Answer Key

Chapter 3-4 Quiz Answer Key

1 Matter and Change

Section: Chemistry is a Physical Science

1. d	2. a
3. b	4. d
5. a	6. d
7. c	8. c
9. b	10. c

Section: Matter and its Properties

1. c	2. a
3. b	4. a
5. c	6. d
7. b	8. a
9. a	10. c

Section: Elements

1. c	2. c
3. a	4. c
5. d	6. a
7. b	8. c
9. d	10. b

2 Measurements and Calculations

Section: Scientific Method

1. d	2. a
3. c	4. a
5. d	6. d
7. c	8. d
9. b	10. b

Section: Units of Measurement

1. c	2. d
3. d	4. b
5. d	6. a
7. a	8. b
9. a	10. c

Section: Using Scientific Measurements

1. b	2. b
3. b	4. b
5. d	6. b
7. b	8. c
9. d	10. c

3 Atoms: The Building Blocks of Matter

Section: The Atom: From Philosophical Idea to Scientific Theory

1. b	2. c
3. a	4. d
5. c	6. d
7. b	8. c
9. d	10. b

Section: The Stucture of the Atoms

1. b	2. d
3. a	4. c
5. b	6. c
7. d	8. b
9. c	10. c

Section: Counting Atoms

1. c	2. c
3. a	4. d
5. b	6. b
7. c	8. a
9. b	10. c

4 Arrangement of Electrons in Atoms

Section: The Development of a New Atomic Model

1. a	2. b
3. c	4. b
5. c	6. a
7. d	8. c
9. b	10. a

Section: The Quantum Model of the Atom

1. c	2. a
3. c	4. b
5. a	6. b
7. d	8. c
9. b	10. c

Section: Electron Configurations

1. d	2. b
3. a	4. c
5. c	6. b
7. d	8. b
9. a	10. c

Name _____ Date _____ Class _____

Chapter 3 Study Guide Answer Key

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 1

SHORT ANSWER **Answer the following questions in the space provided.**

1. Why is Democritus's view of matter considered only an idea, while Dalton's view is considered a theory?

Democritus's idea of matter does not relate atoms to a measurable property, while Dalton's theory can be tested through quantitative experimentation.

2. Give an example of a chemical or physical process that illustrates the law of conservation of mass.

A glass of ice cubes will have the same mass when the ice has completely melted into liquid water, even though its volume will change. (Accept any reasonable process.)

3. State two principles from Dalton's atomic theory that have been revised as new information has become available.

Atoms are divisible into smaller particles called subatomic particles. A given element can have atoms with different masses, called isotopes.

4. The formation of water according to the equation

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

shows that 2 molecules (made of 4 atoms) of hydrogen and 1 molecule (made of 2 atoms) of oxygen produce 2 molecules of water. The total mass of the product, water, is equal to the sum of the masses of each of the reactants, hydrogen and oxygen. What parts of Dalton's atomic theory are illustrated by this reaction? What law does this reaction illustrate?

Atoms cannot be subdivided, created, or destroyed. Also, atoms of different elements combine in simple, whole-number ratios to form compounds. The reaction also illustrates the law of conservation of mass.

Name _____ Date _____ Class _____

SECTION 1 continued

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

5. 16 g If 3 g of element C combine with 8 g of element D to form compound CD, how many grams of D are needed to form compound CD_2?

6. A sample of baking soda, $NaHCO_3$, *always* contains 27.37% by mass of sodium, 1.20% of hydrogen, 14.30% of carbon, and 57.14% of oxygen.

a. Which law do these data illustrate?

the law of definite proportions

b. State the law.

A chemical compound contains the same elements in exactly the same proportions by mass regardless of the sample or the source of the compound.

7. Nitrogen and oxygen combine to form several compounds, as shown by the following table.

Compound	Mass of nitrogen that combines with 1 g oxygen (g)
NO	1.70
NO_2	0.85
NO_4	0.44

Calculate the ratio of the masses of nitrogen in each of the following:

2.0 **a.** $\frac{NO}{NO_2}$ 2.0 **b.** $\frac{NO_2}{NO_4}$ 4.0 **c.** $\frac{NO}{NO_4}$

d. Which law do these data illustrate?

the law of multiple proportions

Name _____ Date _____ Class _____

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 2

SHORT ANSWER **Answer the following questions in the space provided.**

1. In cathode-ray tubes, the cathode ray is emitted from the negative electrode, which is called the cathode.

2. The smallest unit of an element that can exist either alone or in molecules containing the same or different elements is the atom.

3. A positively charged particle found in the nucleus is called a(n) proton.

4. A nuclear particle that has no electrical charge is called a(n) neutron.

5. The subatomic particles that are least massive and most massive, respectively, are the electron and neutron.

6. A cathode ray produced in a gas-filled tube is deflected by a magnetic field. A wire carrying an electric current can be pulled by a magnetic field. A cathode ray is deflected away from a negatively charged object. What property of the cathode ray is shown by these phenomena?

   The particles that compose cathode rays are negatively charged.

7. How would the electrons produced in a cathode-ray tube filled with neon gas compare with the electrons produced in a cathode-ray tube filled with chlorine gas?

   The electrons produced from neon gas and chlorine gas would behave in the same way because electrons do not differ from element to element.

8. **a.** Is an atom positively charged, negatively charged, or neutral?

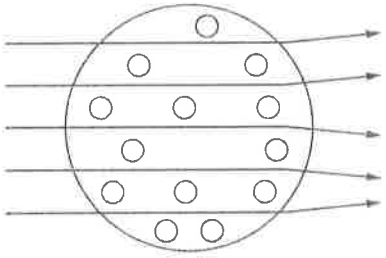
   Atoms are neutral.

   b. Explain how an atom can exist in this state.

   Atoms consist of a positively charged nucleus, made up of protons and neutrons, that is surrounded by a negatively charged electron cloud. The positive and negative charges combine to form a net neutral charge.

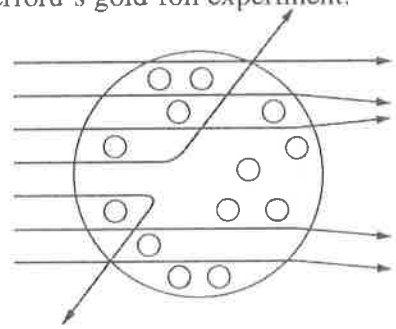
SECTION 2 continued

9. Below are illustrations of two scientists' conceptions of the atom. Label the electrons in both illustrations with a − sign and the nucleus in the illustration to the right with a + sign. On the lines below the figures, identify which illustration was believed to be correct before Rutherford's gold-foil experiment and which was believed to be correct after Rutherford's gold-foil experiment.



(Students should place a − sign inside all circles.)

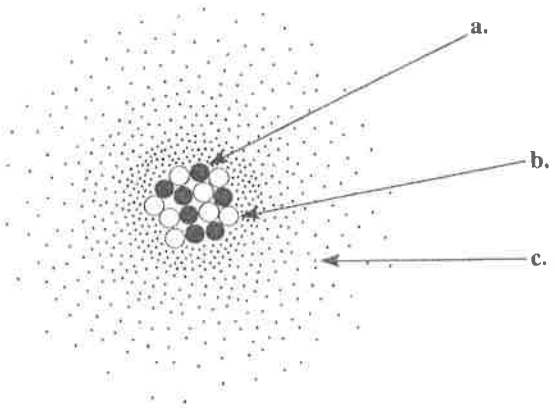
a. **before Rutherford's experiment**



(Students should place a + sign in the center and a − sign inside all circles.)

b. **after Rutherford's experiment**

10. In the space provided, describe the locations of the subatomic particles in the labeled model of an atom of nitrogen below, and give the charge and relative mass of each particle.



a. proton

The proton, a positive and relatively massive particle, should be located in the nucleus.

b. neutron

The neutron, a neutral and relatively massive particle, should be located in the nucleus.

c. electron (a possible location of this particle)

The electron, a negative particle with a low mass, should be located in the cloud surrounding the nucleus.

Name _____ Date _____ Class _____

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. Explain the difference between the *mass number* and the *atomic number* of a nuclide.

Mass number is the total number of protons and neutrons in the nucleus of an isotope. Atomic number is the total number of protons in the nucleus of each atom of an element.

2. Why is it necessary to use the average atomic mass of all isotopes, rather than the mass of the most commonly occurring isotope, when referring to the atomic mass of an element?

Elements rarely occur as only one isotope; rather, they exist as mixtures of different isotopes of various masses. Using a weighted average atomic mass, you can account for the less common isotopes.

3. How many particles are in 1 mol of carbon? 1 mol of lithium? 1 mol of eggs? Will 1 mol of each of these substances have the same mass?

There are 6.022×10^{23} particles in 1 mol of each of these substances. One mole of one substance will not necessarily have the same mass as one mole of another substance.

4. Explain what happens to each of the following as the atomic masses of the elements in the periodic table increase:

a. the number of protons

increases

b. the number of electrons

increases

c. the number of atoms in 1 mol of each element

stays the same

Name _____ Date _____ Class _____

SECTION 3 continued

5. Use a periodic table to complete the following chart:

Element	Symbol	Atomic number	Mass number
Europium-151	$^{151}_{63}Eu$	63	151
Silver-109	$^{109}_{47}Ag$	47	109
Tellurium-128	$^{128}_{52}Te$	52	128

6. List the number of protons, neutrons, and electrons found in zinc-66.

30 protons

36 neutrons

30 electrons

PROBLEMS **Write the answer on the line to the left. Show all your work in the space provided.**

7. 32.00 g What is the mass in grams of 2.000 mol of oxygen atoms?

8. 3.706 mol How many moles of aluminum exist in 100.0 g of aluminum?

9. 1.994×10^{24} atoms How many atoms are in 80.45 g of magnesium?

10. 1.993×10^{-21} g What is the mass in grams of 100 atoms of the carbon-12 isotope?

Name _____ Date _____ Class _____

CHAPTER 3 REVIEW

Atoms: The Building Blocks of Matter

MIXED REVIEW

SHORT ANSWER **Answer the following questions in the space provided.**

1. The element boron, B, has an atomic mass of 10.81 amu according to the periodic table. However, no single atom of boron has a mass of exactly 10.81 amu. How can you explain this difference?

The periodic table reports the average atomic mass, which is a weighted average of all isotopes of boron.

2. How did the outcome of Rutherford's gold-foil experiment indicate the existence of a nucleus?

A few alpha particles rebounded and therefore must have "hit" a dense bundle of matter. Because such a small percentage of particles were redirected, he reasoned that this clump of matter, called the nucleus, must occupy only a small fraction of the atom's total space.

3. Ibuprofen, $C_{13}H_{18}O_2$, that is manufactured in Michigan contains 75.69% by mass carbon, 8.80% hydrogen, and 15.51% oxygen. If you buy some ibuprofen for a headache while you are on vacation in Germany, how do you know that it has the same percentage composition as the ibuprofen you buy at home?

The law of definite proportions states that a chemical compound contains the same elements in exactly the same proportions by mass regardless of the site of the sample or the source of the compound.

4. Complete the following chart, using the atomic mass values from the periodic table:

Compound	Mass of Fe (g)	Mass of O (g)	Ratio of O:Fe
FeO	55.85	16.00	0.2865
Fe_2O_3	111.70	48.00	0.4297
Fe_3O_4	167.55	64.00	0.3820

Name _____ Date _____ Class _____

MIXED REVIEW continued

5. Complete the following table:

Element	Symbol	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
Sodium	Na	11	22	11	11	11
Fluorine	F	9	19	9	10	9
Bromine	Br	35	80	35	45	35
Calcium	Ca	20	40	20	20	20
Hydrogen	H	1	1	1	0	1
Radon	Rn	86	222	86	136	86

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

6. 1.51×10^{24} atoms **a.** How many atoms are there in 2.50 mol of hydrogen?

1.51×10^{24} atoms **b.** How many atoms are there in 2.50 mol of uranium?

4.65 mol **c.** How many moles are present in 107 g of sodium?

7. A certain element exists as three natural isotopes, as shown in the table below.

Isotope	Mass (amu)	Percent natural abundance	Mass number
1	19.99244	90.51	20
2	20.99395	0.27	21
3	21.99138	9.22	22

20 amu Calculate the average atomic mass of this element.

Name _____ Date _____ Class _____

Chapter 4 Study Guide Answer Key

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 1

SHORT ANSWER **Answer the following questions in the space provided.**

1. In what way does the photoelectric effect support the particle theory of light?

In order for an electron to be ejected from a metal surface, the electron must be struck by a single photon with at least the minimum energy needed to knock the electron loose.

2. What is the difference between the ground state and the excited state of an atom?

The ground state is the lowest energy state of the atom. When the atom absorbs energy, it can move to a higher energy state, or excited state.

3. Under what circumstances can an atom emit a photon?

A photon is emitted when an atom moves from an excited state to its ground state or to a lower-energy excited state.

4. How can the energy levels of the atom be determined by measuring the light emitted from an atom?

When an atom loses energy, it falls from a higher energy state to a lower energy state. The frequency of the emitted light, observed in an element's line-emission spectrum, may be measured. The energy of each transition is calculated using the equation $E = hv$, where v is the frequency of each of the lines in the element's line-emission spectrum. From the analysis of these results, the energy levels of an atom atom of each element may be determined.

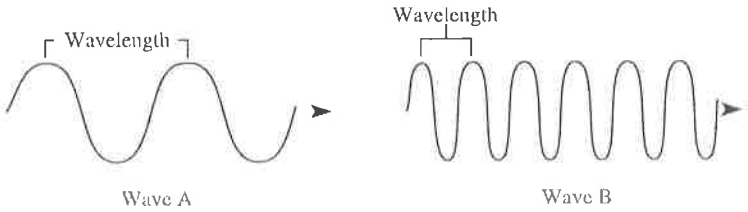
5. Why does electromagnetic radiation in the ultraviolet region represent a larger energy transition than does radiation in the infrared region?

Energy is proportional to frequency, and ultraviolet radiation has a higher frequency than infrared radiation. To produce ultraviolet radiation, electrons must drop to lower energy levels than they do to produce infrared radiation.

Name _____ Date _____ Class _____

SECTION 1 continued

6. Which of the waves shown below has the higher frequency? (The scale is the same for each drawing.) Explain your answer.



Wave B has the higher frequency. Wavelength is inversely proportional to frequency, so as the wavelength decreases, its frequency increases.

7. How many different photons of radiation were emitted from excited helium atoms to form the spectrum shown below? Explain your answer.

Spectrum for helium

Six different photons were emitted. Each time an excited helium atom falls back from an excited state to its ground state or to a lower energy state, it emits a photon of radiation that shows up as this specific line-emission spectrum. There are six lines in this helium spectrum.

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

8. 9.7×10^{14} Hz What is the frequency of light that has a wavelength of 310 nm?

9. 9.4×10^{9} m What is the wavelength of electromagnetic radiation if its frequency is 3.2×10^{-2} Hz?

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 2

SHORT ANSWER **Answer the following questions in the space provided.**

1. __d__ How many quantum numbers are used to describe the properties of electrons in atomic orbitals?

(a) 1
(b) 2
(c) 3
(d) 4

2. __a__ A spherical electron cloud surrounding an atomic nucleus would best represent

(a) an *s* orbital.
(b) a *p* orbital.
(c) a combination of two different *p* orbitals.
(d) a combination of an *s* and a *p* orbital.

3. __a__ How many electrons can an energy level of $n = 4$ hold?

(a) 32
(b) 24
(c) 8
(d) 6

4. __c__ How many electrons can an energy level of $n = 2$ hold?

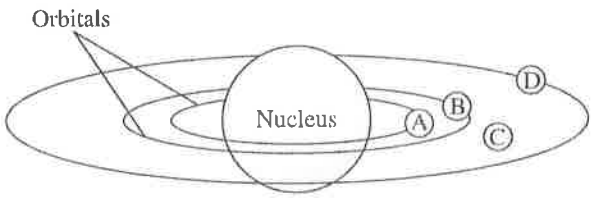
(a) 32
(b) 24
(c) 8
(d) 6

5. __c__ Compared with an electron for which $n = 2$, an electron for which $n = 4$ has more

(a) spin.
(b) particle nature.
(c) energy.
(d) wave nature.

6. __c__ According to Bohr, which is the point in the figure below where electrons cannot reside?

(a) point A
(b) point B
(c) point C
(d) point D



7. __c__ According to the quantum theory, point D in the above figure represents

(a) the fixed position of an electron.
(b) the farthest position from the nucleus that an electron can achieve.
(c) a position where an electron probably exists.
(d) a position where an electron cannot exist.

Name _____ Date _____ Class _____

SECTION 2 continued

8. How did de Broglie conclude that electrons have a wave nature?

Scientists knew that any wave confined to a space could have only certain frequencies. De Broglie suggested that electrons should be considered as waves confined to the space around an atomic nucleus; in this way, electron waves could exist only at specific frequencies. According to the relationship $E = h\nu$, these frequencies correspond to the specific quantized energies of the Bohr orbitals.

9. Identify each of the four quantum numbers and the properties to which they refer.

The principal quantum number refers to the main energy level. The angular momentum quantum number refers to the shape of the orbital. The magnetic quantum number refers to the orientation of an orbital around the nucleus. The spin quantum number indicates the spin state of an electron in an orbital.

10. How did the Heisenberg uncertainty principle contribute to the idea that electrons occupy "clouds," or "orbitals"?

The Heisenberg uncertainty principle states that it is impossible to determine simultaneously both the position and velocity of an electron (or any other particle). Because measuring the position of an electron actually changes its position, there is always a basic uncertainty in trying to locate an electron. Thus, the exact position of the electron cannot be found. An electron cloud or orbital represents the region that is the probable location of an electron.

11. Complete the following table:

Principal quantum number, *n*	Number of sublevels	Types of orbitals
1	1	*s*
2	2	*s,p*
3	3	*s,p,d*
4	4	*s,p,d,f*

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

SECTION 3

SHORT ANSWER **Answer the following questions in the space provided.**

1. State the Pauli exclusion principle, and use it to explain why electrons in the same orbital must have opposite spin states.

The Pauli exclusion principle states that no two electrons in an atom may have the same set of four quantum numbers. If both electrons in the same orbital had the same spin state, each electron would have the same four quantum numbers. If one electron has the opposite spin state, the fourth quantum number is different and the exclusion principle is obeyed.

2. Explain the conditions under which the following orbital notation for helium is possible:

$\underline{\uparrow}_{1s}$ $\underline{\uparrow}_{2s}$

This orbital notation is possible if the helium atom is in an excited state.

Write the ground-state electron configuration and orbital notation for each of the following atoms:

3. Phosphorus

$1s^2 2s^2 2p^6 3s^2 3p^3$; $\underline{\uparrow\downarrow}_{1s}$ $\underline{\uparrow\downarrow}_{2s}$ $\underline{\uparrow\downarrow}_{2p_x}$ $\underline{\uparrow\downarrow}_{2p_y}$ $\underline{\uparrow\downarrow}_{2p_z}$ $\underline{\uparrow\downarrow}_{3s}$ $\underline{\uparrow}_{3p_x}$ $\underline{\uparrow}_{3p_y}$ $\underline{\uparrow}_{3p_z}$

4. Nitrogen

$1s^2 2s^2 2p^3$; $\underline{\uparrow\downarrow}_{1s}$ $\underline{\uparrow\downarrow}_{2s}$ $\underline{\uparrow}_{2p_x}$ $\underline{\uparrow}_{2p_y}$ $\underline{\uparrow}_{2p_z}$

5. Potassium

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$; $\underline{\uparrow\downarrow}_{1s}$ $\underline{\uparrow\downarrow}_{2s}$ $\underline{\uparrow\downarrow}_{2p_x}$ $\underline{\uparrow\downarrow}_{2p_y}$ $\underline{\uparrow\downarrow}_{2p_z}$ $\underline{\uparrow\downarrow}_{3s}$ $\underline{\uparrow\downarrow}_{3p_x}$ $\underline{\uparrow\downarrow}_{3p_y}$ $\underline{\uparrow\downarrow}_{3p_z}$ $\underline{\uparrow}_{4s}$

Name _____ Date _____ Class _____

SECTION 3 continued

6. Aluminum

$1s^2 2s^2 2p^6 3s^2 3p^1$; $\underset{1s}{\uparrow\downarrow}$ $\underset{2s}{\uparrow\downarrow}$ $\underset{2p_x}{\uparrow\downarrow}$ $\underset{2p_y}{\uparrow\downarrow}$ $\underset{2p_z}{\uparrow\downarrow}$ $\underset{3s}{\uparrow\downarrow}$ $\underset{3p_x}{\uparrow}$ $\underset{3p_y}{\quad}$ $\underset{3p_z}{\quad}$

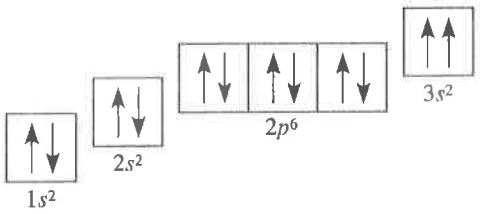
7. Argon

$1s^2 2s^2 2p^6 3s^2 3p^6$; $\underset{1s}{\uparrow\downarrow}$ $\underset{2s}{\uparrow\downarrow}$ $\underset{2p_x}{\uparrow\downarrow}$ $\underset{2p_y}{\uparrow\downarrow}$ $\underset{2p_z}{\uparrow\downarrow}$ $\underset{3s}{\uparrow\downarrow}$ $\underset{3p_x}{\uparrow\downarrow}$ $\underset{3p_y}{\uparrow\downarrow}$ $\underset{3p_z}{\uparrow\downarrow}$

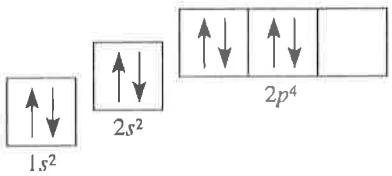
8. Boron

$1s^2 2s^2 2p^1$; $\underset{1s}{\uparrow\downarrow}$ $\underset{2s}{\uparrow\downarrow}$ $\underset{2p_x}{\uparrow}$ $\underset{2p_y}{\quad}$ $\underset{2p_z}{\quad}$

9. Which guideline, Hund's rule or the Pauli exclusion principle, is violated in the following orbital diagrams?



a. Pauli exclusion principle



b. Hund's rule

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW

Arrangement of Electrons in Atoms

MIXED REVIEW

SHORT ANSWER Answer the following questions in the space provided.

1. Under what conditions is a photon emitted from an atom?

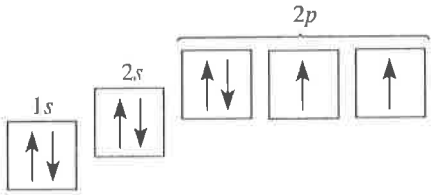
A photon is emitted when an electron moves from a higher energy level to a lower energy level.

2. What do quantum numbers describe?

Quantum numbers describe the energy level, orbital shape, orbital orientation, and spin state of an electron.

3. What is the relationship between the principal quantum number and the electron configuration?

The principle quantum number, *n*, describes the energy level. For example, the electrons at $2p^6$ are at the energy level represented by $n = 2$.



4. In what way does the figure above illustrate Hund's rule?

The most stable arrangement of electrons is one with the maximum number of unpaired electrons.

5. In what way does the figure above illustrate the Pauli exclusion principle?

No two electrons have the same set of four quantum numbers.

Name _____ Date _____ Class _____

MIXED REVIEW continued

6. Elements of the fourth and higher main-energy levels do not seem to follow the normal sequence for filling orbitals. Why is this so?

Electrons from the *s* orbital will sometimes be promoted to a higher energy level in order to form an electron configuration of lowest energy, which is therefore the most stable.

7. How do electrons create the colors in a line-emission spectrum?

The colors are created by photons containing specific amounts of energy, released when an electron makes the transition from a higher energy level to a lower energy level.

8. Write the ground-state electron configuration of the following atoms:

a. Carbon

$1s^2 2s^2 2p^2$

b. Potassium

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

c. Gallium

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$

d. Copper

$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$

PROBLEMS Write the answer on the line to the left. Show all your work in the space provided.

9. 1×10^{12} m What is the wavelength of light that has a frequency of 3×10^{-4} Hz in a vacuum?

10. 3.3×10^{-19} J What is the energy of a photon that has a frequency of 5.0×10^{14} Hz?