes
B. entropy
C. amount of reactants
D. reaction rate

Standards Practice
Chemical Equilibrium

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

9. a. **Students know how to use Le Châtelier's principle to predict the effect of changes in concentration, temperature, and pressure.**

1. Le Châtelier's principle states that equilibrium of a reaction can be affected by a change in
- **A.** concentration of reactants only.
- **B.** pressure or temperature only.
- **C.** pressure, temperature, or volume only.
- **D.** concentration, pressure, temperature, or volume only.

2. Henri Le Châtelier believed a change in a reaction would result in a new
- **A.** equilibrium.
- **B.** pressure.
- **C.** temperature.
- **D.** volume.

3. According to Le Châtelier's principle, if the reactants are decreased in a reaction at equilibrium, there will be a net reaction
- **A.** to the left and no equilibrium.
- **B.** to the left until a new equilibrium occurs.
- **C.** to the right and no equilibrium.
- **D.** to the right until a new equilibrium occurs.

4. According to Le Châtelier's principle, if the products are increased in a reaction at equilibrium, there will be a net reaction
- **A.** to the left and no equilibrium.
- **B.** to the left until a new equilibrium occurs.
- **C.** to the right and no equilibrium.
- **D.** to the right until a new equilibrium occurs.

5. Use the equation below. According to Le Châtelier's principle, what will happen if the concentration of N_2 is increased in the presence of unlimited H_2 at equilibrium?

$$N_2\,(g) + 3H_2\,(g) \rightleftharpoons 2NH_3\,(g)$$

- **A.** There will be a net reaction to the left and no equilibrium.
- **B.** There will be a net reaction to the left until a new equilibrium occurs.
- **C.** There will be a net reaction to the right and no equilibrium.
- **D.** There will be a net reaction to the right until a new equilibrium occurs.

9. b. **Students know equilibrium is established when forward and reverse reaction rates are equal.**

6. In order for equilibrium to occur, a reaction must be
- **A.** organic.
- **B.** reversible.
- **C.** endothermic.
- **D.** exothermic.

7. At equilibrium, the forward reaction rate is
- **A.** greater than the reverse reaction rate.
- **B.** less than the reverse reaction rate.
- **C.** equal to the reverse reaction rate.
- **D.** independent of the reverse reaction rate.

8. The reaction below is exothermic. To produce less product (decrease the forward reaction), what experimental change would be most effective?

$$2SO_2\,(g) + O_2\,(g) \rightleftharpoons 2SO_3\,(g)\Delta H^\circ rxn = -198\text{ kJ}$$

- **A.** decrease volume
- **B.** decrease temperature
- **C.** increase pressure
- **D.** increase temperature

9. Which statement most accurately describes how an endothermic chemical reaction would be affected by an increase in temperature?
- **A.** The reaction would move backward until a new equilibrium occurred.
- **B.** The reaction would move forward until a new equilibrium occurred.
- **C.** The reaction would alternate moving backward and forward.
- **D.** An increase in temperature would not affect an endothermic reaction.

Standards Practice

Chemical Equilibrium

9. c. **Students know how to write and calculate an equilibrium constant expression for a reaction.**

10. Which symbol is usually used to show the equilibrium constant of a reaction?
- **A.** K
- **B.** P
- **C.** Q
- **D.** S

11. How would the equation for an equilibrium constant for which $2NO_2$ is one of the components be expressed?
- **A.** $[NO_2]$
- **B.** $[N] + [O_2]$
- **C.** $4[NO]$
- **D.** $[NO_2]^2$

12. Which of these is the correct way to express the equilibrium constant for the reaction below?

2CO (g) + Cl2 (g) qe COCl2 (g)

- **A.** $[COCl_2] \times ([CO] + [Cl_2])$
- **B.** $[COCl_2] \div [CO]^2[Cl_2]$
- **C.** $[CO]^2 + [Cl_2] + [COCl_2]$
- **D.** $[CO] + 2[Cl] + 2[COCl_2]$

13. Which of these is the correct way to express the equilibrium constant for the reverse of the forward reaction shown below?

$$2HI\ (g) \rightarrow H_2\ (g) + I_2\ (g)$$

- **A.** $[H_2] + [I_2] + [HI]$
- **B.** $[H_2][I_2] \times [HI]^2$
- **C.** $2[H_2][I_2] \div [HI]$
- **D.** $[HI]^2 \div [H_2][I_2]$

14. If the equilibrium constant for a forward reaction is known, K_{eq}, what is the equilibrium constant for the reverse reaction?
- **A.** K_{eq}
- **B.** $-K_{eq}$
- **C.** $1 - K_{eq}$
- **D.** $1 \div K_{eq}$

Standards Practice
Organic Chemistry and Biochemistry

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

10. a. **Students know large molecules (polymers), such as proteins, nucleic acids, and starch, are formed by repetitive combinations of simple subunits.**

1. Which group contains only molecules that are each assembled from smaller organic compounds?
- **A.** proteins, carbon dioxide, DNA, starch
- **B.** proteins, carbon dioxide, starch, water
- **C.** proteins, DNA, fats, starch
- **D.** proteins, DNA, fats, water

2. Which compound is composed of many monosaccharides?
- **A.** lipid
- **B.** nucleic acid
- **C.** protein
- **D.** starch

3. Which compound is composed of many nucleotides?
- **A.** lipid
- **B.** nucleic acid
- **C.** protein
- **D.** starch

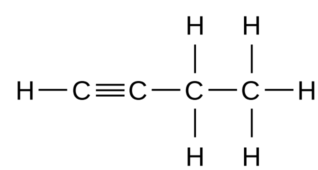
4. Which compound is composed of many amino acids?
- **A.** lipid
- **B.** nucleic acid
- **C.** protein
- **D.** starch

5. Which compound is composed of many fatty acids?
- **A.** lipid
- **B.** nucleic acid
- **C.** protein
- **D.** starch

10. b. **Students know the bonding characteristics of carbon that result in the formation of a large variety of structures ranging from simple hydrocarbons to complex polymers and biological molecules.**

6. Carbohydrates are important molecules in all living things. Which elements are found in all carbohydrates?
- **A.** calcium, carbon, potassium
- **B.** carbon, hydrogen, oxygen
- **C.** hydrogen, nitrogen, phosphorus
- **D.** nitrogen, oxygen, sulfur

7. Which types of bonds are formed by carbon in the molecule shown below?



- **A.** single bond and double bond
- **B.** single bond and benzene ring
- **C.** single bond and triple bond
- **D.** triple bond and benzene ring

8. Which type of bond is not formed by carbon?
- **A.** single
- **B.** double
- **C.** triple
- **D.** quadruple

9. The versatile bonding characteristics of carbon make it ideal for
- **A.** being the standard for a mole.
- **B.** having isotopes.
- **C.** forming a single bond.
- **D.** forming organic molecules.

10. What gives carbon its versatile bonding characteristics?
- **A.** its atomic weight
- **B.** its isotopes
- **C.** its electron configuration in orbitals
- **D.** its affinity for hydrogen

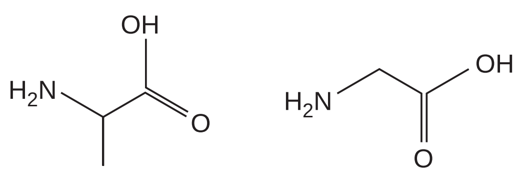
Standards Practice
Organic Chemistry and Biochemistry

10. c. **Students know amino acids are the building blocks of proteins.**

11. The building block that makes a protein is a(n)
A. amino acid.
B. benzene ring.
C. oxygen atom.
D. nitrogen atom.

12. In a protein, which is most important?
A. atomic mass of its building blocks
B. bonds of its building blocks
C. number of its building blocks
D. order of its building blocks

13. The diagram below shows two amino acids. What would biochemists call the result of chaining many of these molecules together?



Alanine   Glycine

A. carbohydrate
B. lipid
C. nucleic acid
D. protein

14. How many amino acids exist in the human body?
A. 1
B. 7
C. 20
D. 50

15. What is the minimum number of building blocks needed to make a protein?
A. 1
B. 10
C. 25
D. 50

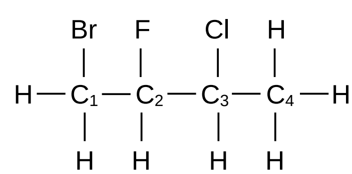
10. d. **Students know the system for naming the ten simplest linear hydrocarbons and isomers that contain single bonds, simple hydrocarbons with double and triple bonds, and simple molecules that contain a benzene ring.**

16. What is the suffix for the name of a carbon-chain molecule containing a double bond?
A. -ane
B. -ene
C. -yne
D. -amide

17. What is the suffix for the name of a carbon-chain molecule containing a triple bond?
A. -ane
B. -ene
C. -yne
D. -amide

18. What is the prefix for the name of a carbon-chain molecule containing five carbon atoms?
A. meth-
B. non-
C. pent-
D. pro-

19. The base molecule of the molecule shown below is



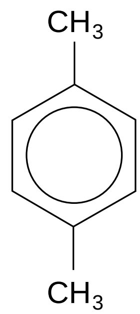
A. butane.
B. butene.
C. propane.
D. propene.

20. What is the name of C_2H_4?
A. butane
B. ethene
C. methane
D. propyne

Standards Practice
Organic Chemistry and Biochemistry

10. e. **Students know how to identify the functional groups that form the basis of alcohols, ketones, ethers, amines, esters, aldehydes, and organic acids.**

21. This compound must be classified as aromatic because it



A. has two methyl groups.
B. contains a propene ring.
C. has ten carbon atoms.
D. has a benzene ring.

22. An aldehyde has a functional group of
A. –CHO.
B. –CO–.
C. –COOH.
D. –OH.

23. An ether has a functional group of
A. –CHO.
B. –CO–.
C. –O–.
D. –OH.

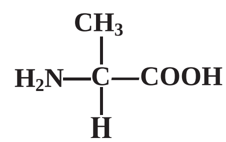
24. –OH is to alcohols as $–NH_2$ is to
A. alkenes.
B. amines.
C. esters.
D. ketones.

25. A compound containing a chain of a carbon atom and an oxygen atom with another oxygen atom double bonded to the carbon atom is an
A. alcohol.
B. alkyne.
C. anime.
D. ester.

10. f. **Students know the R-group structure of amino acids and know how they combine to form the polypeptide backbone structure of proteins.**

26. Which molecule is not found in an amino acid?
A. carbon
B. nitrogen
C. oxygen
D. potassium

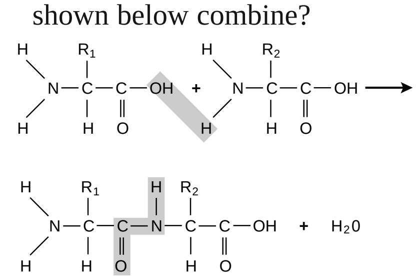
27. Which group of molecules makes up the R group of the amino acid alanine, shown below?



A. CH_3
B. $COO^–$
C. H
D. NH_3^+

28. The R group of an amino acid
A. reacts with peptides.
B. combines to help form proteins.
C. is also the functional group of an amine.
D. identifies the amino acid.

29. Which of these two groups provides the bonding sites when the two amino acids shown below combine?



A. carboxyl and hydrogen
B. aldehyde and amino
C. ketone and amino
D. amino and carboxyl

30. The COO– group and NH3+ group of amino acids combine to form
A. galactose.
B. triglyceride.
C. a fatty acid.
D. a peptide.

Standards Practice

Nuclear Processes

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

11. a. Students know protons and neutrons in the nucleus are held together by nuclear forces that overcome the electromagnetic repulsion between the protons.

1. A nucleus of an atom is partially held together by the attraction between
- **A.** electrons and neutrons.
- **B.** electrons and protons.
- **C.** protons and neutrons.
- **D.** protons and quarks.

2. How do the particles in a nucleus of an atom interact with each other?
- **A.** Protons attract each other as neutrons.
- **B.** Protons attract each other but are repelled by neutrons.
- **C.** Protons do not interact with each other.
- **D.** Protons repel each other but attract electrons.

3. The short range of the nuclear force of atoms means that
- **A.** nuclei with too many protons and neutrons are less stable than smaller nuclei.
- **B.** protons and neutrons repel each other.
- **C.** large nuclei are held together more tightly than are small nuclei.
- **D.** the nuclear force becomes more important for large nuclei.

Use the chart below to answer question 4.

Number of protons v. number of neutrons	Number of stable nuclei
Even number of protons/even number of neutrons	about 160
Even number of protons/odd number of neutrons	about 50
Odd number of protons/even number of neutrons	about 50
Odd number of protons/odd number of neutrons	4

4. Which is the most likely conclusion that can be drawn from the chart? If an element has an
- **A.** even number of protons and an even atomic mass, it will be unstable.
- **B.** even number of protons but an odd atomic mass, it will be unstable.
- **C.** odd number of protons and an odd atomic mass, it will be unstable.
- **D.** odd number of protons but an even atomic mass, it will be unstable.

11. b. Students know the energy release per gram of material is much larger in nuclear fusion or fission reactions than in chemical reactions. The change in mass (calculated by $E = mc^2$) is small but significant in nuclear reactions.

5. What do nuclear fusion and nuclear fission have in common?
- **A.** Both take place regularly in nature.
- **B.** With both, large amounts of energy are released.
- **C.** With both, heavy nuclei are split into lighter ones.
- **D.** With both, light nuclei are formed into heavier ones.

6. In nuclear fission,
- **A.** carbon atoms are released.
- **B.** mass is converted to energy.
- **C.** much energy is absorbed.
- **D.** a nucleus is converted to a larger nucleus.

7. An example of nuclear fusion is
- **A.** the breakdown of uranium.
- **B.** the splitting of an atom into particles.
- **C.** a reaction that takes place in the Sun.
- **D.** the combining of sodium and chlorine ions to form salt.

8. Which type of reaction has the largest release of energy per gram of material?
- **A.** breakdown of complex sugars into carbohydrates
- **B.** formation of proteins from amino acids
- **C.** fission of uranium into thorium and helium
- **D.** formation of a salt from potassium and chlorine

9. The energy release from the nuclear fission of 1 g of material is approximately equivalent to the energy release from
- **A.** breaking down 1 g of salt into its molecular components.
- **B.** breaking down 1 g of sugar into its molecular components.
- **C.** burning 5 large logs of wood.
- **D.** burning 3 tons of coal.

Standards Practice

Nuclear Processes

11. c. Students know some naturally occurring isotopes of elements are radioactive, as are isotopes formed in nuclear reactions.

10. Which statement about isotopes is true?
- **A.** An isotope of an element has a different number of protons in the nucleus.
- **B.** All isotopes have only one neutron in the nucleus.
- **C.** Isotopes of an element have different numbers of electrons.
- **D.** Isotopes of an element can be radioactive.

11. An isotope formed in a nuclear reaction is most likely to
- **A.** be inert.
- **B.** be radioactive.
- **C.** have a lower atomic number than uranium.
- **D.** have a higher atomic number than uranium.

12. Which is most likely to be radioactive?
- **A.** ion of sodium
- **B.** isotope of uranium
- **C.** proton of hydrogen
- **D.** proton of oxygen

13. An isotope of uranium emits a particle composed of two protons and two neutrons. What remains after the isotope emits the particle?
- **A.** new element that has a mass smaller than the uranium mass
- **B.** new element that has a mass greater than the uranium mass
- **C.** new isotope of uranium that has a mass greater than the starting mass
- **D.** new isotope of uranium that has a mass smaller than the starting mass

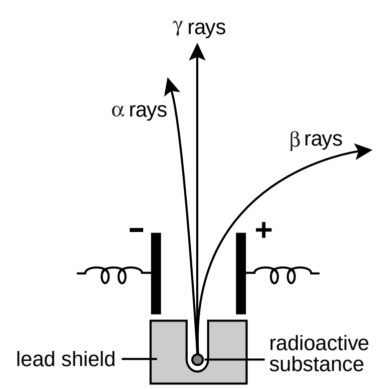
11. d. Students know the three most common forms of radioactive decay (alpha, beta, and gamma) and know how the nucleus changes in each type of decay.

14. Which form of radioactive decay consists of photons, and thus has no matter?
- **A.** alpha
- **B.** beta
- **C.** delta
- **D.** gamma

15. When an alpha particle is emitted, the nucleus of the original element
- **A.** remains unchanged.
- **B.** loses one neutron.
- **C.** loses one proton.
- **D.** loses two protons.

16. What happens when a beta particle is emitted?
- **A.** A proton in the original nucleus is reduced to quarks.
- **B.** A neutron in the original nucleus is converted to a proton and an electron.
- **C.** The original nucleus is reduced in atomic number.
- **D.** The original nucleus remains unchanged.

17. The diagram below shows how alpha, beta, and gamma rays are affected by two electrically charged plates. Based on the paths the rays follow, what are the respective charges of alpha, beta, and gamma rays?

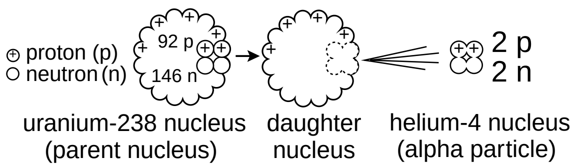


- **A.** negative, positive, none
- **B.** positive, negative, none
- **C.** negative, none, positive
- **D.** positive, none, negative

Standards Practice
Nuclear Processes

11. e. **Students know alpha, beta, and gamma radiation produce different amounts and kinds of damage in matter and have different penetrations.**

18. The particle emitted by the uranium-238 nucleus that contains two protons and two neutrons



A. has an electromagnetic field.
B. has a negative charge.
C. can be stopped by a piece of paper.
D. will not be deflected in an electric field.

19. Alpha particles
A. are more penetrating than beta particles.
B. are more penetrating than gamma particles.
C. are the least penetrating of all radioactive decay.
D. are the most penetrating of all radioactive decay.

20. Beta particles
A. penetrate more than alpha particles.
B. penetrate more than gamma particles.
C. are the least penetrating of all radioactive decay.
D. are the most penetrating of all radioactive decay.

21. Beta particles can
A. penetrate farther than gamma particles.
B. penetrate a 1-mm thick piece of aluminum.
C. penetrate a 1-m lead wall.
D. be stopped by a piece of paper.

22. Gamma particles
A. are deflected by an electric field.
B. can be stopped by a piece of paper.
C. can be stopped by a 1-mm thick piece of aluminum.
D. penetrate materials similar to the way that X rays penetrate materials.

11. f. **Students know how to calculate the amount of a radioactive substance remaining after an integral number of half-lives have passed.**

Use the chart below to answer questions 23–25.

Half Lives Passed	Amount of Radioactive Material
None	100%
1	A
2	B
3	C

23. After one half-life has passed, how much of the original radioactive material will remain (A)?
A. 12.5 percent
B. 25 percent
C. 50 percent
D. 100 percent

24. After two half-lives have passed, how much of the original radioactive material will remain (B)?
A. 12.5 percent
B. 25 percent
C. 50 percent
D. 100 percent

25. After three half-lives have passed, how much of the original radioactive material will remain?
A. 3.125 percent
B. 6.25 percent
C. 12.5 percent
D. 25 percent

26. Uranium-238 has a half-life of 4.51×10^9 y. How much of the original radioactive material will remain after 9.02×10^9 years?
A. 1/2
B. 1/4
C. 1/8
D. 1/16

27. If 25 percent of the original radioactive material of polonium-210 remains after 276 days, what is its half-life?
A. 138 days
B. 276 days
C. 414 days
D. 552 days

Standards Practice

Nuclear Processes

11. g. **Students know protons and neutrons have substructures and consist of particles called quarks.**

28. Which is not smaller than an atom?
- **A.** isotope
- **B.** neutron
- **C.** proton
- **D.** quark

29. Which shows the particles in order from smallest to largest?
- **A.** atom, molecule, proton, quark
- **B.** molecule, atom, proton, quark
- **C.** proton, atom, quark, molecule
- **D.** quark, proton, atom, molecule

30. Which is not the name of a quark?
- **A.** charmed
- **B.** east
- **C.** strange
- **D.** up

31. How many known quarks are there?
- **A.** 1
- **B.** 2
- **C.** 4
- **D.** 6

32. A quark is a substructure of
- **A.** electrons and neutrons.
- **B.** electrons and protons.
- **C.** isotopes and neutrons.
- **D.** neutrons and protons.

33. Individual quarks
- **A.** are often seen.
- **B.** are easily isolated.
- **C.** are inside subatomic particles.
- **D.** have a charge of 1^-.

Sample Test

Read each question, and choose the best answer. Then, on your answer sheet, mark the answer choice that you think is best.

1. What kind of bond exists between the copper and the sulfate ions in $CuSO_4$? **2. b.**
A. covalent
B. ionic
C. metallic
D. nuclear

2. Aloe juice contains $1.0 \times 10^{-6}M$ of hydroxide ions. What is its pH? **5. f.**
A. 1.0
B. 6.0
C. 8.0
D. 16.0

3. How do you determine the enthalpy change for the reaction shown below? **7. e.**

$$H_2\ (g) + (1/2)O_2\ (g) \rightarrow H_2O\ (l)$$

A. (enthalpy of water) + [(enthalpy of hydrogen) – (1/2)(enthalpy of oxygen)]
B. (enthalpy of water) – [–(enthalpy of hydrogen) + –(1/2)(enthalpy of oxygen)]
C. (enthalpy of water) – [(enthalpy of hydrogen) + (1/2)(enthalpy of oxygen)]
D. (enthalpy of water) + [(enthalpy of hydrogen) + (1/2)(enthalpy of oxygen)]

4. In this equation, the question mark should be replaced by **3. a.**

$$2Al + ?HCl \rightarrow 2AlCl_3 + 3H_2$$

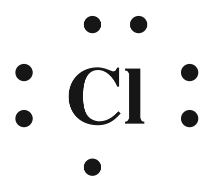
A. 1
B. 2
C. 3
D. 6

5. Max Planck's work with atoms correlated energy to **1. j.**
A. atomic size.
B. frequency.
C. ionic charge.
D. mass.

6. Gamma particles **11. e.**
A. penetrate less than alpha particles.
B. penetrate less than beta particles.
C. are the least penetrating of all radioactive decay.
D. are the most penetrating of all radioactive decay.

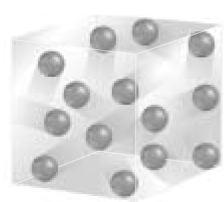
7. The rate of a reaction can be described using **8. a.**
A. evidence of reactant depletion.
B. evidence of product formation.
C. the decrease in the concentration of the reactants with time.
D. the decrease in the concentration of the products with time.

8. The Lewis dot structure for chlorine is shown below. What is the electron configuration for chlorine? **2. e.**



A. $1s^22s^22p^1$
B. $1s^22s^22p^63s^1$
C. $1s^22s^22p^63s^23p^1$
D. $1s^22s^22p^63s^23p^5$

9. In order to decrease the pressure in this container, **4. c.**



A. its volume must increase.
B. its temperature must increase.
C. its volume must decrease and its temperature must increase.
D. a lighter gas must be added.

Go on

Sample Test *(continued)*

10. A strong nuclear force holds protons and neutrons together in the atomic nucleus, yet atomic nuclei with too many protons and neutrons are unstable. Which statement correctly explains this? **11. a.**

A. The electromagnetic force is stronger than the nuclear force.
B. The weak nuclear force becomes more important for large nuclei.
C. The range of the strong nuclear force is too small for large nuclei.
D. The range of the electromagnetic force is too great for large nuclei.

11. A tank containing 1.7 atm of oxygen gas is combined with a tank containing 0.6 atm of oxygen gas and a tank containing 2.3 atm of oxygen gas. What is the final pressure in the tank? **4. i.**

A. 0.0 atm
B. 2.3 atm
C. 4.0 atm
D. 4.6 atm

12. Sodium hydroxide is a strong base. How will it dissociate in water? **5. c.**

A. It will not dissociate in water.
B. It will partially dissociate in water.
C. It will fully dissociate in water.
D. It will not mix with water.

13. The nucleus of an atom **1. e.**

A. is the largest part of an atom.
B. is the lightest part of an atom.
C. contains all of the atom's mass.
D. has a negative charge.

14. A basic solution usually feels **5. a.**

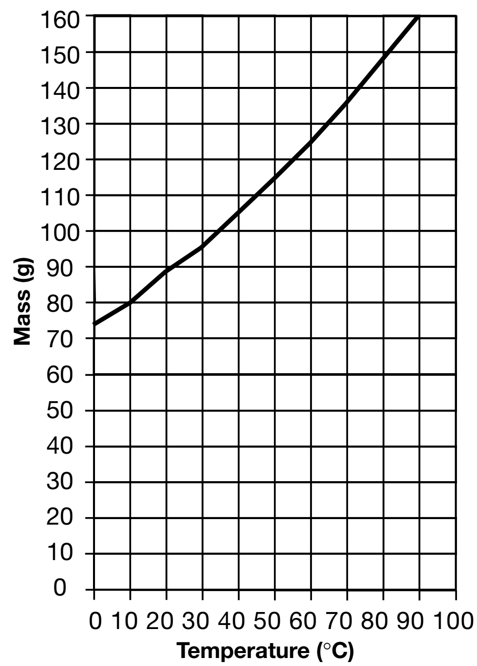
A. cold.
B. gritty.
C. soapy.
D. tingly.

15. Standard temperature and pressure (STP) occurs at **4. d.**

A. 1 atm.
B. 1 mm Hg.
C. 1 pound per square inch.
D. 1 torr.

16. According to these data, what is the approximate number of grams of sodium nitrate that can be dissolved at a temperature of 90°C? **6. c.**

Mass of Sodium Nitrate that can be Dissolved in 100 mL of Water



A. 150 g
B. 155 g
C. 160 g
D. 165 g

17. Higher temperatures will **8. b.**

A. decrease the amount of products in a chemical reaction.
B. decrease the concentration of products in a chemical reaction.
C. increase the reaction rate of a chemical reaction.
D. increase the activation energy of a chemical reaction.

18. What is 50°C in kelvins? **4. e.**

A. 223 K
B. 273 K
C. 323 K
D. 373 K

Go on

19. A ketone has a functional group of **10. e.**
  A. –CHO.
  B. –CO–.
  C. –COOH.
  D. –OH.

20. Use the ideal gas law below. What is the pressure produced by 2.0 mol of H_2 in a 22.4-L container at 273 K? Use R = 0.0821 (L•atm)/(mol•K). **4. h.**

$$PV = nRT$$

  A. 0.5 atm
  B. 2.0 atm
  C. 4.0 atm
  D. 8.0 atm

21. Hydrogen ions are donated by **5. b.**
  A. acids.
  B. bases.
  C. radioactive isotopes.
  D. salts.

22. After four half-lives have passed, how much of the original radioactive material will remain? **11. f.**
  A. 1/2
  B. 1/4
  C. 1/16
  D. 1/32

23. A solution is a uniform mixture. How are molecules distributed in a solution? **6. b.**
  A. Molecules are distributed by their random motion.
  B. Molecules are distributed based on their atomic mass.
  C. Molecules are distributed by ionic attraction.
  D. Molecules are distributed by electromagnetic attraction.

24. An Arrhenius base is a(n) **5. e.**
  A. hydroxide-ion producer.
  B. hydroxide-ion donor.
  C. ionic compound that dissociates into a metal and a hydroxide ion.
  D. substance that, in water, dissociates into hydrogen ions.

25. Henri Le Châtelier believed a new equilibrium after a change in a reaction would **9. a.**
  A. minimize the impact of the change.
  B. maximize the impact of the change.
  C. not occur in cases of a change in pressure.
  D. not occur in cases of a change in temperature.

26. A solution **6. a.**
  A. acts as a buffer.
  B. contains a dissolved substance.
  C. must contain a salt.
  D. reacts with a solvent.

27. Which form of radioactive decay consists of electrons? **11. d.**
  A. alpha
  B. beta
  C. delta
  D. gamma

28. How many atoms does 230 g of sodium (atomic weight = 23) have? **3. d.**
  A. 3.01×10^{23}
  B. 6.02×10^{23}
  C. 1.20×10^{24}
  D. 6.02×10^{24}

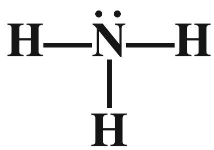
29. Which is true of liquids? **2. d.**
  A. Their molecules move less freely than those in solids.
  B. Their molecules move more freely than those in gases.
  C. The intermolecular forces between the atoms or molecules are weaker than those in solids.
  D. The intermolecular forces between the atoms or molecules are weaker than those in gases.

30. J.J. Thomson's experiments with cathode rays in magnetic and electric fields led to the discovery of **1. h.**
  A. electrons.
  B. neutrons.
  C. protons.
  D. quarks.

Go on

Sample Test *(continued)*

31. The Lewis dot structure for NH3 is shown below. What is the probable shape of NH3? **2. f.**



A. linear
B. pyramidal
C. tetrahedral
D. trigonal planar

32. Amino acids make up **10. c.**

A. DNA.
B. lipids.
C. proteins.
D. sugars.

33. Which of these is the correct way to express the equilibrium constant for the reaction below? **9. c.**

$$2HI(g) \rightleftharpoons H_2\,(g) + I_2\,(g)$$

A. $K = [HI] + [H_2] + [I_2]$
B. $K = [HI]^2 \times [H_2][I_2]$
C. $K = [HI] \div 2[H_2][I_2]$
D. $K = [H_2][I_2] \div [HI]^2$

34. What is the approximate molarity of 109.5 g of HCl (36.5 g/mole) in 750 L of solution? **6. d.**

A. 3*M*
B. 4*M*
C. 6*M*
D. 12*M*

35. Which temperature is unrealistic? **4. f.**

A. –100° C
B. –100° F
C. –100° K
D. 100 K

36. Which statement most accurately describes the state of a reaction after it has reached chemical equilibrium? **9. b.**

A. At chemical equilibrium, equal amounts of products and reactants are present.
B. The forward and reverse reaction rates are producing equal concentrations.
C. The forward and reverse reaction rates are occurring at equal rates.
D. At equilibrium, the reaction is continuing in either the forward or reverse direction.

37. In the reduction reaction shown below, how many electrons are needed to balance the equation? **3. g.**

$$Cl_2 + ?e^- \rightarrow 2Cl^-$$

A. 0
B. 1
C. 2
D. 4

38. Which element sets the standard for the quantity of a mole? **3. b.**

A. carbon
B. hydrogen
C. sodium
D. uranium

39. The diagram below shows a chemical equation representing a chemical reaction. The name and mass of each substance involved in the chemical reaction are also shown. What mass of sodium chloride was produced in this reaction? **3. e.**

36.5 g		40.0 g		? g		18.0 g
HCl	**+**	**NaOH**	**→**	**NaCl**	**+**	**H_2O**
hydrochloric acid		sodium hydroxide		sodium chloride		water

A. 24.0 g
B. 35.5 g
C. 48.0 g
D. 58.5 g

Go on

40. According to the kinetic theory of gases, which has the least influence on gas molecules? **4. g.**
- **A.** molecular attraction
- **B.** volume of the container
- **C.** temperature
- **D.** velocity

41. What holds the atoms of a salt in a fixed pattern? **2. c.**
- **A.** covalent bonds
- **B.** electrostatic attraction
- **C.** magnetic force
- **D.** viscosity

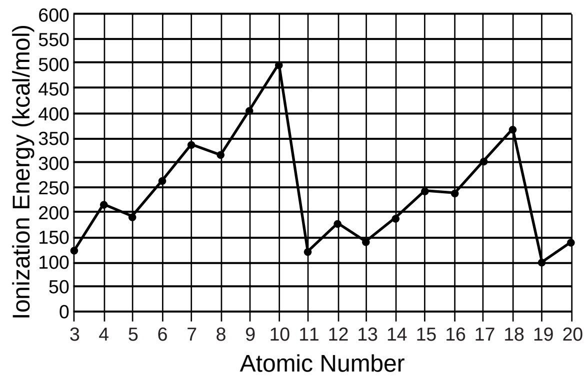
42. Which of these increases as the amount of solute particles in a solution increases? **6. e.**
- **A.** atomic number
- **B.** boiling point
- **C.** electronegativity
- **D.** freezing point

43. When an ice cube melts, energy **7. c.**
- **A.** remains unchanged.
- **B.** is transferred to a catalyst.
- **C.** is absorbed by the ice.
- **D.** is released by the water.

44. Perfume diffuses through the air because the molecules of a gas **4. b.**
- **A.** produce pressure.
- **B.** have random motion.
- **C.** vibrate quickly.
- **D.** are very small.

45. Uranium, located in the last row (elements 90 through 103) of the two rows normally shown as an insert of the periodic table, is a(n) **1. f.**
- **A.** actinide element.
- **B.** alkali metal.
- **C.** alkaline metal.
- **D.** lanthanide element.

Use the graph below to answer question 46.



46. Elements with atomic numbers 4, 12, and 20 are in the same group in the periodic table. As you move down a group, **1. a.**
- **A.** the principal energy level increases and the first ionization energy increases.
- **B.** the principal energy level increases and the first ionization energy decreases.
- **C.** the principal energy level decreases and the first ionization energy increases.
- **D.** the principal energy level decreases and the first ionization energy decreases.

47. An isotope of uranium is most likely to be **11. c.**
- **A.** inert.
- **B.** radioactive.
- **C.** of a lower atomic number than uranium.
- **D.** of a higher atomic number than uranium.

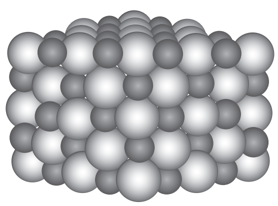
48. Refer to the periodic table on page 56 to answer this question. In the periodic table, the halogens are located in column (group) **1. b.**
- **A.** 1A.
- **B.** 2A.
- **C.** 7A.
- **D.** 8A.

Go on

49. The periodic table is organized into blocks representing the energy sublevel being filled with valence electrons. In the periodic table, which block represents the p sublevel? **1. g.**
A. W
B. X
C. Y
D. Z

50. Why is carbon found in so many molecules? **10. b.**
A. It has versatile bonding characteristics.
B. It has the highest electronegativity of all elements.
C. It has a unique atomic number.
D. Its atomic mass makes it ideal to bond with.

51. Which statement best describes what happens to sodium and chlorine atoms when they combine to form sodium chloride? **2. a.**



A. The sodium atom becomes a positive ion, and the chlorine atom becomes a negative ion.
B. The sodium atom becomes a negative ion, and the chlorine atom becomes a positive ion.
C. The sodium atom becomes a positive ion, and the chlorine atom becomes a positive ion.
D. The sodium atom becomes a negative ion, and the chlorine atom becomes a negative ion.

52. How many molecules does 2 mol of NaOH have? **3. c.**
A. 3.01×10^{23} molecules
B. 6.02×10^{23} molecules
C. 1.20×10^{24} molecules
D. 6.02×10^{24} molecules

53. Nucleic acids are composed of many **10. a.**
A. amino acids.
B. nucleotides.
C. monosaccharides.
D. triglycerides.

54. Which identifies an amino acid? **10. f.**
A. N group
B. R group
C. C terminal
D. N terminal

55. How many electrons do the elements in the first column of the periodic table have available for bonding? **1. d.**
A. 0
B. 1
C. 7
D. 8

56. In chromatography, a solute is separated from a solvent by **6. f.**
A. boiling off the solvent.
B. siphoning off the solvent.
C. electrostatic attraction.
D. viscosity.

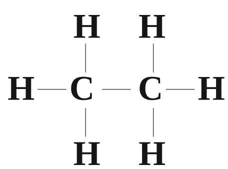
57. How does a catalyst affect reaction rates? **8. c.**
A. It decreases the reaction rate of a chemical reaction.
B. It increases the reaction rate of a chemical reaction.
C. It shifts the equilibrium constant of a chemical reaction to the left.
D. It shifts the equilibrium constant of a chemical reaction to the right.

58. When two atoms of the same kind (and, thus, with the same electronegativity) bond, **2. g.**
A. a hydrogen bond is formed.
B. an ionic bond is formed.
C. a nonpolar covalent bond is formed.
D. a polar covalent bond is formed.

Go on

59. Which is true of nuclear fusion? **11. b.**
A. It releases hydrogen ions.
B. It requires a low initial temperature.
C. It releases a great amount of energy.
D. It occurs regularly on Earth.

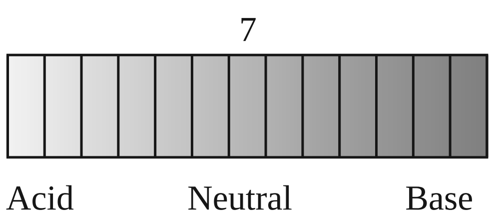
60. The compound shown below is known as **10. d.**



A. ethane.
B. methane.
C. ethene.
D. methene.

61. How do hydrogen bonds contribute to the high boiling-point temperature of water? **2. h.**
A. The bonds give up electrons.
B. The bonds cause protons to lose part of their positive charge.
C. The bonds cause a shorter distance between nuclei of hydrogen and oxygen.
D. The bonds are identical to ionic bonds.

62. Buttermilk has a pH of about 4.6. This makes buttermilk a **5. d.**



A. strong acid.
B. strong base.
C. weak acid.
D. weak base.

63. A reaction has a positive change in free energy. This reaction will **7. f.**
A. be spontaneous.
B. will never reach equilibrium.
C. will occur at 0°C.
D. need energy to occur.

64. According to the graph, which element has the strongest attraction for electrons? **1. c.**
A. beryllium (atomic number = 4)
B. calcium (atomic number = 20)
C. magnesium (atomic number = 12)
D. sodium (atomic number = 11)

65. Which is a substructure of a neutron? **11. g.**
A. electron
B. isotope
C. proton
D. quark

66. How much heat in joules does it take to raise 100 g of water from 25°C to 30°C? (The specific heat of water is 4.180 J/g•˚C.) **7. d.**

$$q = mc\Delta T$$

A. 20.9 J
B. 209 J
C. 2090 J
D. 20,900 J

67. When temperature rises, **7. a.**
A. collisions occur less often.
B. energy decreases.
C. entropy decreases.
D. molecules move faster.

68. How does blood maintain a pH of about 7.4? **5. g.**
A. It is a weak acid.
B. It contains buffers.
C. It contains sugar.
D. It contains red blood cells.

Go on

69. Which is an exothermic reaction? **7. b.**
- **A.** building a chemical bond
- **B.** making a protein
- **C.** oxidation
- **D.** reduction

70. Which is an example of activation energy? **8. d.**
- **A.** energy required to split a nucleus
- **B.** energy required to start a chemical reaction
- **C.** energy required to buffer a solution
- **D.** energy required to initiate radioactive decay

71. Niels Bohr's work with the atom led him to speculate about **1. i.**
- **A.** isotopes.
- **B.** protons.
- **C.** spectral lines.
- **D.** types of bonds.

72. In a chemical reaction, a yield of 1000 g of a substance is predicted based on the amount of reactants. Only 979 g of the substance is produced. What is the percent yield? **3. f.**
- **A.** 2 percent
- **B.** 21 percent
- **C.** 98 percent
- **D.** 979 percent

73. Which demonstrates the collisions of gas molecules? **4. a.**
- **A.** mass of a tank of gas on a balance
- **B.** movement of dye in a glass of water
- **C.** pressure on the inside walls of a balloon
- **D.** smell of a skunk in the air

Stop

Periodic Table of the Elements

Columns of elements are called groups. Elements in the same group have similar chemical properties.

Element — Hydrogen
Atomic number — 1
Symbol — H
Atomic mass — 1.008
State of matter

Gas
Liquid
Solid
Synthetic

The first three symbols tell you the state of matter of the element at room temperature. The fourth symbol identifies elements that are not present in significant amounts on Earth. Useful amounts are made synthetically.

Metal
Metalloid
Nonmetal

The color of an element's block tells you if the element is a metal, nonmetal, or metalloid.

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	Hydrogen 1 H 1.008																	Helium 2 He 4.003
2	Lithium 3 Li 6.941	Beryllium 4 Be 9.012											Boron 5 B 10.811	Carbon 6 C 12.011	Nitrogen 7 N 14.007	Oxygen 8 O 15.999	Fluorine 9 F 18.998	Neon 10 Ne 20.180
3	Sodium 11 Na 22.990	Magnesium 12 Mg 24.305											Aluminum 13 Al 26.982	Silicon 14 Si 28.086	Phosphorus 15 P 30.974	Sulfur 16 S 32.065	Chlorine 17 Cl 35.453	Argon 18 Ar 39.948
4	Potassium 19 K 39.098	Calcium 20 Ca 40.078	Scandium 21 Sc 44.956	Titanium 22 Ti 47.867	Vanadium 23 V 50.942	Chromium 24 Cr 51.996	Manganese 25 Mn 54.938	Iron 26 Fe 55.845	Cobalt 27 Co 58.933	Nickel 28 Ni 58.693	Copper 29 Cu 63.546	Zinc 30 Zn 65.409	Gallium 31 Ga 69.723	Germanium 32 Ge 72.64	Arsenic 33 As 74.922	Selenium 34 Se 78.96	Bromine 35 Br 79.904	Krypton 36 Kr 83.798
5	Rubidium 37 Rb 85.468	Strontium 38 Sr 87.62	Yttrium 39 Y 88.906	Zirconium 40 Zr 91.224	Niobium 41 Nb 92.906	Molybdenum 42 Mo 95.94	Technetium 43 Tc (98)	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.906	Palladium 46 Pd 106.42	Silver 47 Ag 107.868	Cadmium 48 Cd 112.411	Indium 49 In 114.818	Tin 50 Sn 118.710	Antimony 51 Sb 121.760	Tellurium 52 Te 127.60	Iodine 53 I 126.904	Xenon 54 Xe 131.293
6	Cesium 55 Cs 132.905	Barium 56 Ba 137.327	Lanthanum 57 La 138.906	Hafnium 72 Hf 178.49	Tantalum 73 Ta 180.948	Tungsten 74 W 183.84	Rhenium 75 Re 186.207	Osmium 76 Os 190.23	Iridium 77 Ir 192.217	Platinum 78 Pt 195.078	Gold 79 Au 196.967	Mercury 80 Hg 200.59	Thallium 81 Tl 204.383	Lead 82 Pb 207.2	Bismuth 83 Bi 208.980	Polonium 84 Po (209)	Astatine 85 At (210)	Radon 86 Rn (222)
7	Francium 87 Fr (223)	Radium 88 Ra (226)	Actinium 89 Ac (227)	Rutherfordium 104 Rf (261)	Dubnium 105 Db (262)	Seaborgium 106 Sg (266)	Bohrium 107 Bh (264)	Hassium 108 Hs (277)	Meitnerium 109 Mt (268)	Darmstadtium 110 Ds (281)	Roentgenium 111 Rg (272)	Ununbium * 112 Uub (285)		Ununquadium * 114 Uuq (289)				

The number in parentheses is the mass number of the longest-lived isotope for that element.

* The names and symbols for elements 112–114 are temporary. Final names will be selected when the elements' discoveries are verified.

Rows of elements are called periods. Atomic number increases across a period.

The arrow shows where these elements would fit into the periodic table. They are placed below the table to save space.

Lanthanide series	Cerium 58 Ce 140.116	Praseodymium 59 Pr 140.908	Neodymium 60 Nd 144.24	Promethium 61 Pm (145)	Samarium 62 Sm 150.36	Europium 63 Eu 151.964	Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.925	Dysprosium 66 Dy 162.500	Holmium 67 Ho 164.930	Erbium 68 Er 167.259	Thulium 69 Tm 168.934	Ytterbium 70 Yb 173.04	Lutetium 71 Lu 174.967
Actinide series	Thorium 90 Th 232.038	Protactinium 91 Pa 231.036	Uranium 92 U 238.029	Neptunium 93 Np (237)	Plutonium 94 Pu (244)	Americium 95 Am (243)	Curium 96 Cm (247)	Berkelium 97 Bk (247)	Californium 98 Cf (251)	Einsteinium 99 Es (252)	Fermium 100 Fm (257)	Mendelevium 101 Md (258)	Nobelium 102 No (259)	Lawrencium 103 Lr (262)