Name ______________________ Date ____________ Class ____________

CHAPTER 5 REVIEW

The Periodic Law

SECTION 1

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

1. __c__ In the modern periodic table, elements are ordered

(a) according to decreasing atomic mass.
(b) according to Mendeleev's original design.
(c) according to increasing atomic number.
(d) based on when they were discovered.

2. __d__ Mendeleev noticed that certain similarities in the chemical properties of elements appeared at regular intervals when the elements were arranged in order of increasing

(a) density.
(b) reactivity.
(c) atomic number.
(d) atomic mass.

3. __b__ The modern periodic law states that

(a) no two electrons with the same spin can be found in the same place in an atom.
(b) the physical and chemical properties of an element are functions of its atomic number.
(c) electrons exhibit properties of both particles and waves.
(d) the chemical properties of elements can be grouped according to periodicity, but physical properties cannot.

4. __c__ The discovery of the noble gases changed Mendeleev's periodic table by adding a new

(a) period.
(b) series.
(c) group.
(d) level.

5. __d__ The most distinctive property of the noble gases is that they are

(a) metallic.
(b) radioactive.
(c) metalloid.
(d) largely unreactive.

6. __c__ Lithium, the first element in Group 1, has an atomic number of 3. The second element in this group has an atomic number of

(a) 4.
(b) 10.
(c) 11.
(d) 18.

7. An isotope of fluorine has a mass number of 19 and an atomic number of 9.

__9__ **a.** How many protons are in this atom?

__10__ **b.** How many neutrons are in this atom?

__$^{19}_{9}F$__ **c.** What is the nuclear symbol of this fluorine atom, including its mass number and atomic number?

SECTION 1 continued

8. Samarium, Sm, is a member of the lanthanide series.

Pu, plutonium **a.** Identify the element just below samarium in the periodic table.

32 units **b.** By how many units do the atomic numbers of these two elements differ?

9. A certain isotope contains 53 protons, 78 neutrons, and 54 electrons.

53 **a.** What is its atomic number?

131 **b.** What is the mass number of this atom?

Iodine, I **c.** What is the name of this element?

F, Cl, Br, At **d.** Identify two other elements that are in the same group as this element.

10. In a modern periodic table, every element is a member of both a horizontal row and a vertical column. Which one is the group, and which one is the period?

The group is the vertical column, and the period is the horizontal row.

11. Explain the distinction between atomic mass and atomic number of an element.

The atomic number is the number of protons in an atom. The atomic mass is a weighted average of the masses of the naturally occurring isotopes of that element.

12. In the periodic table, the atomic number of I is greater than that of Te, but its atomic mass is less. This phenomenon also occurs with other neighboring elements in the periodic table. Name two of these pairs of elements. Refer to the periodic table if necessary.

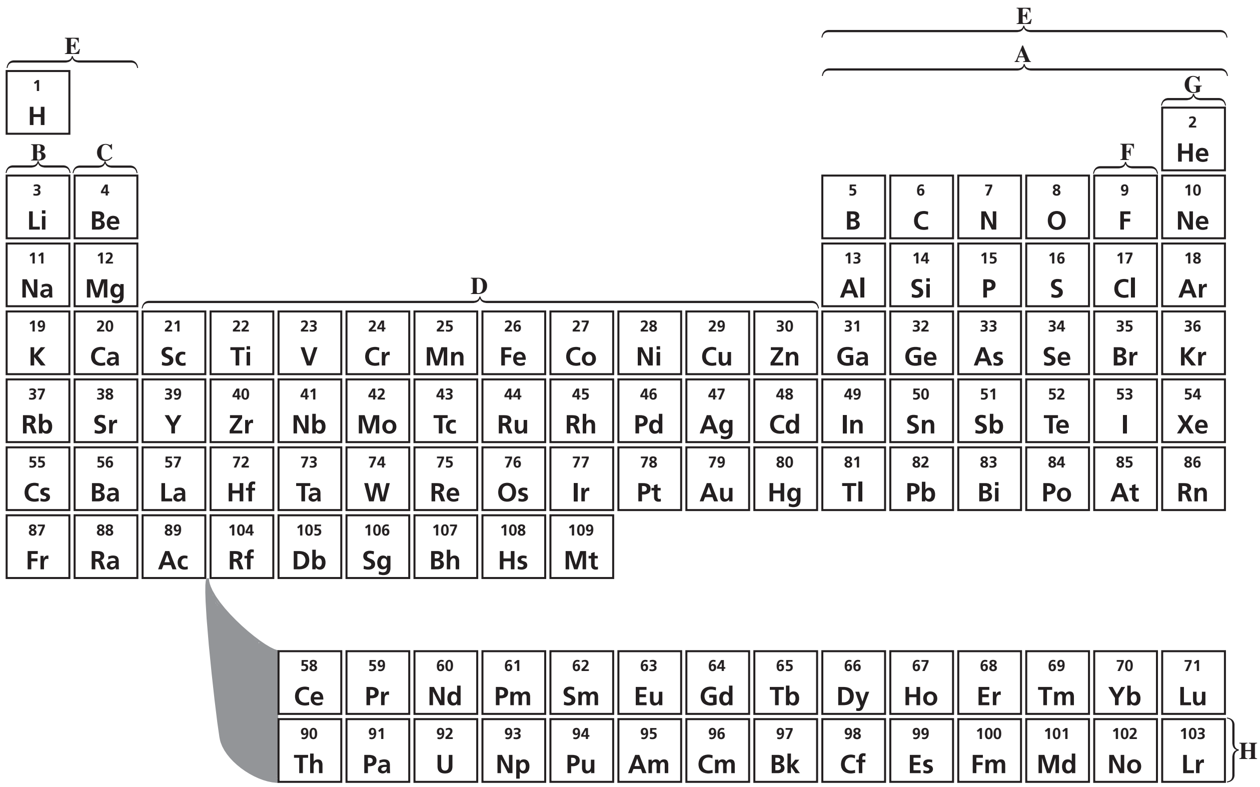
Co and Ni; Ar and K; Th and Pa; U and Np; Pu and Am; Sg and Bh. (The phenomenon occurs here because the mass of only the most stable isotope of each element is given.)

CHAPTER 5 REVIEW

The Periodic Law

SECTION 2

SHORT ANSWER **Use this periodic table to answer the following questions in the space provided.**



1. Identify the element and write the noble-gas notation for each of the following:

a. the Group 14 element in Period 4

Ge; [Ar]3d^{10}4$s^2$4p^2

b. the only metal in Group 15

Bi; [Xe]4f^{14}5d^{10}6$s^2$6p^3

c. the transition metal with the smallest atomic mass

Sc; [Ar]3$d^1$4s^2

d. the alkaline-earth metal with the largest atomic number

Ra; [Rn]7s^2

2. On the periodic table given, several areas are labeled with letters A–H.

_p_ block **a.** Which block does **A** represent, *s, p, d*, or *f*?

b. Identify the remaining labeled areas of the table, choosing from the following terms: *main-group elements, transition elements, lanthanides, actinides, alkali metals, alkaline-earth metals, halogens, noble gases*.

alkali metals **B**

alkaline-earth metals **C**

transition elements **D**

main-group elements (also in B and C) **E**

halogens **F**

noble gases **G**

actinides **H**

3. Give the symbol, period, group, and block for the following:

a. sulfur

S, Period 3, Group 16, _p_ block

b. nickel

Ni, Period 4, Group 10, _d_ block

c. [Kr]$5s^1$

Rb, Period 5, Group 1, _s_ block

d. [Ar]$3d^5 4s^1$

Cr, Period 4, Group 6, _d_ block

4. There are 18 columns in the periodic table; each has a group number. Give the group numbers that make up each of the following blocks:

1–2 **a.** *s* block

13–18 **b.** *p* block

3–12 **c.** *d* block

Name ______________________ Date ____________ Class ____________

CHAPTER 5 REVIEW

The Periodic Law

SECTION 3

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

1. c When an electron is added to a neutral atom, energy is

(a) always absorbed. **(c)** either absorbed or released.
(b) always released. **(d)** neither absorbed nor released.

2. d The energy required to remove an electron from a neutral atom is the atom's

(a) electron affinity. **(c)** electronegativity.
(b) electron energy. **(d)** neither absorbed nor released.

3. From left to right across a period on the periodic table,

negative **a.** electron affinity values tend to become more (negative or positive).

increase **b.** ionization energy values tend to (increase or decrease).

smaller **c.** atomic radii tend to become (larger or smaller).

4. At **a.** Name the halogen with the least-negative electron affinity.

Li **b.** Name the alkali metal with the highest ionization energy.

Ar **c.** Name the element in Period 3 with the smallest atomic radius.

C **d.** Name the Group 14 element with the largest electronegativity.

5. Write the electron configuration of the following:

a. Na

$1s^22s^22p^63s^1$

b. Na^+

$1s^22s^22p^6$

c. O

$1s^22s^22p^4$

d. O^{2-}

$1s^22s^22p^6$

e. Co^{2+}

$1s^22s^22p^63s^23p^63d^7$

SECTION 3 continued

6. **a.** Compare the radius of a positive ion to the radius of its neutral atom.

The radius of a positive ion is smaller than the radius of its corresponding neutral atom.

b. Compare the radius of a negative ion to the radius of its neutral atom.

The radius of a negative ion is larger than the radius of its corresponding neutral atom.

7. **a.** Give the approximate positions and blocks where metals and nonmetals are found in the periodic table.

Metals are on the left side of the periodic table, mostly in the *s*, *d*, and *f* blocks. Nonmetals are on the right side of the periodic table, all in the *p* block (except for hydrogen).

b. Of metals and nonmetals, which tend to form positive ions? Which tend to form negative ions?

Metals tend to form positive ions; nonmetals tend to form negative ions.

8. **Table 3** on page 155 of the text lists successive ionization energies for several elements.

$3s^2$ **a.** Identify the electron that is removed in the first ionization energy of Mg.

$3s^1$ **b.** Identify the electron that is removed in the second ionization energy of Mg.

$2p^6$ **c.** Identify the electron that is removed in the third ionization energy of Mg.

d. Explain why the second ionization energy is higher than the first, the third is higher than the second, and so on.

As electrons are removed in successive ionizations, fewer electrons remain within the atom to shield the attractive force of the nucleus. Each electron removed from an ion experiences a stronger effective nuclear pull than the electron removed before it.

9. Explain the role of valence electrons in the formation of chemical compounds.

Valence electrons are the electrons most subject to the influence of nearby atoms or ions. They are the electrons available to be lost, gained, or shared in the formation of chemical compounds.

Name ______________________ Date ______________ Class ______________

CHAPTER 5 REVIEW

The Periodic Law

MIXED REVIEW

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

1. Consider the neutral atom with 53 protons and 74 neutrons to answer the following questions.

53 **a.** What is its atomic number?

127 **b.** What is its mass number?

atomic number **c.** Is the element's position in a modern periodic table determined by its atomic number or by its atomic mass?

2. Consider an element whose outermost electron configuration is $3d^{10}4s^24p^x$.

Period 4 **a.** To which period does the element belong?

5 **b.** If it is a halogen, what is the value of x?

True **c.** The group number will equal $(10 + 2 + x)$. True or False?

3. ***p*** **a.** In which block are metalloids found, *s*, *p*, *d*, or *f*?

d **b.** In which block are the hardest, densest metals found, *s*, *p*, or *d*?

4. **fluorine, F** **a.** Name the most chemically active halogen.

$1s^22s^22p^5$ **b.** Write its electron configuration.

$1s^22s^22p^6$ **for 1− ion** **c.** Write the configuration of the most stable ion this element makes.

5. Refer only to the periodic table at the top of the review of Section 2 to answer the following questions on periodic trends.

In **a.** Which has the larger radius, Al or In?

Ca **b.** Which has the larger radius, Se or Ca?

Ca **c.** Which has a larger radius, Ca or Ca^{2+}?

nonmetals **d.** Which class has greater ionization energies, metals or nonmetals?

Cl **e.** Which has the greater ionization energy, As or Cl?

negative ion **f.** An element with a large negative electron affinity is most likely to form a (positive ion, negative ion, or neutral atom)?

MIXED REVIEW continued

small g. In general, which has a stronger electron attraction, a large atom or a small atom?

O h. Which has greater electronegativity, O or Se?

O i. In the covalent bond between Se and O, to which atom is the electron pair more closely drawn?

6 j. How many valence electrons are there in a neutral atom of Se?

6. **Ca^{+} and Zn^{2+}** Identify all of the following ions that do not have noble-gas stability.
K^{+} S^{2-} Ca^{+} I^{-} Al^{3+} Zn^{2+}

7. Use only the periodic table in the review of Section 2 to give the noble-gas notation of the following:

$[Ar]3d^{10}4s^{2}4p^{5}$ a. Br

$[Ar]3d^{10}4s^{2}4p^{6}$ b. Br^{-}

$[Kr]4d^{10}5s^{2}5p^{1}$ c. the element in Group 13, Period 5

$[Xe]4f^{1}5d^{1}6s^{2}$ d. the lanthanide with the smallest atomic number

8. Use electron configuration and position in the periodic table to describe the chemical properties of calcium and oxygen.

Calcium is a Group 2 alkaline-earth metal with $[Ar]4s^{2}$ configuration. It forms a stable 2+ ion, has relatively low ionization energy, and forms salt-like ionic compounds. Oxygen, with$[He]2s^{2}2p^{4}$ configuration, is a typical Group 16 nonmetal, making a stable 2− ion; it has high electronegativity and ionization energy and quite negative electron affinity.

9. Copper's electron configuration might be predicted to be $3d^{9}4s^{2}$. But in fact, its configuration is $3d^{10}4s^{1}$. The two elements below copper in Group 11 behave similarly. (Confirm this in the periodic table in **Figure 6** on pages 140–141 of the text.)

$3d^{10}4s^{1}$ a. Which configuration for copper is apparently more stable?

Yes b. Is the *d* sublevel completed in the atoms of these three elements?

True c. Every element in Period 4 has four levels of electrons established. True or False?