Chapter 1 and 2 Homework Answer Key

1.3 pg. 17 #8 to 11

#8) Answers will vary. Sample Answer:
 Observation: My Stomach is Growling!
 Hypothesis: Maybe I am hungry!
 Experiment: Eat something and see what happens!

#9) A law is a concise statement that summarizes the results of a broad variety of observations and experiments. A theory is a tested (by experiments) hypothesis. A hypothesis is a proposed explanation or reason for what is observed.

#10) The hypothesis guides the design of the experiment.

#11) a law; The statement summarizes facts; it does not give an explanation

2.1 pg. 31 #1 to 4

#1) No; a substance is a particular kind of matter that has a uniform and definite composition. Pure substances contain only one kind of matter. Matter is anything that has mass and occupies space.

#2) Solids have definite shape and volume and are nearly incompressible. Liquids have definite volume but no definite shape, are nearly incompressible, and can flow. Gases have neither definite shape nor definite volume and are easily compressed.

#3) b., c., and d.

#4) a. Mercury b. Melting point and Density c. All are colorless

2.2 pg. 34 #5, 6; pg. 35 #7 to 12

#5) Magnet attracts iron, but not salt. Salt dissolves in water, but not iron.

#6) a. Heterogeneous b. Homogeneous c. Heterogeneous

d. Homogeneous e. Homogeneous

#7) Heterogeneous mixtures have non-uniform composition consisting of two or more phases. Homogeneous mixtures have a uniform composition throughout the sample.

#8) Add water to dissolve the salt. Pour the resulting mixture onto a piece of closely-woven cloth. The sand will remain on the cloth, and the salt solution will pass through. Use evaporation to remove the water from the solution, leaving solid salt behind.

#9) a. Substance b. Mixture c. mixture d. Substance

#10) A pure substance contains only one kind of matter. A mixture contains tow or more kinds of matter that may or may not be uniform in composition.

#11) The components of a mixture can be separated by physical means such as filtration or distillation.

#12) A phase is any part of a system with uniform composition. A homogeneous mixture consists of one phase; a heterogeneous mixture consists of two or more phases.

2.3 pg. 39 #13; pg. 40 #14 to 18

#13) It cannot be an element because it separated into at least one solid and one liquid. It must have been a mixture. The liquid components evaporated, leaving behind the solid components. Compounds cannot be separated by physical means such as evaporation. Evaporation is one way to physically separate the components of a mixture.

#14) Compounds can be separated by chemical means into elements. Elements cannot be separated into simpler substances by chemical techniques.

#15)	a. C	u	b. O	C.	Р	d.	Ag	e.	Na	f. He
#16)	a. Ti	in		b.	Calcium			C.	Sulfur	
(d. C	admium		e.	Phosphorus	S		f.	Chlorine	
#17)	a. M	lixture		b.	Mixture			C.	Compound	
(d. M	lixture		e.	Mixture			f.	Element	

#18) carbon, hydrogen, oxygen, and nitrogen; Hydrogen is present in the greatest proportion by number of atoms.

2.4 pg. 43 #19 to 23

#19) a. In a chemical change, the chemical composition of the reactants changes as one or more different products is formed. In a physical change, the chemical composition of the substance remains the same even if its physical appearance changes. Indicators are (1) a change in color or odor, or production of a gas; (2) energy released or absorbed; (3) irreversibility. Out of these indicators, (3) is the most reliable observation that a chemical change had taken place.

b. In any physical change or chemical reaction, mass is neither created nor destroyed; it is conserved. The mass of the products equals the mass of the reactants in a chemical reaction.

20) a.	Chemical	b.	Physical	C.	Physical
d.	Chemical	e.	Chemical	f.	Chemical

21) 18 g

- 22) a. Color, odor, reaction upon heating, boiling point
 - b. Color, melting point, reactions with other substances, hardness, brittleness, strength
 - c. boiling point, freezing point, density
 - d. density, melting point, magnesium

23) 43.2 g

Chapter 2 Review pg. 47-48 #24 to 43

24) solid, metallic luster, gray color, high melting point, malleable

25) a.	Solid	b. Liquid	c. Gas	d. Solid	e. Liquid	f. Liquid
26) a.	Solid	b. Gas	c. Liquid	d. Liquid	e. Solid	f. Gas

27) vapor; The term "vapor" is used to refer to the gaseous state of a substance which normally exists as a liquid or a solid at room temperature.

28) water, gasoline, acetone (fingernail polish remover), aromatic vaporizers (such as aerosol deodorants)

29) chlorine, mercury, bromine, and water; Chlorine condenses, and mercury, bromine and water all freeze when the temperature drops within the stated range.

30) a. Heterogeneous	b. Heterogeneous	c. Homogeneous
d. Homogeneous	e. Homogeneous	

31) one; A solution is a system with uniform composition and properties. Solutions are homogeneous mixtures, consisting of a single phase.

32) a.	Element	b. Mixture	c. Mixture	e d. Element	e. Mixture	f. Mixture
33) а. с.	Nitrogen, hydro Carbon, hydrog	ogen chlorine gen, oxygen	b. d.	Potassium, mang Calcium, iodine	ganese, oxygen	
34) co	lor change; ene	ergy absorbed o	or released	; gas produced; o	dor change	
35) a.	Physical	b. Chemical	C.	Chemical	d. Physical	

36) The iron combines with oxygen in the air, and oxygen has mass

37) As the wax burns, the chemical composition of the wax changes, producing the products water and carbon dioxide, which are released into the surrounding air.

38) Add sufficient water to dissolve all of the sugar. Separate the charcoal and sand from the sugar water by filtration. Large pieces of charcoal could be separated on the basis of color. Small pieces of charcoal could be burned.

39) a.	Mixtures	b. Mix	ture	S				
40) a.	Color	b. Six		c. Sc	odiun	n chloride	d.	Sulfur
41) a. d. g.	Homogeneous Mixtur Homogeneous Mixtur Homogeneous Mixtur	e e e	b. e. h.	Homogeneous M Heterogeneous N Heterogeneous N	lixtur ⁄lixtu ⁄lixtu	re c. Heterogene ire f. Compound ire	ou	s Mixture
42) a.	Physical b. Phy	/sical	C.	Physical	d.	Physical	e.	Chemical
43) a. d.	Color and odor chang color and texture char	e ige	b. e.	Gas is produced Energy change, o	odor,	c. formation o , irreversible	fa	precipitate
44) a.	(1) product	b. (3)	com	pound				

45) In gases, particles are far apart. In liquids, particles are in contact. In solids, particles are tightly packed

46) The appearance of a substance will change during a change of state, which is a physical change.

47) a. Yes, because the graph is a straight line, the proportion of iron to oxygen is a constant.

b. No, plotting these values on the graph would not give a point on the line indicating that the mass ratio of iron to oxygen is different from the other four samples.

48) a. Oxygen and calcium b. Silicon, aluminum, and iron

c. Different. The second most abundant element in the Earth's crust, silicon, is not present in the human body, and the second most abundant element in the human body, carbon, is not among the most abundant elements of Earth's crust. If the elements are different then the compounds must also be different.

3-3 to 3-5 Homework Answer Key

3.3 pg. 67 #17 to 22

#17) a.	Amount of substance, mol	b. Density, kg/m ³	C.	time, s
d.	Pressure, Pa	e. Length, m	f.	Mass, kg

#18) Mass is a measure of the amount of matter in an object. Weight is a measure of the force of gravity on an object. Weight is a measure of the force of gravity on an object.

#19) a. (m), 10^{-3} b. (n), 10^{-9} c. (d), 10^{-1} d. (c), 10^{-2}

#20) Your weight would decrease; your mass would remain constant.

#21) 8.8 x 10² cm³

#22) a. and d. > f. > e. > c. > b.

3.4 pg. 71 #23, 24; pg. 72 #25, 26, 28

#23) 2.50 g/cm³, no

#24) 6.5 cm³

#25) Mass is divided by volume.

#26) a. 1.7 x 10⁻² g/L

#28) 0.802 g/cm³; It would sink.

3.5 pg. 75 #30 to 35

#30) - 196°C

#31) melting point: 1234 K; boiling point: 2485 K

#32) °C = K - 273

#33) 463 K

#34) 443 K

#35) - 186°C

Chapter 3

Answers

- 36. a. qualitative
 - b. quantitative
 - **c.** qualitative
 - d. quantitative
- 37. a. precision
 - **b**. accuracy
 - C. precision
 - **d**. precision
 - e. accuracy
 - f. accuracy
- 38. when using an improperly calibrated measuring device
- 39. Lissa: inaccurate and imprecise Lamont: accurate and precise Leigh Anne: inaccurate but precise
- 40. a. accurate and precise **b**. inaccurate but precise c. inaccurate and imprecise
- 41. a. infinite
 - **b**. infinite c. infinite
 - **d**. 4
 - e. infinite
 - f. 3
- 42. a. 98.5 L
 - **b.** 0.000 763 cg **c.** 57.0 m
 - d. 12.2°C
 - **e.** 0.007 50 \times 10⁴ mm
 - f. 1760 mL
- **43. a.** 9.85×10^{1} L **b.** 7.63×10^{-4} cg
- **c.** 5.70×10^{1} m
- **d.** 1.22×10^{1} °C
- **e.** 7.50×10^{1} mm
- f. 1.76×10^3 mL
- 44. a. 43 g
 - **b.** $7.3 \, \mathrm{cm}^2$ C. 225.8 L **d**. 92.0 kg
 - **e.** 32.4 m³
 - f. 104 m³
- **45. a.** 4.3×10^{1} g **b**. $7.3 \times 10^{\circ} \text{ cm}^2$ **c.** 2.258×10^{2} L
 - **d.** 9.20×10^{1} kg
 - **e.** $3.24 \times 10^{1} \text{ m}^{3}$
 - f. $1.04 \times 10^2 \,\mathrm{m}^3$

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46. The error is the difference between the accepted and the experimental values. The percent

Chapter 3 REVIEW

CONCEPT PRACTICE

- 36. Identify the following as quantitative or qualitative measurements. 3.1
- a. A flame is hot.
- b. A candle has a mass of 90 g.
- c. Wax is soft.
- d. A candle's height decreases 4.2 cm/hr.
- 37. Which of these synonyms or characteristics apply to the concept of accuracy? Which apply to the concept of precision? 3.2
 - a. multiple measurements
 - b. correct
 - c. repeatable
 - d. reproducible
 - e. single measurement
 - f. true value
- 38. Under what circumstances could a series of measurements of the same quantity be precise but inaccurate? 3.2
- **39**. Three students made multiple weighings of a
 - copper cylinder, each using a different balance.
- The correct mass of the cylinder had been previously determined to be 47.32 g. Describe the accuracy and precision of each student's measurements. 3.2

Mass of Cylinder (g)							
Lissa Lamont Leigh Anne							
Weighing 1	47.13	47.45	47.95				
Weighing 2	47.94	47.39	47.91				
Weighing 3	46.83	47.42	47.89				
Weighing 4	47.47	47.41	47.93				

- 40. Comment on the accuracy and precision of these basketball free-throw shooters. 3.2 a. 99 of 100 shots are made.
 - **b**. 99 of 100 shots hit the front of the rim and bounce off.
 - c. 33 of 100 shots are made: the rest miss.
- 41. How many significant figures are in each underlined measurement? 3.2
 - **a**. <u>60 s</u> = 1 min
 - b. 9 innings in a baseball game
 - **c.** 1 km = 1000 m
 - d. 47.70 g of copper
 - e. 25 computers
 - f. 0.0950 m of gold chain

error is the absolute value of the error divided by the accepted value multiplied by 100%.

- 47. The absolute value of the error is used.
- 48.4%
- 49. 23.9 g
- 50. Possible answers are: Units are based on multiples of ten; prefixes have the same meaning when attached to different units of measure.
- 51. a. second, s b. meter, m

- 42. Round each of these measurements to three significant figures. 3.2
 - d. 12.17 °C a. 98.473 L
 - **e**. 0.007 498 3×10^4 mm b. 0.000 763 21 cg f. 1764.9 mL
- c. 57.048 m
- 43. Write each of the rounded measurements in Problem 42 in scientific notation. 3.2
- 44. Round each of the answers correctly. 3.2
 - **a.** 8.7 g + 15.43 g + 19 g = 43.13 g
 - **b.** 4.32 cm \times 1.7 cm = 7.344 cm²
 - c. 853.2 L 627.443 L = 225.757 L **d.** $38.742 \text{ kg} \div 0.421 = 92.02375 \text{ kg}$
 - **e.** 5.40 m × 3.21 m × 1.871 m = 32.431 914 m³
 - f. $5.47 \text{ m}^3 + 11 \text{ m}^3 + 87.300 \text{ m}^3 = 103.770 \text{ m}^3$
- 45. Express each of the rounded answers in Problem 44 in scientific notation. 3.2
- 46. How are the error and the percent error of a measurement calculated? 3.2
- 47. Why is the percent error of a measurement always positive? 3.2
- 48. A student estimated the volume of a liquid in a beaker as 200 mL. When she poured the liquid into a graduated cylinder, she measured the volume as 208 mL. What is the percent error of the estimated volume from the beaker, taking the measurement in the graduated cylinder as the accepted value? 3.2
- 49. Water with a mass of 35.4 g is added to an empty flask with a mass of 87.432 g. The mass of the flask and the water is 146.72 g after a rubber stopper is added. Express the mass of the stopper to the correct number of significant figures. 3.2
- 50. List at least two advantages of using SI units for measuring. 3.3
- 51. List the SI base unit of measurement for each of these quantities. 3.3
 - a. time c. temperature
 - b. length d. mass
- 52. Use the tables in this chapter to order these lengths from smallest to largest. Give each measurement in terms of meters, 3.3
 - a. centimeter e. meter
 - **b**. micrometer f. nanometer
 - c. kilometer g. decimeter d. millimeter

C. kelvin, K

d. kilogram, kg

kilometer (10³ m)

h. picometer

52. picometer (10^{-12} m) , nanometer (10^{-9} m) , mi-

crometer (10^{-6} m) , millimeter (10^{-3} m) , cen-

timeter (10^{-2} m) , decimeter (10^{-1} m) , meter,

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- Measure each dimension using a unit with the appropriate prefix. 3.3
 - a. the height of this letter: I
 - b. the width of Table 3.3
 - the height of this page
 - om what unit is a measure of volume rived? 3.3
 - bat is the volume of a glass cylinder with an le diameter of 6.0 cm and a height of Cm ? (The volume of a cylinder equals pi \times bus squared × height, or $V = 3.14 r^2 h$.) 3.3

Match he approximate volume with each item.

3.3	
a. orange	(1) 30 m^3
b. basketball	(2) 200 cm ³
c. van	(3) 20 L
d. aspirin tablet	(4) 200 mm ³
-	

How many grams are in each of these quantities? 3.3

c. 1 kg a.1cg b.1µg d. 1 mg

- Astronauts in space are said to have apparent weightlessness. Explain why it is incorrect to say that they are massless. 3.3
- 59. Match the approximate mass with each item.

3.3	
a. peanut	(1) 400 cg
b. pear	(2) 50 mg
c. stamp	(3) 60 kg

- d. person (4) 150 g
- 60. Would the density of a person be the same on the surface of Earth as it is on the surface of the moon? Explain. 3.4
- 61. A shiny, gold-colored bar of metal weighing 57.3 g has a volume of 4.7 cm³. Is the metal bar pure gold? 3.4
- 62. Why doesn't a measure of specific gravity have a unit? 3.4
- Use the values in Table 3.7 to calculate the specific gravity of the following substances at 20 °C. 3.4
 - a. aluminum
 - b. mercury
- c. ice
- Three balloons filled with neon, carbon dioxide, and hydrogen are released into the atmosphere. Using the data in Table 3.7, describe the movement of each balloon. 3.4

65. Which would melt first, germanium with a melting point of 1210 K or gold with a melting point of 1064 °C? 3.5

CONCEPT MASTERY

- 66. List two possible reasons for precise, but inaccurate, measurements.
- 67. Rank these numbers from smallest to largest.

$\mathbf{a}.5.3 imes10^4$	d . 0.0057
b . $57 imes 10^3$	e . 5.1×10^{-3}
c. 4.9×10^{-2}	f. 0.0072×10^2

- 68. Criticize this statement: "When two measurements are added together, the answer can have no more significant figures than the measurement with the least number of significant figures."
- 69. Fahrenheit is a third temperature scale. Use the data in the table or the graph to derive an equation for the relationship between the Fahrenheit and Celsius temperature scales.

	°C	°F
Melting point of selenium	221	430
Boiling point of water	100	212
Normal body temperature	37	98.6
Freezing point of water	0	32
Boiling point of chlorine	-34.6	-30.2



- 70. Which is larger?
 - a. 1 centigram or 1 milligram
 - b. 1 liter or 1 centiliter
 - c. 1 calorie or 1 kilocalorie
 - d. 1 millisecond or 1 centisecond
 - e. 1 microliter or 1 milliliter
 - f. 1 cubic millimeter or 1 cubic decimeter

Scientific Measurement 79

53. a. 2.4 mm

- **b**. 17.6 cm
- c. 27.6 cm
- 54. Volume is a length unit cubed.
- 55. a. 7.9×10^2 cm³
- 56. a. 2
 - b. 3

 - C. 1
 - d. 4

- 57. a. 0.01 g **b.** 0.000 001 g
 - **c.** 1000 g
 - **d.** 0.001 g
- 58. The mass of an object is constant. The weight of an object varies with location.
- 59. a. 1
 - **b**. 4
 - **C.** 2
 - **d**. 3

Chapter 3

- 60. Yes, neither mass nor volume changes with location.
- 61. No; the density of the metal bar is 12 g/cm³, but the density of gold is 19 g/cm³.
- 62. Specific gravity is a ratio of two density measurements, so the density units cancel.
- 63. a. 2.70
 - **b.** 13.6
 - **C.** 0.917
- 64. The carbon dioxide-filled balloon would sink. The neon- and hydrogen-filled balloons would rise, the hydrogen at a much faster rate.
- 65. germanium
- 66. improper calibration or improper use of the measuring device
- 67. e., d., c., f., a., b.
- 68. Significant figures in the answer of an addition problem depend on the measurement with the least number of decimal places.
- **69.** °F = 1.8 °C + 32
- 70. a. cg
 - b. L
 - c. kcal
 - d. cs
 - e.mL
 - f. 1 dm³

5.1 and 5.2 Homework Answer Key

5.1 & 5.2 pg. 108 #1 to 3 & pg. 112 #4 to 6

#1) Answers will vary but should include the ideas that all matter is composed of atoms. Atoms of different elements differ, and chemical change involves a rearrangement of atoms.

#2) Answers will vary but should emphasize the smallness of atoms.

#3) Dalton used reasoning based on the results of scientific experiments. Democritus used mental reasoning only.

#4) proton, positive charge, relative mass = 1; electron, negative charge, relative mass = 1/1840; neutron, no charge, relative mass =1

#5) An atom has a central core composed of protons and neutrons, called the nucleus. Electrons surround the nucleus and occupy most of the volume of the atom.

#6) Thomson passed electric current through sealed glass tubes filled with gases. The resulting glowing beam was described as a stream of tiny negatively charged particles moving at high speed. He concluded that electrons must be parts of the atoms of all elements. Milliken determined the charge and mass of the electron. Rutherford's gold-foil experiments indicated that the atom had a positively charged, dense nucleus which is tiny compared to the atom as a whole.

5.3 and 5.4 Homework Answer Key

5.3 pg. 115 #7, 8; pg. 116 #9 to 11; pg. 117 #12, 13; pg. 120 #14, 15; pg. 121 #16 to 26

#7) a.	9 protons and 9 electrons	b.	13 protons and 13	electrons
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c. 20 protons and 20 electrons

#8) K - 19 protons; B - Atomic Number 5 and 5 protons; S - 16 protons and 16 electrons; V - Atomic Number 23 and 23 electrons

#9) a. 8	b. 16	c. 61	d. 45	e. 125
#10) a. ¹² ₆ C	b. ¹⁹ ₉ F	c. ⁹ ₄ Be		
#11)a.8	b. 16	c. 47	d. 35	e. 82
#12) ¹⁶ ₈ O ¹⁷ ₈ O	¹⁸ 8 O			
#13) 26, 28, 29				
#14) boron-11				

#15) Silicon-28 must be by far the most abundant. The others must be present in only small amounts.

#16) 63.5 amu

#17) 79.90 amu

#18) All atoms of an element have the same number of protons even though the number of neutrons may vary. The protons and electrons are responsible for the chemical behavior of atoms. In a neutral atom, both the number of protons and the number of electrons equal the atomic number.

#19) In a neutral atom, the atomic number is the number of protons and the number of electrons. The mass number is the total number of protons plus neutrons. To find the number of neutrons, subtract the atomic number from the mass number.

#20) a. Mass number b. ¹⁹⁵/₇₈ Pt

#21) Isotopes of the same element have identical numbers of protons; they have different masses, mass numbers, and numbers of neutrons.

#22)) zinc-64: 30 p⁺, 30 e⁻, 34 n zinc-68: 30 p⁺, 30 e⁻, 38 n	zinc-66: 30 p ⁺ , 30 e ⁻ , 36 n zinc-70: 30 p ⁺ , 30 e ⁻ , 40 n	zinc-67: 30 p⁺, 30 e⁻, 37 n
#23)) a. lithium-6: 3 p⁺, 3 e⁻, 3 n	lithium-7: 3 p⁺, 3 e⁻,	4 n
	b. calcium-42: 20 p⁺, 20 e⁻, 22	n calcium-44: 20 p⁺, 2	0 e⁻, 24 n
	c. selenium-78: 34 p⁺, 34 e⁻, 44	4 n selenium-80: 34 p⁺,	34 e⁻, 46 n

#24) The atomic mass is the weighted average of the masses of its isotopes.

#25) The mass of the isotopes in a sample of the element are averaged based on relative abundance. The result is the element's atomic mass.

#26) ${}^{14}_{7}$ N - 14.003 amu at 99.63% ${}^{15}_{7}$ N - 15.000 amu at 0.37% Average Atomic mass = 14.01 amu

5.4 pg. 126 #28 to 30, 32

#28) increasing atomic mass and similarities of properties

#29) A group is a vertical column. A period is a horizontal row. A transition metal is a Group B element.

#30) a. Metal b. Metalloid c. Metal d. Non-mental e. Metal

#32) beryllium, magnesium, strontium, barium

Chapter 5 REVIEW

CONCEPT PRACTICE

- **33**. With which of these statements would John Dalton have agreed in the early 1800s? For each, explain why or why not? *5.1*
 - a. Atoms are the smallest particles of matter.
 - **b**. The mass of an iron atom is different from the mass of a copper atom.
 - c. Every atom of silver is identical to every other atom of silver.
 - d. A compound is composed of atoms of two or more different elements.
- 34. What experimental evidence did Thomson have for each statement? 5.2
 - a. Electrons have a negative charge.
 - b. Atoms of all elements contain electrons.
- 35. Would you expect two electrons to attract or repel each other? 5.2
- **36**. How did the results of Rutherford's gold foil experiment differ from his expectations? *5.2*
- **37**. What is the charge, positive or negative, of the nucleus of every atom? 5.2
- 38. Why is an atom electrically neutral? 5.3
- **39**. What does the atomic number of each atom represent? 5.3
- **40**. How many protons are in the nuclei of the following atoms? 5.3
 - a. phosphorus d. cadmium
 - **b**. molybdenum **e**. chromium
 - c. aluminum f. lead
- **41**. What is the difference between the mass number and the atomic number of an atom? *5.3*

42. Complete this table. 5.3

Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons	Symbol of element
9			10		
		14	15		
	47		25		
	55	25			

- **43**. Name two ways that isotopes of an element differ. *5.3*
- **44**. How can there be more than 1000 different atoms when there are only about 100 different elements? *5.3*

- **45**. What data must you know about the isotopes of an element to calculate the atomic mass of the element? 5.3
- 46. What is the atomic mass of an element? 5.3
- **47.** Look up the word *periodic* in the dictionary. Propose a reason for the naming of the periodic table. *5.4*
- **48**. How did Moseley's arrangement of the elements differ from that of Mendeleev? 5.4
- 49. Give the symbol of each element. 5.4
 - a. any nonmetal in Group 4A
 - **b**. the inner transition metal with the lowest atomic number
 - **c**. all of the nonmetals for which the atomic number is a multiple of five
 - **d**. the two elements that are liquid at room temperature
 - e. any metal in Group 5A

CONCEPT MASTERY

- **50**. Compare the relative size and relative density of an atom with its nucleus.
- 51. Imagine you are standing on the top of a boron-11 nucleus. Describe the numbers and kinds of subatomic particles you would see looking down into the nucleus, and those you would see looking out from the nucleus.
- **52**. What parts of Dalton's atomic theory no longer agree with the current picture of the atom?
- **53.** The four isotopes of lead are shown below, each with its percent by mass abundance and the composition of its nucleus. Using these data, calculate the approximate atomic mass of lead.

	1997		
F.1.4	8'-	820	82p
		125	1260
1.37%	25 26%	20.82%	51.55%

- 54. Datton's atomic theory was not correct in every detail. Should this be taken as a criticism of Dalton as a scientist? Explain.
- **55.** Why are atoms considered the basic building blocks of matter even though smaller particles, such as protons and electrons, exist?

Atomic Structure and the Periodic Table 129

Answers

- **33.** Dalton would agree with all four statements because they all fit his atomic theory.
- **34. a.** A beam of electrons (cathode rays) is deflected by an electric field toward the positively charged plate.

b. The cathode rays were always composed of electrons regardless of the metal used in the electrodes or the gas used in the cathode-ray tube.

- 35. repel
- **36.** He did not expect any alpha particles to be deflected over a large angle.
- **37**. Every atomic nucleus is positively charged.
- **38.** It has equal numbers of protons and electrons.
- **39**. number of protons in the nucleus

40.	a. 15	-	d. 48
	b. 42		e. 24
	C. 13		f. 82

Chapter 5 REVIEW

41. The atomic number is the number of protons. The mass number is the sum of the number of protons and number of neutrons.

42.	9	19	9	10	9	F
	14	29	14	15	14	Si
	22	47	22	25	22	Ti
	25	55	25	30	25	Mn

- **43.** mass numbers, atomic masses, number of neutrons, relative abundance
- **44.** because of the existence of isotopes
- **45.** which isotopes exist, their masses, and their natural percent abundance
- **46.** The atomic mass is the weighted average of the masses of all of its isotopes.
- 47. Answers will vary.
- **48.** Moseley arranged the elements in order of increasing atomic number, not atomic mass.
- 49. a. C, Si d. Hg, Br b. La e. Bi, Sb c. B, Ne, P, Br
- **50.** The nucleus is very small and very dense compared with the atom.
- **51.** five protons and six neutrons in the nucleus; five electrons outside the nucleus
- **52.** All atoms of the same element are not identical (isotopes). The atom is not the smallest particle of matter.
- 53. 207 amu
- **54.** No; in general, he proposed a valid theory in line with the experimental evidence he had available to him.
- **55.** Atoms are the smallest particle of an element that retains the properties of that element.

28.1 and 28.2 Homework Answer Key

28.1 pg. 844 #1 to 3

#1) Radioactivity is the process by which an atomic nucleus gives off radiation. Radioactive decay is the process in which an unstable nucleus distingrates.

- #2) a. Mass: Alpha = 4 amu; Beta = 1/1837 amu; Gamma = 0 amu
 - b. Charge: Alpha = 2+; Beta = 1-; Gamma = 0
 - c. Penetrating Power: Alpha = low ; Beta = moderate ; Gamma = high

#3) The nucleus undergoes change.

28.2 pg. 844 #6, 8 and 10

#6) Because a half-life is the time required for one-half of the nuclei of a radioisotope to decay to products, use *x* to represent the half-lives that have passed. The $(\frac{1}{2})^x$ will represent the fraction of original isotope that remains.

#8) a.	$^{27}_{13}$ Al +	4_2 He $\rightarrow {}^{30}_{14}$ Si \cdot	+ ${}^{1}_{1}$ H
b.		$^{214}_{83}$ Bi $\rightarrow ^{4}_{2}$ He	+ ²¹⁰ ₈₁ TI
C.		$^{27}_{14}$ Si \rightarrow $^{0}_{-1}$ e	+ ²⁷ ₁₅ P
d.		$^{66}_{29}$ Cu \rightarrow $^{0}_{-1}$ e	+ $\frac{66}{30}$ Zn

#10) 5.25 years

28.3 and 28.4 Homework Answer Key

28.3 pg. 856 #11 to 15

#11) Fission involves splitting nuclei into smaller fragments. This is a reliable, controllable source of energy but poses operational dangers and produces radioactive wastes. Fusion occurs when two nuclei combine to produce a nucleus of greater mass. Fusion reactions require very high temperatures and are difficult to contain.

#12) Fission is controlled through neutron moderation and neutron absorption. The heat produced is removed from the reactor core and is used to generate steam to drive a turbine. The spinning turbine generates electricity.

#13) A nuclear chain reaction involves the splitting of atomic nuclei that releases energetic neutrons that split more nuclei.

#14) by using neutron moderation and neutron absorption.

#15) abundant, cheap fuels and no radioactive waste products.

28.4 pg. 861 #16 to 21

#16) Geiger counters detect beta radiation that ionizes gas in the counter's tube, producing a current and an audible or visible signal. Scintillation counters use phosphors to convert a portion of the energy from ionizing radiation into easily detectable signals. Scintillation counters detect all types of ionizing radiation. Film badges are enclosed layers of photographic film that are developed to reveal the strength and type of radiation exposure.

#17) Radioisotopes can be used to analyze samples for age and content; study chemical reactions, molecular structure, and agricultural assimilation of herbicides, pesticides, and fertilizers; and diagnose and treat diseases.

#18) A Geiger counter only detects beta radiation; a scintillation counter can detect all types of ionizing radiation.

#19) No. because a Geiger counter counter cannot detect alpha radiation.

#20) Tracers are used to study reaction mechanisms and the uptake of substances by organisms.

#21) The seed can be placed at the location of the tumor, which minimizes the effect on normal cells, and there are no radioactive waste products to be disposed of.

Chapter 28 REVIEW

Answers

- **22.** Each isotope of an element has the same atomic number but a different atomic mass. A radioisotope is an isotope that is radioactive.
- **23.** $^{226}_{88}$ Ra $\rightarrow ^{222}_{86}$ Rn + $^{4}_{2}$ He
- **24.** $^{210}_{82}$ Pb $\rightarrow ^{210}_{83}$ Bi $+ ^{0}_{-1}$ e
- **25. a.** α, 2+ **b.** β, 1- **c.** γ, 0
- **26. a.** ${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}He;$ thorium-234
 - **b.** $^{230}_{90}$ Th $\rightarrow ^{226}_{88}$ Ra $+ ^{4}_{2}$ He; radium-226
 - **c.** ${}^{235}_{92}U \rightarrow {}^{231}_{90}Th + {}^{4}_{2}He;$ thorium-231
 - **d.** $^{222}_{86}$ Rn $\rightarrow ^{218}_{84}$ Po + $^{4}_{2}$ He; polonium-218
- **27. a.** ${}^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$
 - **b.** ${}^{90}_{38}\text{Sr} \rightarrow {}^{90}_{39}\text{Y} + {}^{0}_{-1}\text{e}$
 - **c.** ${}^{40}_{19}\text{K} \rightarrow {}^{40}_{20}\text{Ca} + {}^{0}_{-1}\text{e}$
 - **d.** $^{13}_{7}N \rightarrow ^{13}_{8}O + ^{0}_{-1}e$
- 28. a. mass number: unchanged; atomic number: increases by 1
 b. mass number: decreases by 4; atomic number: decreases by 2
 c. mass number and atomic number: unchanged
- **29.** a. ${}^{234}_{92}$ U c. ${}^{206}_{82}$ Pb b. ${}^{206}_{81}$ Tl d. ${}^{226}_{88}$ Ra
- **30.** It undergoes radioactive decay.
- **31.** ${}^{17}_{9}F \rightarrow {}^{17}_{8}O + {}^{0}_{+1}e$
- **32.** a. ¹³₆C c. ¹⁶₈O b. ¹₁H d. ¹⁴₇N
- 33. a. platinum
 b. thorium
 c. francium
 g. vanadium
 d. technetium
 h. palladium
 Francium (c), technetium (d), and californium (f) have no stable isotopes.
- **34.** so exposure to the radioactivity is limited in time
- **35.** One half-life is the time required for one-half of the atoms of a radioisotope to emit radiation and decay.
- **36.** $6.3 \times 10^{-1} \text{ mg}$
- Natural radioactivity comes from elements in nature. Artificial radioactivity comes from elements created in nuclear reactors and accelerators.

Chapter 28 REVIEW

CONCEPT PRACTICE

- 22. Explain the difference between an isotope and a radioisotope. 28.1
- **23.** The disintegration of the radioisotope radium-226 produces an isotope of the element radon and alpha radiation. The atomic number of radium (Ra) is 88; the atomic number of radon (Rn) is 86. Write a balanced equation for this transformation. *28.1*
- **24.** A radioisotope of the element lead (Pb) decays to an isotope of the element bismuth (Bi) by emission of a beta particle. Complete the equation for the decay process by supplying the missing atomic number and mass number. *28.1*

 210 Pb $\longrightarrow {}_{83}$ Bi $+ {}^{0}_{-1}$ e

- 25. Write the symbol and charge for each. 28.1a. alpha particleb. beta particle
 - c. gamma ray
- **26.** Alpha radiation is emitted during the disintegration of the following isotopes. Write balanced nuclear equations for their decay processes. Name the element produced in each case. *28.1*
 - a. uranium-238 (²³⁸₉₂U)
 - **b.** thorium-230 (²³⁰₉₀Th)
 - c. uranium-235 $\binom{^{235}{92}}{92}$ U)
 - d. radon-222 (²²²₈₅Rv)
- 27. The following real-bisotopes are beta entirers. Write balanced nuclear equations for their decay processes. 251
 a. carbon-14 (¹/₆C²)
 - **b.** strontium-90 (38
 - c. potassium-40 $\binom{M}{13}$
 - d. nitrogen-13 (¹³₇N)
- 28. How are the mass number and atomic number of a nucleus affected by the loss of the following? *28.1*a. beta particle
 - b. alpha particle
 - c. gamma ray
- 29. The following radioactive nuclei decay by emitting alpha particles. Write the product of the decay process for each. 28.1
 a. ²³⁸₂₄Pu
 c. ²¹⁰₂₈Po

a. 94 Pu	C. 84 PO
b . ²¹⁰ ₈₃ Bi	d . ²³⁰ ₉₀ Th

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- **38.** The elements with atomic number greater than 92; none occurs in nature and all are ra-dioactive.
- **39.** The nuclei of certain isotopes are bombarded with neutrons. The nuclei break into two fragments and release more neutrons. The released neutrons hit other nuclei to start a chain reaction that releases large amounts of energy.
- 40. Answers will vary. Uranium-235 is one natural

- **30**. What happens to an atom with a nucleus that falls outside the band of stability? *28.2*
- **31**. Write an equation for the radioactive decay of fluorine-17 by positron emission. *28.2*
- **32**. Identify the more stable isotope in each pair. *28.2*
 - **a.** ${}^{14}_{6}C$, ${}^{13}_{6}C$
 - **b**. ³₁H, ¹₁H
 - c. ¹⁶/₈O, ¹⁸/₈O
 - d. 14 N, 15 N
- **33**. Name the elements represented by the following symbols and indicate which of them would have no stable isotopes. *28.2*
 - a. Pt e. Xe
 - b. Th f. Cf
 - c. Fr g. V
 - d. Tc h. Pd
- **34**. Why is it important that radioactive isotopes used internally for diagnosis or treatment have relatively short half-lives? *28.2*
- 35. Explain half-life. 28.2
- **36.** A patient is administered 20 mg of iodine-131. How much of this isotope will remain in the body after 40 days if the half-life for iodine-131 is 8 days? *28.2*
- **37.** What is the difference between natural and artificial radioactivity? *28.2*
- **38.** What are the transuranium elements? Why are they unusual? *28.2*
- **39.** Describe the process of nuclear fission and define a nuclear chain reaction. *28.3*
- **40.** Name one fissionable material that occurs in nature and one fissionable material that is artificial. *28.3*
- **41.** Fusion reactions produce enormous amounts of energy. Why is fusion not used to generate electrical power? *28.3*
- 42. Why are x-rays and the radiation emitted by radioisotopes called ionizing radiation? 28.4
- 43. What is the purpose of wearing a film badge when working with ionizing radiation sources? 28.4
- Explain how iodine-131 is used in both the diagnosis and treatment of thyroid disease. 28.4

fissionable material. Plutonium-239 is one artificially produced fissionable material.

- **41.** Fusion requires extremely high temperatures, which would destroy any container.
- **42**. Ionizing radiation, such as x-rays and gamma radiation, has sufficient energy to remove electrons from the atoms it hits.
- **43**. The film badge measures radiation exposure; an exposed film badge indicates how much radiation a worker has received.
- 44. In diagnosis, the amount of iodine uptake in

(ONCEPT MASTERY

- 45. Write nuclear equations for these conversions.
 - a. 30 P to 30 Si
 - b. 13C to 14C
 - c. ¹³¹/₅₄ I to ¹³¹/₅₄ Xe
- 46. What is the difference between the nuclear reactions taking place in the sun and the nuclear reactions taking place in a nuclear reactor?
- 47. Complete these nuclear equations.

a. $^{32}_{15}P \longrightarrow _ + _{-1}^{0}e$ **b**. $\longrightarrow \frac{14}{7}N + \frac{0}{-1}e$ c. $^{238}_{92}U \longrightarrow ^{234}_{90}Th + _$ d. $^{141}_{56}$ Ba \longrightarrow ____ + $^{0}_{-1}$ e e. $\longrightarrow \frac{181}{77}$ Ir + $\frac{4}{2}$ He

- 48. Write nuclear equations for the beta decay of the following isotopes.
 - a. 38 Sr
 - b. 14 C
 - c. 37 Cs
 - d. 239 Np

 - e. 50 Ti
- 49. The following graph shows the radioactive decay curve for thorium-234. Use the graph to answer the questions.



- a. What percent of the isotope remains after 60 days?
- b. How many grams of a 250-g sample of thorium-234 would remain after 40 days had passed?
- c. How many days would pass while 44 g of thorium-234 decayed to 4.4 g of thorium-234?
- d. What is the half-life of thorium-234?
- the thyroid is measured; in treatment, the radioactive iodine-131 is concentrated in and by the thyroid.
- **45.** a. ${}^{30}_{15}P + {}^{0}_{-1}e \rightarrow {}^{30}_{14}Si$
- **b.** ${}^{13}_{6}C + {}^{1}_{0}n \rightarrow {}^{14}_{6}C$
 - **c.** $^{131}_{53}I \rightarrow ^{131}_{54}Xe + ^{0}_{-1}e$
- 46. Nuclear fusion takes place in the sun. A nu-
- clear reactor utilizes nuclear fission. 47. a. 32/16S e. 185 Au
- C.⁴₂He

- 50. Write a nuclear equation for each word equation.
 - a. Radon-222 emits an alpha particle to form polonium-218.
 - b. Radium-230 is produced when thorium-234 emits an alpha particle.
 - c. When polonium-210 emits an alpha particle, the product is lead-206.
- 51. Describe the various contributions the following people made to the fields of nuclear and radiation chemistry.
 - a. Marie Curie
 - b. Antoine Henri Becquerel
 - c. James Chadwick
 - d. Ernest Rutherford

CRITICAL THINKING

- 52. Choose the term that best completes the second relationship.

a. decomposition : combin	nation
fission :	
(1) energy	(3) radioactivity
(2) nuclei	(4) fusion
b. anion : cation electron :	
(1) positron	(3) beta particle
(2) neutron	(4) alpha particle
c. umbrella : rain wood :	
(1) radioactive decay	(3) beta radiation
(2) gamma radiation	(4) x-radiation

- 53. Compare the half-life of an element to a singleelimination sports tournament.
- 54. Why does the relatively large mass and charge of an alpha particle limit its penetrating power?
- 55. Why might radioisotopes of C, N, and O be especially harmful to living creatures?

CUMULATIVE REVIEW

- 56. What is the Pauli exclusion principle? What is Hund's rule?
- 57. Balance the following equations.

a.
$$Ca(OH)_2 + HCI \longrightarrow CaCl_2 + H_2O$$

b. $Fe_2O_3 + H_2 \longrightarrow Fe + H_2O$ NoHCO + H SO

c. NaHCO₃ + H₂SO₄
$$\longrightarrow$$
 No

$$\label{eq:solution} \begin{split} & Na_2SO_4+CO_2+H_2O\\ \textbf{d}.\ C_2H_6+O_2 & \longrightarrow CO_2+H_2O \end{split}$$

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b. ¹⁴₆C d. 141 La **48.** a. ${}^{90}_{38}\text{Sr} \rightarrow {}^{90}_{39}\text{Y} + {}^{0}_{-1}\text{e}$ **b.** ${}^{14}_{6}C \rightarrow {}^{14}_{7}N + {}^{0}_{-1}e$ **c.** $^{137}_{55}$ Cs $\rightarrow ^{137}_{56}$ Ba + $^{0}_{-1}$ e **d**. $^{239}_{93}Np \rightarrow ^{239}_{94}Pu + ^{0}_{-1}e$ $e_{22}^{50}Ti \rightarrow {}^{50}_{23}Sc + {}^{0}_{-1}e$

c. about 83 days 49. a. about 20% d. 25 days b. about 85 g

Chapter 28 REVIEW

- **50.** a. ${}^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He}$
 - **b.** $^{234}_{90}$ Th $\rightarrow ^{230}_{88}$ Ra + $^{4}_{2}$ He
 - **c.** ${}^{210}_{84}\text{Po} \rightarrow {}^{206}_{82}\text{Pb} + {}^{4}_{2}\text{He}$
- 51. a. Named radioactivity and discovered several radioactive elements

b. Discovered natural radioac-

- tivity from uranium ores.
- c. Discovered the neutron.
- d. Transmuted elements.
- 52. a. 4 b. 1 C. 3 53. In every round of the tournament, one-half the teams are eliminated; in every half-life, one-half of the substance decays. In the tournament a single team eventually emerges as the winner and the tournament stops, but radioactive decay continues (almost) indefinitely.
- 54. An alpha particle is much more likely than other kinds of radiation to collide with another particle and be stopped. At the atomic level, the larger the size of a particle, the greater is the chance of its striking another particle. The greater the magnitude of a particle's charge, the more strongly it will be attracted to particles of opposite charge.
- 55. Radioactive isotopes of these elements can be incorporated into the body tissues of organisms. When the isotopes decay, they can damage tissues very easily.
- 56. The Pauli exclusion principle states that no two electrons in an atom can have the same quantum numbers. Hund's rule states that electrons occupying orbitals of equal energy are distributed with unpaired spins as much as possible.
- 57. a. $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 +$ $2H_2O$ **b.** $Fe_2O_3 + 3H_2 \rightarrow 2Fe + 3H_2O$ **c.** $2NaHCO_3 + H_2SO_4 \rightarrow Na_2SO_4 +$ $2CO_2 + 2H_2O$ **d.** $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$

13.1, 13.2 and 13.3 Homework Answer Key

13.1 pg. 366 #1 to 4

#1) Dalton: elements are composed of atoms; Thomson: discovered electrons; Rutherford: discovered nucleus; Bohr: quantized energies of electrons; Schrödinger:quantum mechanics model

#2) The quantum mechanical states that electrons can have only fixed energy levels. Electrons are located in orbitals that may be visualized as clouds of various shapes at different distances from the nucleus.

#3) In an atom, electrons can only exist in certain fixed energy levels. To move from one energy level to another requires the emission of absorption of an exact amount of energy, or quantum.

#4) a.	3	b.	1	c. 7	d.	3	3 e. :	5
13.2 #5) a.	pg. C =	369 # 1s²2s²2	#5, 2p²	6; pg. 370 #7 to	10 b.	Þ	Ar = 1s²2s²2p ⁶ 3	3s²3p ⁶
#6) a.	в =	1s ² 2s ² 2	2p1	(1 unpaired electron)	b.	S	Si = 1s²2s²2p ⁶ 3	3s ² 3p ² (2 unpaired electrons)
#7) a.	Li =	1s ² 2s ¹		b. $F = 1s^2 2s^2 2$	2p⁵		C.	Rb = [Kr] 5s ¹

#8) Half-filled energy sublevels (orbitals) are more stable than partially filled sublevels (orbitals).

#9) 3d, 4s, 3p, 3s, 2p

#10) The 3s and 3p orbitals are already filled; therefore the last electron must go to the next energy sublevel (orbital), which is 4s.

13.3 pg. 383 #16 to 19

#16) Electrons in atoms absorb energy, then lose the energy and emit it as light.

#17) Metals eject electrons when certain wavelengths of light shine on them. Classical physics assumed any wavelength of light could cause the photoelectric effect. However, only light with some minimum frequency and threshold energy can cause an electron to be ejected.

#18) The ground state is the lowest energy level of an electron. The excited state is an energy state higher than the ground state.

#19) c (shortwave radio) > a. (infrared) > b. (x-rays)

Chapter 13 REVIEW

Answers

- **20.** electrons and positively charged particles
- **21.** Electrons have fixed energies. To move to another level, they must emit or absorb a quantas of energy.
- **22.** In Rutherford's model, negatively charged electrons surround a dense, positively charged nucleus. In Bohr's model, electrons are assigned to circular orbits of fixed energy.
- **23.** 90% of the time an electron is inside this boundary.
- **24.** a region beyond the nucleus where there is a high probability of finding an electron
- **25.** The 1*s* orbital is spherical. The 2*s* orbital is spherical with a diameter larger than that of the 1*s* orbital. The 2*p* orbital is dumbbell shaped and reaches beyond the diameter of the 2*s* orbital.

26. a. 2 b. 1 c. 3 d. 6

- 27. Aufbau principle: electrons occupy the lowest possible energy levels; Pauli exclusion principle: an orbital can hold at most two electrons; Hund's rule: before pairing of electrons occurs, one electron occupies each of a set of orbitals with equal energies.
- **28.** a. $1s^22s^22p^63s^23p^3$ b. $1s^22s^22p^63s^2$ c. $1s^22s^22p^5$ d. $1s^22s^22p^5$
- **29.** The *p* orbitals in the third quantum level have three electrons.
- 30. b and c are invalid.

31.	a. 2	c. 2	e .6	g . 14
	b. 6	d . 10	f. 2	h. 6

- 32. a. 8 b. 8 c. 8
- 33. a. 1s²2s²2p⁶3s²3p⁶3d¹⁰4s²4p⁴
 b. 1s²2s²2p⁶3s²3p⁶3d⁸4s²
 c. 1s²2s²2p⁶3s²3p⁶3d⁸4s²
 d. 1s²2s²2p⁶3s²3p⁶3s⁶3d⁶4s²
- **34.** violet, indigo, blue, green, yellow, orange, red
- **35.** Frequency is the number of wave cycles that pass a given point per unit time. Frequency units are cycles, reciprocal seconds (s⁻¹), or hertz. Frequency and wavelength are inversely related.

Chapter 13 **REVIEW**

CONCEPT PRACTICE

- 20. Which subatomic particles did Thomson include in the plum-pudding model of the atom? 13.1
- **21**. How did Bohr answer the objection that an electron traveling in a circular orbit would radiate energy and fall into the nucleus? *13.1*
- 22. Describe Rutherford's model of the atom and compare it with the model proposed by his student Niels Bohr. 13.1
- 23. What is the significance of the boundary of an electron cloud? 13.1
- 24. What is an atomic orbital? 13.1
- 25. Sketch 1s, 2s, and 2p orbitals using the same scale for each. 13.1
- **26**. How many electrons are in the highest occupied energy level of these atoms? *13.2*
 - a. barium
 - b. sodiumc. aluminum
 - d. oxygen
- 27. What are the three rules that govern the filling of atomic orbitals by electrons? 13.2
- **28.** Write electron configurations for the elements that are identified only by these atomic numbers. *13.2*
 - **a**. 15
 - **b**. 12
 - **c**. 9 **d**. 18
- 29. What is meant by 3p³? 13.2
- **30.** Which of these orbital designations are invalid?
 - a. 4s
 - **b**. 3f
 - **c**. 2d
 - **d**. 3d
- 31. What is the maximum number of electrons the can go into each of the following sublevels? 13
 - a. 2s
 - **b**. 3p
 - c. 4s
 - e. 4n
 - f. 5s
 - g. 4f
 - h. 5p

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- **32**. How many electrons are in the second energy level of an atom of each element? 13.2
 - a. chlorine
 - b. phosphorus
 - c. potassium
- Write electron configurations for atoms of these elements. 13.2
 - a. selenium
 - b. vanadium
 - c. nickel d. calcium
- List the colors of the visible spectrum in order of increasing wavelength. 13.3
- **35.** What is meant by the frequency of a wave? What are the units of frequency? Describe the relationship between frequency and wavelength. 13.3
- 36. Use a diagram to illustrate each term. 13.3
- a. wavelength
 b. amplitude
- c. wave cycle
- **37.** Explain the difference between the laws of classical physics and the quantum concept when describing the energy lost or gained by an object. *13.3*
- **38.** What is the energy of a photon of green light with a frequency of $5.80 \times 10^{14} \text{ s}^{-1}$? 13.3
- **39.** How did Planck influence the development of modern atomic theory? *13.3*
- 40. What will happen if the following occur? 13.3 a. Monochromatic light shining on the allest
 - Monochromatic light shining on the alkali metal cesium is just above the threshold frequency.
 - b. The intensity of light increases, but the frequency remains the same.
 - c. Monochromatic light of a shorter wavelength is used.
- 41. Explain the difference between a photon and a quantum. 13.3
- 42. What happens when a hydrogen atom absorbs a quantum of energy? 13.3
- **43.** When white light is viewed through sodium vapor in a spectroscope, the spectrum is continuous except for a dark line at 589 nm. How can you explain this observation? *13.3*
- 44. What is the wavelength of a 2500-kg truck traveling at a rate of 75 km/h? 13.3
- **37.** Classical physics viewed energy changes as continuous, occurring in any quantity. In the quantum concept, energy changes occur in discrete units called quanta.
- **38.** 3.84×10^{-19} J
- **39.** He showed that the quanta of radiation ^{emil-} ted depends on the frequency of radiation.
- 40. a. Emitted electrons have a low velocity.
 b. More electrons are emitted; velocity stays low.
 - c. Emitted electrons have higher velocity.

CONCEPT MASTERY

- Give the symbol for the atom that corresponds
 - to each electron configuration.
 - a. 1s22s22p63s23p6
 - b. 1s²2s²2p⁶3s²3p⁶3d¹⁰4s²4p⁶4d⁷5s¹
 - c. 1s²2s²2p⁶3s²3p⁶3d¹⁰4s²4p⁶4d¹⁰4f⁷
 - 5s25p65d16s2
- M. Write the electron configuration for an arsenic atom. Calculate the total number of electrons in each energy level and state which energy levels are not full.
- 4. How many paired electrons are there in an atom of each element?
- a, helium
- b. boron
- c. sodium
- d. oxygen
- 48. An atom of an element has two electrons in the first energy level and five electrons in the second energy level. Write the electron configuration for this atom and name the element. How many unpaired electrons does an atom of this element have?
- 4. Suppose your favorite AM radio station broadcasts at a frequency of 1150 kHz. What is the wavelength, in centimeters, of the radiation from the station?
- 9. A mercury lamp, such as the one below, data radiation with a wavelength of 4.36×10^{-5}



- a. What is the wavelength of this radiation in centimeters?
- b. In what region of the electromagnetic spectrum is this radiation?
- ^{c.} Calculate the frequency of this radiation.
- 51. Calculate the energy of a photon of red light with a wavelength of 6.45×10^{-5} cm. Compare your answer with the answer to Problem 38. Is red light of higher or lower energy than green light?

52. Give the symbol and name of the elements that correspond to these configurations.

a. 1s²2s²2p⁶3s¹ **b**. $1s^2 2s^2 2p^3$

- c. 1s²2s²2p⁶3s²3p²
- d. $1s^2 2s^2 2p^4$
- e. 1s²2s²2p⁶3s²3p⁶4s¹
- f. 1s²2s²2p⁶3s²3p⁶3d²4s²

CRITICAL THINKING

53. Choose the term that best completes the second relationship.

a. orbital : energy level apartment :	
(1) floor	(3) building
(2) room	(4) stairway
b. water : container electrons :	
(1) frequency	(3) nuclei
(2) sublevel	(4) orbitals
• • • • • • • • • • • • • • • • • • •	Alam endaile la link

- c. electromagnetic radiation : visible light student body : (1) eleventh graders (3) teachers
 - (2) principal (4) textbooks
- 54. Traditional cooking methods make use of infrared radiation (heat). Microwave radiation cooks food faster. Could radio waves be used for cooking? Explain.
- 55. Think about the currently accepted models of the atom and of light. In what ways do these models seem strange to you? Why are these models not exact or definite?

CUMULATIVE REVIEW

- 56. A potassium atom has a diameter of about 0.406 nm. Express this in meters and micrometers. If the measurement is always the same, explain why the two numbers are different.
- 57. Balance the following chemical equations.
 - a. $KNO_3 + H_2SO_4 \longrightarrow K_2SO_4 + HNO_3$
 - **b.** $Cu_2O + H_2 \longrightarrow Cu + H_2O$ **c.** $NO + Br_2 \longrightarrow NOBr$

 - **d**. $SnO_2 + CO \longrightarrow Sn + CO_2$
- 58. Calculate the volume of O2 at STP required for the complete combustion of 5.00 L of acetylene (C2H2) at STP.

 $2C_2H_2(g) + 5O_2 \longrightarrow 4CO_2 + 2H_2O(l)$

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- [¶], Aquantum is a discrete amount of energy. Photons are light quantas.
- k, Its electron is raised to a higher energy level. R. The outermost electron of sodium absorbs
- photons of wavelength 589 nm.
- 4. 1.3 × 10⁻³⁸ m
- 45. a. Ar b. Ru c. Gd
- 4. $1s^22s^22p^63s^23p^64s^23d^{10}4p^3$; Total = 33; the first three levels are full; the fourth level is partially filled. 47. a.2 b.4 c. 10 d. 6

- **b**. visible
- **C.** $6.88 \times 10^{14} s^{-1}$
- **51.** 3.08×10^{-19} J; lower energy
- 52. a. Na, sodium
 - b. N, nitrogen c. Si. silicon
- d. O, oxygen e. K, potassium f. Ti, titanium

Chapter 13

53. a. 1 b. 4 c. 1

- 54. Answers will vary. Students may note that radio waves have the lowest energy in the electromagnetic spectrum, and thus would not be energetic enough to cook food. Others may reason that if microwaves cook food faster than infrared radiation, then radio waves would cook food even faster.
- 55. The model of the atom uses the abstract idea of probability; light is considered a particle and a wave at the same time. Atoms and light cannot be compared with familiar objects because humans cannot experience them directly. Because matter and energy behave differently at the atomic level than at the level humans can observe directly.
- **56.** 4.06×10^{-10} m; 4.06×10^{-4} µm; The units are different, but the values are all equal.
- 57. a. 2, 1, 1, 2 **b**. 1, 1, 2, 1 **c.** 2, 1, 2 d. 1, 2, 1, 2
- 58. 12.5 L

49. 2.61×10^4 cm **50. a.** 4.36×10^{-5} cm

48. $1s^2 2s^2 2p^3$; nitrogen; 3 unpaired electrons

14.1 and 14.2 Homework Answer Key

14.1 pg. 396 #1 to 5

#1) a. $C = 1s^22s^22p^2$ b. $V = 1s^22s^22p^63s^23p^64s^23d^3c$. $Sr = [Kr] 5s^2$

#2) a. He, Be, Mg, Ca, Sr, Ba, Ra b. F, Cl, Br, I, At c. Ti, Zr, Hf, Rf

#3) Sodium and potassium have similar chemical and physical properties because they have similar electron configurations with a single electron in the outermost *s*-orbital.

#4) a.	Transition metal (Ag)	b.	Noble gas (Kr)
C.	Transition Metal (Cr)	d.	Representative element (Si)

#5) Cu, Cd, Au, Co; These elements have partially filled *d*-orbitals.

14.2 pg. 406 #6 to 9

#6) a. - first ionization energy and c. - electronegativity

#7) sodium, aluminum, sulfur, chlorine; periodic trend

#8) The anion is larger than its parent atom.

#9) a. Sodium b. Phosphorus

Chapter 14 REVIEW

CONCEPT PRACTICE

- 10. What are the noble gases, the representative elements, the transition elements, and the inner transition elements? 14.1
- 11. Use Figure 14.2 to write the electron configuration of these elements. 14.1 b. magnesium a. boron c. arsenic
- 12. Which of the following are representative elements: Na, Mg, Fe, Ni, Cl? 14.1
- 13. Write the electron configuration of these elements. 14.1
 - a. the inert gas in period 3
 - b. the element in Group 4A, period 4
 - c. the element in Group 2A, period 6
- 14. Explain how an element's outer electron configuration is related to its position in the periodic table. 14.1
- 15. Use Figure 14.2 to write the electron configuration of these atoms. 14.1
 - a. fluorine b. zinc c. aluminum d. tin
- 16. What are the symbols for all the elements with the following outer configurations? 14.1 **b**. $s^2 p^4$ **c**. $s^2 d^{10}$ a. s1
- 17. Explain why fluorine has a smaller atomic radius than both oxygen and chlorine. 14.2
- 18. Indicate which element in each pair has the greater atomic radius. 14.2 a. sodium, lithium
 - b. strontium, magnesium
 - c. carbon, germanium
 - d. selenium, oxygen
- 19. Distinguish between the first and second ionization energy of an atom. 14.2
- 20. Indicate which element in each pair has the greater first ionization energy. 14.2 a. lithium, boron
 - b. magnesium, strontium
 - c. cesium, aluminum
- 21. Would you expect metals or nonmetals to have higher ionization energies? Why? 14.2
- 22. Arrange the following elements in order of increasing ionization energy. 14.2 a. Be, Mg, Sr b. Bi. Cs. Ba c. Na, Al, S
- 23. Why is there a large increase between the first and second ionization energies of the alkali metals? 14.2

Answers

- 10. noble gases: Group 0; representative elements: Groups 1A-7A; transition elements: Groups 1B-8B; inner transition elements: a separate section between Groups 2A and 3B
- 11. a. 1s²2s²2p¹
 - **b.** $1s^22s^22p^63s^2$
- c. $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$
- 12. Na, Mg, Cl 13. a. Ar: $1s^22s^22p^63s^23p^6$

- 24. Which particle has the larger radius in each atom/ion pair? 14.2
- b. S. S²⁻ a. Na. Na⁺ c. I, I⁻ d. Al, Al³⁺
- 25. How does the ionic radius of a typical metallic atom compare with its atomic radius? 14.2
- 26. Explain why the noble gases do not appear in Table 14.2. 14.2
- 27. Which element is more electronegative? 14.2 a. Cl. F b. C, N c. Mg, Ne d. As. Ca

CONCEPT MASTERY

- 28. The Mg²⁺ and Na⁺ ions each have ten electrons surrounding the nucleus. Which ion would you expect to have the smaller radius? Why?
- 29. Explain why it takes more energy to remove a 4s electron from zinc than from calcium.
- 30. The graphs show the relationship between the electronegativities and first ionization energies for period 2 and period 3 elements.
 - a. State the general trend between these two values in each period.
 - b. Propose an explanation for this trend.





Chemical Periodicity 409

- **b.** Ge: $1s^22s^22p^63s^23p^63d^{10}4s^24p^2$
- **c.** Ba: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^6$ $6s^2$
- 14. An element's outer electron configuration places it in a particular column (group) of the periodic table.
- 15. a. 1s²2s²2p⁵
 - **b**. 1s²2s²2p⁶3s²3p⁶3d¹⁰4s²
 - **c.** $1s^2 2s^2 2p^6 3s^2 3p^1$
 - d. $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}5s^25p^2$
- 16. a. H, Li, Na, K, Rb, Cs, Fr

Chapter 14 REVIEW

b. O, S, Se, Te, Po c. Zn, Cd, Hg, Uub

- **17.** Fluorine has a smaller atomic radius than oxygen because fluorine has one more nuclear charge. Fluorine has a smaller radius than chlorine because fluorine has eight fewer electrons.
- 18. a. sodium **c**. germanium **b**. strontium d. selenium
- 19. The first ionization energy is the energy needed to remove the outermost electron. The second ionization energy is the energy needed to remove the secondoutermost electron.
- 20. a. boron C. aluminum **b**. magnesium
- 21. nonmetals; The nuclear charge increases, but the shielding effect is the same, thus creating greater electron attraction.
- 22. a. Sr, Mg, Be c. Na, Al, S b. Cs, Ba, Bi
- 23. An atom of an alkali metal becomes stable by losing one electron. Removing the second electron involves removing an electron from an ion that has a stable noble-gas configuration. This requires much more energy.
- **24. a**. Na **b**. S²⁻ **c**. I⁻ d. Al
- 25. The radius of a cation is smaller than the atom from which it forms.
- 26. Noble gases generally do not form compounds.
- 27. a.F b.N c.Mg d.As
- **28.** Mg^{2+} has more protons in its nucleus; its electron attraction is therefore greater.
- 29. Zinc has more protons than calcium, thus attracting the 4s electrons more.
- 30. a. The general trend is that the first ionization energy increases as electronegativity increases. This is true for both period 2 and period 3.
 - b. A positive correlation is expected because both properties measure the interaction between the nucleus and the surrounding electrons.



Chapter 14

REVIEW

Answers

- **31. a.** potassium, K, [Ar]4s¹ **b.** aluminum, Al, [Ne]3s²3p¹
 - **c.** sulfur, S, [Ne] $3s^23p^4$
 - **d**. barium, Ba, [Kr] $6s^2$
- **32. a.** Ca²⁺ **b.** P³⁻ **c.** Cu⁺
- **33.** scandium, Sc, $[Ar]3d^{1}4s^{2}$ titanium, Ti, $[Ar]3d^{2}4s^{2}$ vanadium, V, $[Ar]3d^{3}4s^{2}$ chromium, Cr, $[Ar]3d^{5}4s^{1}$ manganese, Mn, $[Ar]3d^{5}4s^{2}$ iron, Fe, $[Ar]3d^{6}4s^{2}$ cobalt, Co, $[Ar]3d^{7}4s^{2}$ nickel, Ni, $[Ar]3d^{8}4s^{2}$ copper, Cu, $[Ar]3d^{10}4s^{1}$ zinc, Zn, $[Ar]3d^{10}4s^{2}$
- **34. a.** 2 **b.** 1 **c.** 3
- **35.** Magnesium achieves a stable electron configuration by losing two electrons; aluminum achieves a stable electron configuration by losing three electrons.
- 36. a. 1851–1900: 25 elements
 - **b.** Mendeleev's periodic table helped scientists predict the existence of undiscovered elements.
 - **c.** None of these elements are found in nature.
- 37. a. Possible cations are Rb⁺ and Sr²⁺; possible anions are Br⁻, Se²⁻, and As³⁻.
 - **b**. No; a cation is isoelectronic with the noble gas in the preceding period; an anion is isoelectronic with the noble gas in the same period.
- **38.** $P_{\rm CO_2} = 1.23$ atm; $P_{\rm N_2} = 0.879$ atm
- **39. a.** $2Ag + S \rightarrow Ag_2S$ **b.** $Na_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2NaOH$ **c.** $Zn + 2HNO_3 \rightarrow Zn(NO)_2 + H_2$ **d.** $2H_2O + 2SO_2 + O_2 \rightarrow 2H_2SO_4$ **40.** 69.9 g Fe **41. a.** Li_2SO_4 **b.** $Zn_3(PO_4)_2$
 - c. KMnO₄
 - d. SrCO₃

- 31. Give the electron configuration of the element found at each location in the periodic table.
 a. Group 1A, period 4
 b. Group 3A, period 3
 c. Group 2A, period 6
- 32. In each pair, which ion is larger? a. Ca^{2+} , Mg^{2+} b. Cl^{-} , P^{3-} c. Cu^{+} , Cu^{2+}
- Give the names, symbols, and electron configurations for the ten period-4 transition metals.

CRITICAL THINKING

34. Choose the term that best completes the second relationship.

a. sister : brother	oxygen :
(1) hydrogen	(3) silicon
(2) sulfur	(4) Group 6A
b. potassium : cation	sulfur :
(1) anion	(3) yellow
(2) nonmetal	(4) solid
c. magnesium : s orbital	zinc :
(1) s orbital	(3) d orbital
(2) p orbital	(4) forbital

- **35.** There is a large jump between the second and third ionization energies of magnesium. The corresponding large jump is between the third and fourth ionization energies of aluminum. Explain.
- The following graph shows how many elements were discovered before 1750 and in tact 50-year period since then.
 - a. In which 50-year period were the result elements discovered?
 - b. How did Mendeleev's work contribute to the discovery of so many elements?
 - c. What characteristic do all the elements discovered since 1950 have in common?



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- 37. Atoms and ions with the same number of electrons are called isoelectronic.
 - Name a cation and an anion that are isoelectronic with krypton.
 - b. Is it possible for a cation to be isoelectronic with an anion from the same period? Explain

CUMULATIVE REVIEW

- 38. A 2.00-L flask at 27 °C contains 4.40 g of carbon dioxide and 2.00 g of nitrogen gas. What is the pressure (in atm) of each of the two components?
- **39.** Balance the following chemical equations. **a.** $Ag + S \longrightarrow Ag_2S$ **b.** $Na_2SO_4 + Ba(OH)_2 \longrightarrow BaSO_4 + NaOH$ **c.** $Zn + HNO_3 \longrightarrow Zn(NO_3)_2 + H_2$ **d.** $H_2O + SO_2 + O_2 \longrightarrow H_2SO_4$
- 40. The smelting of iron ore consists of heating the ore with carbon.

 $2Fe_2O_3 + 3C \longrightarrow 4Fe + 3CO_2$

What mass of iron can be obtained from \log_g of the ore?

- 41. Write chemical formulas for the following compounds.
 - a. lithium sulfate
 - **b**. zinc phosphate
 - c. potassium permanganate
 - d. strontium carbonate
- 42. If a gas sample at 25 °C occupies a volume of 2.93 L, what will be the volume at 500 °C if the pressure is unchanged?

CONCEPT CHALLENGE

- 43. The ions S²⁻, Cl⁻, K⁺, Ca²⁺, and Sc³⁺ have be same total number of electrons as the noble gas argon. How would you expect the radiiot these ions to vary? Would you expect to see the same variation in the series O²⁻, F⁻, Na⁺, Mg⁺, and Al³⁺, in which each ion has the same total number of electrons as the noble gas neon? Why or why not?
- 44. Using a chemistry reference book, make a table for the Group 2A elements. Include densities atomic masses, formulas of the chlorides and oxides, and first ionization energies. Can you justify placing these elements in one group on the basis of these data?

42. 7.60 L

- **43.** The ionic radii decrease from S^{2-} , Cl^- , Ar, K^+ , Ca^{2+} , to Sc^{3+} as the number of protons increases. The radii decrease from O^{2-} , F^- , Ne, Na^+ , Mg^{2+} , to Al^{3+} for the same reason.
- 44. The table shows gradually increasing atomic masses and decreasing ionization energies for these elements. All need two atoms of chlorine and one atom of oxygen for compounds. These trends justify placing these elements in one group.

6.1 to 6.4 Homework Answer Key

6.1 pg. 137 #1 to 9

#1) a. Sulfide, S ^{2−}	b. Aluminum ion, Al ³⁺	c. Calcium ion, Ca ²⁺			
#2) a. Ba ²⁺ lost 2 electrons	b. As ³⁻ gained 3 electrons	c. Cu ²⁺ lost 2 electrons			
#3) Ionic compounds are usually sol	ids with high melting points formed fr	om a metal and a nonmetal.			
#4) Metals tend to form positively ch	arged cations; nonmetals tend to for	n negatively charged anions.			
#5) It is an anion.					
#6) They are monatomic noble gase	S.				
#7) A molecule is the smallest neutral particle of a substance that retains all the substance's properties; diatomic O_2 has two atoms; triatomic O_3 has three atoms.					

b. Zinc ion, Zn²⁺

#9) a. Fluoride, F¹⁻ b. Sulfide, S²⁻

6.2 pg. 142 #10, 11, 14 and 15

#10) 2:1

#11) No. The ratio in the compound is 1.62: 1 rather than 1.75:1 or 7:4

#14) law of definite proportions

#8) a. Potassium ion, K⁺

#15) law of multiple proportions

6.3 pg. 145 #16, 17; pg. 146 #18, 19; pg. 148 #20, 22 and 23 #16) a. Se²⁻ b. Ba²⁺ c. Cs¹⁺ d. P³⁻ b. O²⁻ gained 2 electrons #17) a. Fe^{3+} lost 3 electrons c. Cu¹⁺ lost one electron d. Cd²⁺ lost 2 electrons #18) a. Selenide, anion b. Barium ion, cation c. Cesium ion, cation d. Phosphide, anion #19) a. Iron (III) ion b. Oxide c. copper (I) ion d. Cadmium ion

#20) by noting the group in which the element is found

#22) a.	Potassium ion, K⁺	b.	Sulfide, S ²⁻	c. no ion formed		
d.	Bromide, Br ¹⁻	e.	Beryllium, Be ²⁺	f. Sodium, Na ¹⁺		
#23) a.	NH_{4}^{1+}	b. Sn ²⁺	c. C	rO₄²-	d.	NO ₃ ¹⁻
e.	CN ¹⁻	f. Fe ³⁺	g. N	InO ₄ ¹⁻	h.	Mn ²⁺

6.4 pg. 151 #24, 3	25; pg. 153 #26,	27; pg. 155 #28, 2	.9; pg. 156 #30 to	o 36
#24) a. BaS	b. Li ₂ O	c. Ca ₃ N ₂	d. Cul ₂	

#25) a.	Nal	b. SnCl ₂	C.	K ₂ S	d.	Cal ₂
#26) a. c.	Zinc sulfide Barium oxide		b. d.	Potassium chloride Copper (II) bromid	e e	
#27) а. с.	Calcium oxide Iron (II) sulfide		b. d.	Copper (I) selenide Aluminum fluoride	e	
#28) a.	$(NH_4)_2SO_3$	b. $Ca_{3}(PO_{4})_{2}$	C.	AI(NO ₃) ₃	d.	K ₂ CrO ₄
#29) a.	LiHSO ₄	b. Cr(NO ₂) ₃	c.	HgBr ₂	d.	$(NH_4)_2Cr_2O_7$
#30) a. c.	Calcium oxalate Potassium permanga	anate	b. d.	Potassium hypoch Lithium sulfite	lori	te
#31) a. c.	Aluminum hydroxide tin (II) phosphate		b. d.	Sodium iodide sodium chromate		

#32) The formula must be written so the net ionic charge is zero. The cation is written first, the anion second. To write the name form the formula, name the cation followed by the anion.

#33) Write the formula (symbol and charge) for each ion. Then use the criss cross method. To do the reverse, name the cation followed by the anion.

#34) a.	$Cr(NO_2)_3$	 Manganese phosphate 	c. Lithium fluoride
d.	NaClO ₄	e. Lead (II) acetate	f. Mg(HCO ₃) ₂

#35) When more than one polyatomic ion is needed to balance a formula

#36) The net ionic charge must be zero.

6.5 and 6.6 Homework Answer Key

6.5 pg. 159 #37, 38; pg. 160 #39 to 42

#37) a.	Oxygen diflue	oride	b.	Dichlorine octaoxide	C.	sulfur trioxide
#38) a.	NF ₃	b. S_2Cl_2		c. N ₂ O ₄		

#39) Use the prefixes to determine the subscript for each element in the formula. Write the correct symbols for the two elements. When there are multiple atoms of an element, place a

prefix before the element name. The name of the second element will end in ~ide.

#40) a.	Sulfuric acid	b. Carbonic acid	c. HNO ₃	d. H ₃ PO ₄
#41) a.	Carbon disulfide	b. Dichlorine heptaoxide	c. CBr ₄	d. P_2O_3

#42) hydrogen

6.6 pg. 163 #43 and 44

#43) a.	Calcium carbonate	. Potassi	um permanganate
C.	lead (II) chromate	l. Calciun	n hydrogen phosphate
e.	Tin (II) dichromate	Magnes	ium phosphide

#44) a. Sn(OH) ₂	b. BaF ₂	c. I ₄ O ₉
d. $Fe_2(C_2O_4)_3$ or $Fe_2(OOCCOO)_3$	e. CaS	f. Al(HCO ₃) ₃

Gianard

Answers

- 45. Ions are formed when atoms lose or gain electrons.
- 46. a. 1 gained
 - b. 1 lost
 - c. 3 gained
 - d. 2 lost
 - e. 1 lost
 - f. 1 gained
- 47. a. bromide ion, anion
 - b. sodium ion, cation
 - c. arsenide ion, anion
 - d. calcium ion, cation
 - e. copper(1) ion, cation
 - f. hydride ion, anion
- 48. The net positive charge on the cations is exactly balanced by the net negative charge on the anions.
- 49. a. ionic d. molecular
 - b. molecular e. ionic
 - f. molecular C. ionic
- 50. a. Yes: the ratio of the mass of the colorless gas to the mass of the white powder is a constant. b. law of definite proportions c. 3.5 g colorless gas
- 51. a. carbon 6, hydrogen 8, oxygen 6
 - b. carbon 5, hydrogen 8, oxygen 4. sodium 1
 - c. carbon 12, hydrogen 22, oxygen 11

d. carbon 7, hydrogen 5, nitro-

gen 3, oxygen 6 e. nitrogen 2, hydrogen 4, oxy-

gen 3

- 52. ionic; high melting point
- 53. a. O²⁻ e. Cu2+ C. Li **b**. Pb²⁺ **d**. N³⁻ f. F
- 54. a. barium ion b. iodide ion c. silver ion d. mercury(II) ion e. phosphide ion f. tin(IV) ion
- 55. cyanide (CN⁻) and hydroxide (OH^{-})
- 56. a. hydroxide e. hydrogen b. lead(IV) phosphate C. sulfate f. dichromate
 - d. oxide q. aluminum h. chlorite
- 57. The net ionic charge is zero. Ionic compounds are electrically neutral.
- 58. a. Na₂O C. KCl b. SnS₂ d. Mg₃N₂

CONCEPT PRACTICE

- 45. Describe two ways that an ion forms from an atom. 6.1
- 46. State the number of electrons either lost or gained in forming each ion. 6.1
 - d. Ca2+ a. Br
 - b. Na* e. Cu*
 - c. As3f. H-
- 47. Name each ion in Problem 46. Identify each as an anion or a cation. 6.1
- 48. If ionic compounds are composed of charged particles (ions), why isn't every ionic compound either positively or negatively charged? 6.2
- 49. Would you expect the following pairs of atoms to combine chemically to give an ionic or a molecular compound? 6.2
 - a. Li and S d. F and Cl
 - b. O and S e. I and K
 - c. Al and O f. H and N
- 50. Four students heat different masses of a blue substance. The blue substance decomposes into a white powder and a colorless gas. The mass of the white powder and the mass of the colorless gas for each sample is plotted on the graph below. 6.2



- a. Does the graph support the idea that all four samples of blue substance are the same material? Explain.
- b. Assuming the blue powder is a compound, what law do the data presented here demonstrate?
- c. On heating, a fifth sample of the same blue substance gives 6.2 g of white powder. How many grams of colorless gas are also produced?
- 166 Chapter 6
- 59. By knowing the number of each ion in the formula and the ionic charge of the anion; net ionic charge is zero.
- 60. Parentheses are used to indicate more than one polyatomic ion.
- 61. a. $CaCO_3$ b. $Ba(HCO_3)_2$ c. LiClO d. $Sn(Cr_2O_7)_2$
- 62. NH₄NO₃, ammonium nitrate (NH₄)₂CO₃, ammonium carbonate NH4CN, ammonium cyanide (NH₄)₃PO₄, ammonium phosphate Sn(NO3)4, tin(IV) nitrate

- 51. Identify the number and kinds of atoms pits ent in a molecule of each compound. 6.2 a. ascorbic acid (vitamin C), C6H806
 - b. monosodium glutamate (MSG), C₅H₄O₄V₂
 - c. sucrose (table sugar), C12H22O11
 - d. trinitrotoluene (TNT), C7H5N3O6
 - e. ammonium nitrate (fertilizer), NH,NO,
- 52. The melting point of a compound is 1240 r Is this compound an ionic or a molecular compound? 6.2
- 53. Write the symbol for each ion. Be sure to include the charge. 6.3
 - a. oxide ion d. nitride ion
 - b. lead(II) ion e. cupric ion
 - 1. fluoride ion c. lithium ion
- 54. Name the following ions. Use Table 6.3 if necessary. 6.3
 - a. Ba2+ e. P3c. Ag*
 - d. Hg²⁺ 1. Sn4+ b. I-
- 55. Write the names and formulas of the two putatomic anions in Table 6.4 with names that do not end in -ite or -ate. 6.3
- 56. Without consulting Table 6.4, name the follow ing ions. 6.3
 - e. HPO42a. OH
 - b. Pb4+ 1. Cr₂O₇²⁻
 - c. SO42g. Al3*
 - d. O2h. ClO₂⁻
- 57. What is the net ionic charge of every ionic compound? Explain. 6.4
- 58. Write formulas for compounds composed of these pairs of ions. 6.4
 - c. K*, Cla. Na*, O2-
 - b. Sn⁴⁺, S²⁻ d. Mg2+, N3-
- 59. How do you determine the charge of a transition metal cation from the formula of an ionic conground? 6.4
- 68 When must parentheses be used in a formal
- fr. formulas for these compounds. 64 ate tan carbonate 2
 - an hydrogen carbonate 51
 - e lit : am hypochlorite
 - 6 distriction dichromate
- 62. Complete the table on the following page by writing correct formulas for the compounds formed by combining positive and negative ions. Then name each compound. 6.4

Sn(CO₃)₂, tin(IV) carbonate Sn(CN)4, tin(IV) cyanide Sn₃(PO₄)₄, tin(IV) phosphate Fe(NO₃)₃, iron(III) nitrate Fe₂(CO₃)₃, iron(III) carbonate Fe(CN)3, iron(III) cyanide FePO₄, iron(III) phosphate Mg(NO₃)₂, magnesium nitrate MgCO₃, magnesium carbonate Mg(CN)2, magnesium cyanide Mg₃(PO₄)₂, magnesium phosphate



	NO ₃ -	CO32-	CNT	P043-
NH4+				
Sn4+				
Fe ³⁺				
Mg ²⁺				

- 63. What are the components of a binary molecular compound? 6.5
- 64. Write the formula or name for these compounds. 6.5
 a. boron trichloride c. N₂O₅

b. dinitrogen tetrahydride d. CCl₄

65. Give the name or the formula for these acids. 6.5 a. HCl c. sulfuric acid

b. HNO₃ d. acetic acid

66. Which prefix indicates each of the following numbers of atoms in a molecular compound formula? 6.5

a.3 b.1 c.2 d.6 e.5 f.4

CONCEPT MASTERY

67. Name these compounds.

a. NaClO3	d. AlI ₃	g. KHSO4
b. Hg ₂ Br ₂	$e. SnO_2$	h. CaH ₂
c. K ₂ CrO ₄	f. $Fe(C_2H_3O_2)_3$	

68. Write formulas for these compounds.

- a. potassium permanganate
- b. calcium hydrogen carbonate
- c. dichlorine heptoxide
- d. trisilicon tetranitride
- e. sodium dihydrogen phosphate
- f. phosphorus pentabromide
- g. carbon tetrachloride
- Name each substance. Use Figure 6.21 if necessary.

a. LiClO₄ d. CaO g. SrSO₄ b. Cl₂O e. Ba₃(PO₄)₂ h. CuC₂H₃O₂ c. HgF₂ f. I₂

70. Write formulas for these compounds.

a. magnesium sulfidee. sulfite ionb. nitrogen gasf. calcium carbonatec. barium hydroxideg. sodium bromided. copper(II) nitriteh. ferric sulfate

71. Name each compound.

a. Mg(MnO ₄) ₂	$d. N_2H_4$	g . PI ₃
b . Be(NO ₃) ₂	e. LiOH	h. ZnO
c. K ₂ CO ₃	f. BaF ₂	

- 63. two nonmetals
- 64. a. BCl₃
 b. N₂H₄
 c. dinitrogen pentoxide
 d. carbon tetrachloride
 65. a. hydrochloric acid
 b. nitric acid
 c. H₂SO₄
 d. HC₂H₃O₂
- 66. a. tri- d. hexa-
- b. mono- e. penta
 - c. di- f. tetra-
- 67. a. sodium chlorate
- **b**. mercury(I) bromide
 - c. potassium chromate

72. The United States produces thousands of different kinds of inorganic chemicals. The following data table shows the amounts (in billions of kg) of the top ten inorganic chemicals produced in a recent year.

Chemical	Amount produced (billions of kg)
Sulfuric acid	39.4
Nitrogen	26.9
Oxygen	17.7
Ammonia	16.5
Lime	16.3
Phosphoric acid	11.2
Sodium hydroxide	11.0
Chlorine	10.3
Sodium carbonate	9.3
Nitric acid	6.8

- a. For what percentage of the total production of the top ten did lime (calcium oxide) account?
- b. Three diatomic gases are on the list. What are their names and what was their total production?
- c. For what percentage of the total production of the top ten did the three acids account?
- Write formulas for the top ten inorganic chemicals.

73. Write formulas for these compounds.

e. tin(IV) cyanide
f. lithium hydride
g. strontium acetate
h. sodium silicate

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- - b. plant:oak tree anion:______
 (1) sodium ion (3) cation

(1) sodium ion	(3) cation
(2) magnesium ion	(4) chloride ion
c. molecular compound	i:molecule
ionic compound:	
(1) cation	(3) anion

Chemical Names and Formulas 167

(4) metalloid

- d. aluminum iodide
- e. tin(IV) oxide

(2) formula unit

- f. iron(III) acetate
- g. potassium hydrogen sulfate
- h. calcium hydride

e. NaH ₂ PO ₄	
f. PBr ₅	
g. CCl ₄	

- d. Si₃N₄
- 69. a. lithium perchlorate
 - b. dichlorine monoxide

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REVIEW

- c. mercury(II) fluoride d. calcium oxide e. barium phosphate f. iodine g. strontium sulfate h. copper(I) acetate e. SO32-70. a. MgS b. N2 f. CaCO₃ **c**. $Ba(OH)_2$ q. NaBr d. $Cu(NO_2)_2$ **h**. $Fe_2(SO_4)_3$ 71. a. magnesium permanganate b. beryllium nitrate c. potassium carbonate d. dinitrogen tetrahydride e. lithium hydroxide f. barium fluoride g. phosphorus triiodide h. zinc oxide 72. a. 9.85% b. nitrogen, oxygen, chlorine; 54.9 Kg C. 34.7% d. H₂SO₄, N₂, O₂, NH₃, CaO, H₃PO₄, NaOH, Cl₂, Na₂CO₃, HNO₃ 73. a. CaBr₂ e. Sn(CN)4 f. LiH b. AgCl **c.** Al_4C_3 **g**. $Sr(C_2H_3O_2)_2$ \mathbf{d} . NO₂ h. Na₂SiO₃
- 74. a. 1 b. 4 c. 2

15-1 and 15-2 Homework Answer Key

15.1 pg. 418 #1 to 6

#1) The group number equals the number of valence electrons for representative elements.

#2) It is easier for a metal to lose electrons and for a nonmetal to gain electrons to achieve the electron configurations of a noble gas



15.2 pg. 421 #7 and 8; pg. 425 #9 to 13

#7) a. Kl b. Al_2O_3

#8) a. Potassium iodide b. Aluminum oxide

#9) characterized by attraction between oppositely charged ions formed through electron transfer.

#10) Ionic compounds conduct electricity when melted and in aqueous solution because their ions are free to move about.

#11) a.	K_2S	b. CaO	c. Na ₂ SO ₄	d. AIPO ₄
#12) a.	KNO ₃	b. $BaCl_2$	c. MgSO ₄	
d.	Li ₂ O	e. $(NH_4)_2CO_3$	f. Ca ₃ (PO ₄) ₂	

#13) c. and d.

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Answers

- 20. electrons in the highest occupied energy level
- **21.** fluorine (F), chlorine (Cl), bromine (Br), iodine (I); Group 7A; seven valence electrons
- **22. a**. 7; 5A **b**. 3; 1A
 - **c**. 15; 5A
 - **d**. 56; 2A
- 23. a. :Cl
 - **b.** :S·
 - C. Al.
 - d. Li
- 24. Their outermost occupied energy levels are filled.
- **25.** a. 2 b. 3 c. 1 d. 2
- 26. a. Al³⁺
 - b. Li⁺
 - **C**. Ba²⁺
 - **d**. K⁺
 - e. Ca2+
 - f. Sr²⁺
- 27. a. 1s²2s²2p⁶3s²3p⁶3d³
 b. 1s²2s²2p⁶3s²3p⁶3d⁴
 c. 1s²2s²2p⁶3s²3p⁶3d⁵
- **28.** Most nonmetals gain 1, 2, or 3 electrons to achieve a noble-gas electron configuration.
- **29.** a. S²⁻ b. Na⁺ c. F⁻ d. P³⁻
- **30.** a. 3 b. 2 c. 1 d. 3
- **31. a**. Br⁻ **b**. H⁻ **c**. As³⁻ **d**. Se²⁻ **32.** All are $ls^2 2s^2 2p^6$. All have the
- same configuration as neon.33. The positive charges of the cations equal the negative
- charges of the anions. 34. a., b., and d.
- **35. a**. K⁺Cl⁻ **b**. Ba²⁺SO₄²⁻
 - C. Mg²⁺Br⁻
 - **d**. $\text{Li}^+\text{CO}_3^{2-}$
- **36.** No; the packing of ions in a crystalline structure depends on a number of factors, including the relative sizes of the ions. The coordination number of an element can vary from compound to compound.
- **37.** Their network of electrostatic attractions and repulsions forms a rigid structure.

Chapter 15 REVIEW

CONCEPT PRACTICE

- 20. Define valence electrons. 15.1
- 21. Name the first four halogens. What group are they in, and how many valence electrons does each have? 15.1
- 22. How many electrons does each atom have? What group is each in? 15.1
 - a. nitrogen c. phosphorus
 - b. lithium d. barium
- 23. Write electron dot structures for each of the following elements. 15.1
 a. Cl b. S c. Al d. Li
- 24. The atoms of the noble gas elements are stable. Explain. 15.1
- 25. How many electrons must each atom lose to attain a noble-gas electron configuration? 15.1
 a. Ca b. Al c. Li d. Ba
- **26.** Write the formula for the ion formed when each of the following elements loses its valence electrons. *15.1*
 - a. aluminum c. barium e. calcium b. lithium d. potassium f. strontium
- 27. Write electron configurations for the tripositive ions (3+) of these elements. 15.1
 - a. chromium b. manganese c. iron
- 28. Why do nonmetals tend to form anions when they react to form compounds? 15.1
- 29. What is the formula of the ion formed when the following elements gain or lose valence electrons and attain noble-gas configurations? 2014
 a. sulfur
 c. fluorine
 - b. sodium d. phosphorus
- How many electrons must be gained by each of the following atoms to achieve a stable electron configuration? 15.1
 - a. N b. S c. Cl d. P
- 31. Write the formula for the ion formed when each element gains electrons and attains a noble-gas configuration. 15.1
 a. Br b. H c. As d. Se
- 32. Write electron configurations for the following and comment on the result. 15.1
 a. N³⁻ b. O²⁻ c. F⁻ d. Ne
- **33**. Explain why ionic compounds are electrically neutral. *15.2*

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- 34. Which of the following pairs of elements in not form ionic compounds? 15.2
 - a. sulfur and oxygen
 - **b**. sodium and calcium **c**. sodium and sulfur
 - d. oxygen and chlorine
- **35.** Write the formula for the ions in the following compounds. *15.2*
 - a. KCl b. BaSO4 c. MgBr2 d. Li2CO,
- Can you predict the coordination number of a ion from the formula of an ionic compound Explain. 15.2
- 37. Most ionic substances are brittle. Why? 152
- Explain why molten MgCl₂ does conduct an electric current although crystalline MgCl₂ does not. 15.2
- Explain briefly why metals are good conducting of electricity. 15.3
- 40. Name the three crystal arrangements of closely packed metal atoms. Give an example of a treat that crystallizes in each arrangement. 15.3
- 41. Name some alloys that you have used or seen today. 15.3
- 42. Why aren't the properties of all steels identical Explain. 15.3
- 43. The properties of all samples of brass areast identical. Explain. 15.3

CONCEPT MASTERY

- 44. Which of the following substances are most likely not ionic?
 - a. H_2O c. CO_2 e. NH_3
 - **b**. Na_2O **d**. CaS **f**. SO_2
- 45. Construct a table that shows the relationship among the group number, valence electrons lost or gained, and the formula of the cation or anion produced for the following metallic and nonmetallic elements: Na, Ca, Al, N. S. R.
- Write electron dot formulas for the following atoms.
 - a.C b.Be c.O d.F e.Na 1.P
- 47. Show the relationship between the electron dot structure of an element and the location of the element in the periodic table.
- **48.** In terms of electrons, why does a cation have a positive charge?

- **38.** Ions are free to move in molten $MgCl_2$.
- **39.** They have many mobile valence electrons. Electrons in the current replace the electrons leaving the metal.
- **40.** body-centered cubic: Na, K, Fe, Cr, or W face-centered cubic: Cu, Ag, Au, Al, or Pb hexagonal closest-packed: Mg, Zn, or Cd
- 41. Answers will vary and could include tableware, steel in cars and buses, high-speed dental drill bits, solder in stereos and televisions, and structural steel in buildings.
- **42.** The properties of steel vary according to its composition. In addition to iron, steel can contain varying amounts of carbon and such metals as chromium, nickel, and molyb-denum.
- **43.** Brass is a mixture of copper and zinc. The properties of a particular sample of brass not vary with the relative proportions of the not metals.
- 44. a., c., e., and f.

- 49. Why does an anion have a negative charge?
- 50. Metallic cobalt crystallizes in a hexagonal close-packed structure. How many neighbors will a cobalt atom have?
- 51. Write electron configurations for the dipositive ions (2+) of these elements.
 a. Fe b. Co c. Ni
- 52. Write electron configurations for these atoms and ions, and comment on the result.
 a. Ar b. Cl⁻ c. S²⁻ d. P³⁻
- Explain how hexagonal close-packed, facecentered cubic, and body-centered cubic unit cells are different from one another.
- 54. The spheres below represent the relative diameters of atoms or ions. Rearrange the sequences in a. and b. so the relative sizes of the particles correspond to the increasing size of the particles in the illustration.



- a. oxygen atom, oxide ion, sulfur atom, sulfide ion
- sodium atom, sodium ion, potassium atom, potassium ion
- 55. Write complete electron configurations for the following atoms and ions. For each group, comment on the results.

a. Ar, K⁺, Ca²⁺ b. Ne, Na⁺, Mg²⁺, Al³⁺

CRITICAL THINKING

56. Choose the term that best completes the second relationship.

a. cow : horse	ionic bond :
(1) anion	(3) chemical bond
(2) cation	(4) covalent bond
b. Cl : Cl ⁻	Mg :
(1) Mg^{2+}	(3) Al^{3+}
(2) Mg^{2-}	(4) Mn
c. pipe : water	metal :
(1) ions	(3) electricity
(2) ionic bond	(4) conductor

57. Compare the physical and chemical characteristics of metals and ionic compounds.

CUMULATIVE REVIEW

 Hydrogen and oxygen react to produce water according to this equation.

 $2H_2 + O_2 \longrightarrow 2H_2O$

How many liters of hydrogen at STP are needed to produce 0.50 mol H_2O ?

- Distinguish among gases, liquids, and solids with respect to shape, volume, relative density, and motion of particles.
- 60. What is the volume, in liters, occupied by 8.0 g of oxygen gas at STP?
- **61.** If you raise the temperature of a gas, what actually happens to the gas particles?
- 62. A gas occupies 750 cm³ at 27 °C and 1.6 kPa. Find its volume at STP.
- 63. Explain each of these observations on the basis of the kinetic theory and the forces of attraction that exist between the particles in matter.
 - a. Water evaporates faster at 40 °C than at 20 °C.
 - **b**. A burn from steam at 100 °C is worse than a burn from water at 100 °C.
 - c. A cap pops off a bottle of root beer that has been kept in the trunk of a car on a hot day.
 - **d**. Diethyl ether ($C_4H_{10}O$) has a vapor pressure of 58.9 kPa at 20 °C, whereas the vapor pressure of water at 20 °C is only 2.3 kPa.
 - e. A melting ice cube cools a glass of tea.
 - Foods are dehydrated (water is removed) by using low pressures rather than high heat.
- 64. Use the gas laws and kinetic theory to complete these statements. Unless otherwise stated, assume a constant amount of gas.
 - As the volume of a gas increases at constant temperature, its pressure _____.
 - b. As the temperature of a gas increases and its pressure decreases, its volume ______.
 - c. At constant pressure, a decrease in the volume of a gas is caused by a(n) ________ in its temperature.
 - **d**. An increase in the volume and pressure of a gas is caused by a(n) _____ in its temperature.
 - At constant volume, a decrease in the temperature of a gas causes the pressure to_____.
 - If the volume of a gas is increased while its temperature is increased, the pressure will ______.

Ionic Bonding and Ionic Compounds 433

- 45. Group
- 6A 7A number 2A 3A 5A 1A Valence electrons 2 1 3 lost or gained 1 2 3 Na⁺ Ca²⁺ Al³⁺ N³⁻ S²⁻ Br⁻ lon formula 46. a. C d. :F. b. Bee. Na-C. :Ö. f. P.

47. For the representative elements, the number

of electrons in the electron dot structure is the group number.

- 48. It has lost valence electrons.
- 49. It has gained valance electrons.
- 50. 12
- 51. a. 1s²2s²2p⁶3s²3p⁶3d⁶
 - **b**. 1s²2s²2p⁶3s²3p⁶3d⁷ **c**. 1s²2s²2p⁶3s²3p⁶3d⁸
- **52.** All have the noble gas configuration of $1s^22s^22p^63s^23p^6$

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REVIEW

- **53.** Hexagonal closest-packed unit cells have 12 neighbors for every atom or ion. Face-centered cubic unit cells also have 12 neighbors for every atom or ion, with an atom or ion in the center of each face. Body-centered cubic units cells have 8 neighbors for every atom or ion, with an atom or ion at the center of each cube.
- 54. a. oxygen atom, sulfur atom, oxide ion, sulfide ion
 b. sodium ion, potassium ion, sodium atom, potassium atom
- **55. a.** $1s^22s^22p^63s^23p^6$ **b.** $1s^22s^22p^6$ Each has a noble-gas electron

configuration. 56. a. 4

b. 1

C. 3

- **57.** Both are composed of ions and are held together by electro-
- static bonds. Metals always conduct electricity, but ionic compounds conduct only when melted or in aqueous solution. Ionic compounds are composed of cations and anions, but metals are composed of cations and free-floating valence electrons. Metals are malleable, but ionic compounds are brittle.
- 58. 11 L
- 59. Gases and liquids assume the shapes of their containers. Solids have a definite shape. Liquids and solids have a definite volume; gases do not. Gases have a low density; liquids have an intermediate density; solids have a high density. The molecules in a gas move freely and randomly. The molecules of a liquid flow. The molecules of a solid vibrate and rotate around a fixed position.
- 60. 5.6 L
- **61.** The average kinetic energy and the speed of the gas molecules increase.
- **62.** 11 cm³

8.1 to 8.3 Homework Answer Key

8.1 pg. 206 #1, 2; pg. 209 #3, 4; pg. 21	0 #5 to 7; pg. 211 #8 to 11
#1) a. $S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$	b. $\text{KCIO}_{3 \text{ (s)}} \xrightarrow{MnO_2,\Delta} \text{KCI}_{(s)} + \text{O}_{2 \text{ (g)}}$
#2) a. Mixing aqueous potassium hydroxide and su potassium sulfate.	Ifuric acid results in water and aqueous
b. When solid sodium is added to water, hydro	gen gas and aqueous sodium hydroxide is produced.
#3) a. 2 AgNO ₃ + H ₂ S \rightarrow Ag ₂ S + 2 HNO ₃ c. 3 Zn(OH) ₂ + 2 H ₃ PO ₄ \rightarrow Zn ₃ (PO ₄) ₂ + 6 H ₂ O	b. $MnO_2 + 4 HCI \rightarrow MnCI_2 + 2 H_2O + CI_2$
#4) a. $H_2 + S \rightarrow H_2S$	b. 2 FeCl ₃ + 3 Ca(OH) ₂ \rightarrow 2 Fe(OH) ₃ + 3 CaCl ₂
#5) 3 CO + Fe ₂ O ₃ \rightarrow 2 Fe + 3 CO ₂	
#6) 2 C + $O_2 \rightarrow 2$ CO	
#7) a. FeCl ₃ + 3 NaOH → Fe(OH) ₃ + 3 NaCl c. CH ₄ + Br ₂ → CH ₃ Br + HBr	b. $CS_2 + 3 Cl_2 \rightarrow CCl_4 + S_2Cl_2$
#8) a. 2 Na + 2 H ₂ O \rightarrow 2NaOH + H ₂	b. $Ca(OH)_2 + H_2SO_4 \rightarrow CaSO_4 + 2 H_2O$
 #9) a. copper (II) sulfide + oxygen → copper + sulfu c. sodium hydrogen carbonate → sodium carbo 	ur dioxide b. hydrogen + oxygen → water onate + carbon dioxide + water
#10) a. $2 \text{ SO}_2 + \text{O}_2 \rightarrow 2 \text{ SO}_3$	b. $Fe_2O_3 = 3 H_2 \rightarrow 2 Fe + 3 H_2O$
c. 4 P + 5 $O_2 \rightarrow P_4 O_{10}$	d. 2 AI + $N_2 \rightarrow 2$ AIN
#11) a. SO _{3 (g)} b. KNO _{3 (aq)} c. $-\Delta$	$\rightarrow \qquad \text{d. } Cu_{(s)} \qquad \text{e. } Hg_{(l)} \qquad \text{f. } \xrightarrow{ZnCl_2}$
8.2 pg. 214 #13, 14; pg. 216 #15, 16; p	g. 218 #17;
pg. 220-221 #18 to 21; pg.224 #2 #13) a 2 Be + $\Omega_2 \rightarrow 2$ BeO	$2, 23 \text{ (OMIT 23C)}$ $b SO_2 + H_2O \rightarrow H_2SO_2$
#14) a. Sr + $I_2 \rightarrow SrI_2$	b. 3 Mg + $N_2 \rightarrow Mg_3N_2$
#15) a. 2 HI → H ₂ + I ₂	b. $Mg(CIO_3)_2 \rightarrow MgCI_2 + 3 O_2$

#16) a. HBr b. NaCl

b. $CI_{2(q)} + 2 NaI_{(aq)} \rightarrow 2 NaCI_{(aq)} + I_{2(s)}$ #17) a. Fe $_{(s)}$ + Pb(NO₃)_{2 (aq)} \rightarrow Fe(NO₃)_{2 (aq)} + Pb $_{(s)}$ c. Ca $_{(s)}$ + 2 H₂O $_{(l)} \rightarrow$ Ca(OH)_{2 (aq)}+ H_{2 (g)} #18) a. 3 NaOH + Fe(NO₃)₃ \rightarrow Fe(OH)₃ + 3 NaNO₃ b. 3 Ba(NO₃)₃ + 2 H₃PO₄ \rightarrow Ba₃(PO₄)₂ + 6 HNO₃ #19) a. 3 KOH + $H_3PO_4 \rightarrow K_3PO_4$ + 3 HOH b. $3 H_2SO_4 + 2 AI(OH)_3 \rightarrow AI_2(SO_4)_3 + 6 HOH$ b. $C_7H_{16} + 11 O_2 \rightarrow 7 CO_2 + 8 H_2O$ #20) a. 2 HCOOH + $O_2 \rightarrow 2 CO_2 + 2 H_2O$ #21) $C_6H_{12}O_6 + 6 O_2 \rightarrow 6 CO_2 + 6 H_2O$ #22) a. Pb(NO₃)₂ + K₂CrO₄ \rightarrow PbCrO₄ + 2 KNO₃, double replacement b. $CI_2 + 2 KI \rightarrow 2 KCI + I_2$, single replacement c. $2 C_3 H_6 + 9 O_2 \rightarrow 6 CO_2 + 6 H_2 O_3$, hydrocarbon combustion d. 2 AI(OH)₃ \rightarrow AI₂O₃ + 3 H₂O, decomposition e. 4 Li + $O_2 \rightarrow 2$ Li₂O, combination f. 6 HCl + Fe₂O₃ \rightarrow 2 FeCl₃ + 3 H₂O, double replacement g. MgCO₃ \rightarrow MgO + CO₂, decomposition h. $Ba(CN)_2 + H_2SO_4 \rightarrow BaSO_4 + 2 HCN$, double replacement b. 2 AI + 3 $CI_2 \rightarrow 2 AICI_3$ #23) a. $Cal_2 + Hg(NO_3)_2 \rightarrow Hgl_2 + Ca(NO_3)_2$ d. $2 C_2H_2 + 5 O_2 \rightarrow 4 CO_2 + 2 H_2O$ e. MgCl₂ $\xrightarrow{electricity}$ Mg + Cl₂ f. $H_2O_2 \rightarrow 2 H_2O + O_2$

8.3 pg. 226 #25, pg. 228 #26 to 31

#25) a. $Pb^{2+}_{(aq)} + 2 I^{-}_{(aq)} \rightarrow PbI_{2 (aq)}$ the c. $H^{+}_{(aq)} + OH^{-}_{(aq)} \rightarrow HOH_{(l)}$ or $H_2O_{(l)}$

b.
$$Zn_{(s)} + 2 H^{+}_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

- #26) $Pb^{2+}(aq) + 2 Cl^{-}(aq) \rightarrow PbCl_{2(s)}$
- #27) Complete Ionic: $Fe^{3^{+}}_{(aq)} + 3 NO_{3}^{1^{-}}_{(aq)} + 3 Na^{+}_{(aq)} + 3 OH^{-}_{(aq)} \rightarrow Fe(OH)_{3 (s)} + 3 Na^{+}_{(aq)} + 3 NO_{3}^{1^{-}}_{(aq)}$ Net-Ionic: $Fe^{3^{+}}_{(aq)} + 3 OH^{-}_{(aq)} \rightarrow Fe(OH)_{3 (s)}$
- #28) a. $Pb^{2+}_{(aq)} + SO_4^{2-}_{(aq)} \rightarrow PbSO_{4 (s)}$ b. $Pb^{2+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow PbCl_{2 (s)}$ c. $Fe^{3+}_{(aq)} + PO_4^{3-}_{(aq)} \rightarrow FePO_{4 (s)}$ d. $Co^{2+}_{(aq)} + S^{2-}_{(aq)} \rightarrow CoS_{(s)}$
- #29) a. $Ag^{+}_{(aq)} + CI^{-}_{(aq)} \rightarrow AgCI_{(s)}$ b. $Pb^{2+}_{(aq)} + CI^{-}_{(aq)} \rightarrow PbCI_{2(s)}$ c. $Cr^{3+}_{(aq)} + PO_{4}^{3-}_{(aq)} \rightarrow CrPO_{4(s)}$
- #30) a. H^+ and NO_3^{1-} b. Li^+ and $C_2H_3O_2^{1-}$ (or CH_3COO^{1-}) c. Na^+ and Cl^{1-}
- #31) a. BaSO_{4 (s)} b. Al(OH)_{3 (s)} c. Ag₂S _(s) d. PbCl_{2 (s)} e. CaCO_{3 (s)}

MIRITIAN

Answers

32. a. reactants: sodium and water; products: hydrogen and sodium hydroxide

b. reactants: carbon dioxide and water; products: oxygen, glucose

- 33. Dalton said that the atoms of reactants are rearranged to form new substances as products.
- 34. The arrow separates the reactants from the products and indicates a reaction that progresses in a forward direction. A plus sign separates individual reactants and individual products from one another.
- 35. a. Gaseous ammonia and oxygen react in the presence of platinum to produce nitrogen monoxide gas and water vapor b. Aqueous solutions of sulfuric acid and barium chloride are mixed to produce a precipitate of barium sulfate and aqueous hydrochloric acid.

c. The gas dinitrogen trioxide reacts with water to produce an aqueous solution of nitrous acid.

- 36. A catalyst speeds up a chemical reaction.
- **37. a.** $C + 2F + 2G \rightarrow CF_2G_2$ **b**. $F + 3W + S + 2P \rightarrow FW_3SP_2$
- 38. It is incorrect because a formula is a unique identifier of a substance. A different formula would indicate a different substance, not the one that is taking part in the reaction you are trying to balance.
- **39.** a. $2PbO_2 \rightarrow 2PbO + O_2$ **b**. $2Fe(OH)_3 \rightarrow Fe_2O_3 + 3H_2O$ c. $(NH_4)_2CO_3 \rightarrow 2NH_3 + H_2O +$ CO_2 d. 2NaCl + $H_2SO_4 \rightarrow Na_2SO_4$ **e**. $4H_2 + Fe_3O_4 \rightarrow 3Fe + 4H_2O$ f. 2Al + 3CuSO₄ \rightarrow Al₂(SO₄)₃ + 3Cu
- 40. It helps in predicting products of reactions.
- 41. Use ionic charges to write an electrically neutral formula.
- 42. a single product
- 43. a. $2Mg + O_2 \rightarrow 2MgO$ **b**. $4P + 5O_2 \rightarrow 2P_2O_5$

Chapter 8 REVIEW

CONCEPT PRACTICE

- 32. Identify the reactants and products in each chemical reaction. 8.1
 - a. Hydrogen gas and sodium hydroxide are formed when sodium is dropped into water.
 - b. In photosynthesis, carbon dioxide and water react to form oxygen gas and glucose.
- 33. How did John Dalton explain a chemical reaction using his atomic theory? 8.1
- 34. What is the function of an arrow (\rightarrow) and a plus sign (+) in a chemical equation? 8.1
- 35. Write sentences that completely describe each of the chemical reactions shown in these skeleton equations. 8.1
 - a. $NH_3(g) + O_2(g) \xrightarrow{Pt} NO(g) + H_2O(g)$ b. H SO(gg) + BaCL(gg)b. H-SO

$$H_2SO_4(aq) + BaCl_2(aq) \longrightarrow BaSO_4(s) + HCl(aq)$$

c.
$$N_2O_3(g) + H_2O(l) \longrightarrow HNO_2(aq)$$

36. What is the purpose of a catalyst? 8.1

37. Balance equations for each item. The formula for each product (object) is given. 8.1 a. a basketball team center + forward + guard \longrightarrow team

 $C + F + G \longrightarrow CF_2G_2$ b. a tricycle

frame + wheel + seat + pedal \longrightarrow tricycle $F + W + S + P \longrightarrow FW_3SP_2$

- 38. The equation for the formation of water from its elements, $H_2(g) + O_2(g) \longrightarrow H_2O(l)$, can easily be "balanced" by changing the formula of the product to H2O2. Explain why this is incorrect. 8.1
- 39. Balance the following equations. 8.1
 - a. $PbO_2 \longrightarrow PbO + O_2$
 - **b**. $Fe(OH)_3 \longrightarrow Fe_2O_3 + H_2O$
 - c. $(NH_4)_2CO_3 \longrightarrow NH_3 + H_2O + CO_2$
 - d. NaCl + $H_2SO_4 \longrightarrow Na_2SO_4 + HCl$
 - $e. H_2 + Fe_3O_4 \longrightarrow Fe + H_2O$
 - f. Al + CuSO₄ \longrightarrow Al₂(SO₄)₃ + Cu
- 40. Explain why it is useful to classify reactions by their type. 8.2
- 41. How do you predict the correct formula for the combination reaction between a nonmetal and a Group A metal? 8.2
- 42. What is a characteristic of every combination reaction? 8.2

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c. Ca + S \rightarrow CaS

- **d**. 2Fe + $O_2 \rightarrow 2FeO$
- e. $N_2O_5 + H_2O \rightarrow 2HNO_3$
- 44. a. $2Ag_2O \rightarrow 4Ag + O_2$
- 46. a. sodium displaces iron b. copper displaces
- 47. a. no reaction
 - **b.** $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

- 43. Write balanced chemical equations for the following combination reactions. 8.2
 - a. Mg + $O_2 \longrightarrow$
 - **b.** $P + O_2 \longrightarrow$ diphosphorus pentoxide **c.** $Ca + S \longrightarrow$

 - d. Fe + $O_2 \longrightarrow iron(II)$ oxide e. $N_2O_5 + H_2O \longrightarrow$
- 44. Write a balanced chemical equation for each decomposition reaction. 8.2
 - a. Ag₂O(s) $\xrightarrow{\Delta}$?
- **b.** nickel(II) carbonate $\xrightarrow{\Delta}$
- nickel(II) oxide + ? **c**. ammonium nitrate $\xrightarrow{\Delta}$ dinitrogen monoxide + wate
- 45. What is a distinguishing feature of every decomposition reaction? 8.2
- 46. For each of the following pairs, predict which element as an atom would displace the other element as an ion from a compound in aqueous solution. 8.2
 - a. iron and sodium
 - b. silver and copper
 - c. zinc and hydrogen (in HCl)
- 47. Use the activity series of metals to write a balanced chemical equation for each singlereplacement reaction. 8.2
 - a. Au(s) + KNO₃(aq) \longrightarrow
 - b. $Zn(s) + AgNO_3(aq) \longrightarrow$
 - c. Al(s) + H₂SO₄(aq) \longrightarrow
 - $d. Cu(s) + H_2O(l) \longrightarrow$
 - e. Al(s) + CuSO₄(aq) \longrightarrow
- 4. Write a balanced equation for each of the following reactions. 8.2
 - a. $HCl(aq) + Ca(OH)_2(aq) \longrightarrow$
 - $\mathfrak{d}_{\mathrm{Ag}_{2}}\mathrm{SO}_{4}(aq) + \mathrm{AlCl}_{3}(aq) \longrightarrow$ (Silver chloride is a precipitate.)
 - c. $H_2C_2O_4(aq) + KOH(aq) \longrightarrow$
 - d. CdBr₂(aq) + Na₂S(aq) \longrightarrow
 - (Cadmium sulfide is a precipitate.)
- 49. What substance is common to all combustion reactions? 8.2
- 50. Write a balanced equation for the complete combustion of each compound. 8.2
 - a. butene (C4H8)
 - b. octane (C₈H₁₈)
 - c. glycerol (C₃H₈O₃)
 - d. acetone (C₃H₆O)

c. 2Al(s) + 3H₂SO₄(aq) → Al₂(SO₄)₃(aq) + 3H₂(g)

- d. no reaction
- **e.** $2AI(s) + 3CuSO_4(aq) \rightarrow AI_2(SO_4)_3(aq) + 3Cu(s)$
- **48.** a. $2HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) +$ 2H2O(1)
 - **b.** $3Ag_2SO_4(aq) + 2AICl_3(aq) \rightarrow 6AgCl(s) +$
- $Al_2(SO_4)_3(aq)$
 - **c.** $H_2C_2O_4(aq) + 2KOH(aq) \rightarrow K_2C_2O_4(aq) +$ 2H2O(1)
- **d.** CdBr₂(aq) + Na₂S(aq) \rightarrow CdS(s) + 2NaBr(aq) 49. oxygen

- **b.** NiCO₃ \rightarrow NiO + CO₂ C. $NH_4NO_3 \rightarrow N_2O + 2H_2O$
- 45. a single reactant
 - silver C. zinc displaces hydrogen

51. Balance each equation and identify its type. 8.2 a. Hf + N₂ \longrightarrow Hf₃N₄ **b**. Mg + H₂SO₄ \longrightarrow MgSO₄ + H₂ $\mathbf{c}. C_2 \mathbf{H}_6 + \mathbf{O}_2 \longrightarrow \mathbf{CO}_2 + \mathbf{H}_2 \mathbf{O}$

- d. $Pb(NO_3)_2 + NaI \longrightarrow PbI_2 + NaNO_3$ e. Fe + $O_2 \longrightarrow Fe_3O_4$
- 1. $Pb(NO_3)_2 \longrightarrow PbO + NO_2 + O_2$

 $g_{Hg}(NO_3)_2 + NH_4SCN \longrightarrow$

Hg(SCN)₂ + NH₄NO₃ h. $(NH_4)_2SO_4 + NaOH -$

 $NH_3 + H_2O + Na_2SO_4$ (Hint: There are two stepwise reactions in this equation.)

52. Complete each equation and then write a net ionic equation. 8.3 a. Al(s) + $H_2SO_4(aq)$ - \rightarrow

b. $HCl(aq) + Ba(OH)_2(aq) \longrightarrow$ c. Au(s) + HCl(aq) \longrightarrow

53. What is a spectator ion? 8.3

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- 54. Each of these equations is "balanced" but incorrect. Find the errors and correctly balance each equation. a. $Cl_2 + NaI \longrightarrow NaCl_2 + I$ $b. NH_3 \longrightarrow N + H_3$ c. Na + $O_2 \longrightarrow NaO_2$
 - $\begin{array}{l} \text{d. 2Mg} + \text{H}_2\text{SO}_4 \longrightarrow \text{Mg}_2\text{SO}_4 + \text{H}_2 \\ \text{e. MgCl} + \text{CaOH} \longrightarrow \text{MgOH} + \text{CaOl} \\ \end{array}$

 - $f: H_2 + Cl_2 \longrightarrow H_2Cl_2$

55. Write a balanced chemical equation for each reaction. Use the necessary symbols from Table 8.1 to describe the reaction completely.

- a. Bubbling chlorine gas through a solution of potassium iodide gives elemental iodine and a solution of potassium chloride.
- b. Bubbles of hydrogen gas and aqueous iron(III) chloride are produced when metallic iron is dropped into hydrochloric acid.
- c. Solid tetraphosphorus decoxide reacts with water to produce phosphoric acid.
- d. Solid silver oxide can be heated to give silver and oxygen gas.
- e. lodine crystals react with chlorine gas to form solid iodine trichloride.
- 1. Mercury metal is produced by heating a mixture of mercury(II) sulfide and calcium oxide. Additional products are calcium sulfide and calcium sulfate.
- 50. a. $\mathrm{C_4H_8} + 6\mathrm{O_2} \rightarrow 4\mathrm{CO_2} + 4\mathrm{H_2O}$
 - **b.** $2C_8H_{18} + 250_2 \rightarrow 16CO_2 + 18H_2O$
 - **c**. $2C_3H_8O_3 + 7O_2 \rightarrow 6CO_2 + 8H_2O_3$
 - $\textbf{d}.\ \textbf{C}_3\textbf{H}_6\textbf{O} + 4\textbf{O}_2 \rightarrow 3\textbf{CO}_2 + 3\textbf{H}_2\textbf{O}$
- 51. a. $3Hf + 2N_2 \rightarrow Hf_3N_4$; combination **b**. Mg + $H_2SO_4 \rightarrow MgSO_4 + H_2$; singlereplacement
 - **c.** $2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$; combustion **d**. $Pb(NO_3)_2 + 2NaI \rightarrow PbI_2 + 2NaNO_3;$
 - double-replacement
 - e. 3Fe + $2O_2$ → Fe₃O₄; combination

- 56. Write balanced chemical equations for these double-replacement reactions that occur in aqueous solution.
 - a. Zinc sulfide is added to sulfuric acid.
 - b. Sodium hydroxide reacts with nitric acid.
 - c. Solutions of potassium fluoride and calcium nitrate are mixed.
- 57. Write a balanced chemical equation for each combination reaction.
 - a. sodium oxide + water →
 - **b.** hydrogen + bromine \longrightarrow
 - c. dichlorine heptoxide + water \longrightarrow
- 58. Write a balanced chemical equation for each single-replacement reaction that takes place in aqueous solution. Write "no reaction" if a reaction does not occur.
 - a. A piece of steel wool (iron) is placed in sulfuric acid.
 - b. Mercury is poured into an aqueous solution of zinc nitrate.
 - c. Bromine reacts with aqueous barium iodide.
- 59. Pieces of sodium and magnesium are dropped into separate water-filled test tubes (A and B). Tube A bubbles vigorously; Tube B does not.
 - a. Which tube contains the sodium metal? b. Write an equation for the reaction in the tube containing the sodium metal. What type of reaction is occurring in this tube?
- 60. Write a balanced equation for the complete combustion of each compound. Assume that the products are carbon dioxide and water.
 - a. octane (C8H18) b. glucose (C₆H₁₂O₆)
 - c. acetic acid (HC₂H₃O₂)
- 61. Write balanced chemical equations for these decomposition reactions.
 - a. Aluminum can be obtained from aluminum oxide with the addition of a large amount of electrical energy.
 - b. Heating tin(IV) hydroxide gives tin(IV) oxide and water.
 - c. Silver carbonate decomposes into silver oxide and carbon dioxide when it is heated.
- 62. Write a balanced net ionic equation for each reaction. The product that is not in solution is given.

a. $H_2C_2O_4 + KOH \longrightarrow$	$[H_2O]$
b . $CdBr_2 + Na_2S \longrightarrow$	[CdS]
c. NaOH + Fe(NO ₃) ₃ \longrightarrow	[Fe(OH ₃)]

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f. $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$; decomposition g. $Hg(NO_3)_2 + 2NH_4SCN \rightarrow Hg(SCN)_2 +$ 2NH4NO3; double-replacement h. $(NH_4)_2SO_4 + 2NaOH \rightarrow$ $2NH_3 + 2H_2O + Na_2SO_4;$ double-replacement then decomposition

- **52.** a. $2AI(s) + 6H^+(aq) \rightarrow 2AI^{3+}(aq) + 3H_2(g)$ **b.** $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ C. no reaction
- 53. an ion that does not participate in the reaction

54. a. $Cl_2 + 2NaI \rightarrow 2NaCl + I_2$ **b**. $2NH_3 \rightarrow N_2 + 3H_2$ **c.** $4Na + O_2 \rightarrow 2Na_2O$ **d**. Mg + H₂SO₄ \rightarrow MgSO₄ + H₂ e. MgCl₂ + Ca(OH)₂ \rightarrow Mg(OH)₂ + CaCl₂ f. $H_2 + Cl_2 \rightarrow 2HCl$ 55. a. $Cl_2(g) + 2KI(aq) \rightarrow I_2(s) +$ 2KCl(aq) **b.** 2Fe(s) + 6HCl(aq) \rightarrow $2\text{FeCl}_3(aq) + 3\text{H}_2(g)$ **c**. $P_4O_{10}(s)$ + 6 $H_2O(l)$ → 4H3PO4(aq) **d**. $2Ag_2O(s) \xrightarrow{\Delta} 4Ag(s) + O_2(g)$ **e.** $I_2(s) + 3Cl_2(g) \rightarrow 2ICl_3(s)$ f. $4HgS(s) + 4CaO(s) \rightarrow 4Hg(l)$ $+ 3CaS(s) + CaSO_4(s)$ **56.** a. $ZnS(aq) + H_2SO_4(aq) \rightarrow$ $H_2S(g) + ZnSO_4(aq)$ **b.** NaOH(aq) + HNO₃(aq) \rightarrow $H_2O(l) + NaNO_3(aq)$ **c.** 2KF(aq) + Ca(NO₃)₂(aq) → $CaF_2(s) + 2KNO_3(aq)$ 57. a. Na₂O(s) + H₂O(l) \rightarrow 2NaOH(aq) **b.** $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$ c. Cl₂O₇(l) + H₂O(l) → 2HClO₄(aq) **58.** a. $Fe(s) + H_2SO_4(aq) \rightarrow$ $FeSO_4(aq) + H_2(g)$ b. no reaction **c.** $Br_2(l) + Bal_2(aq) \rightarrow$ $BaBr_2(aq) + I_2(s)$ 59. a. A **b.** $2Na(s) + 2H_2O(l) \rightarrow$ $2NaOH(aq) + H_2(g)$ single-replacement **60.** a. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 +$ 18H2O **b.** $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 +$ 6H₂O c. $HC_2H_3O_2 + 2O_2 \rightarrow 2CO_2 +$ $2H_2O$ 61. a. $2Al_2O_3 \xrightarrow{\text{energy}} 4Al + 3O_2$ **b.** $Sn(OH)_4 \xrightarrow{\Delta} SnO_2 + 2H_2O$ c. $Ag_2CO_3 \rightarrow Ag_2O + CO_2$ 62. a. $H^+(aq) + OH^-(aq) \rightarrow H_2O(h)$

b. $\operatorname{Cd}^{2+}(aq) + \operatorname{S}^{2-}(aq) \to \operatorname{CdS}(s)$ **c.** 3OH⁻(aq) + Fe³⁺(aq) → Fe(OH)₃(s)