8. ELECTRON CONFIGURATIONS AND PERIODICITY

■ Solutions to Exercises

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

8.1 a. Possible orbital diagram.

   b. Possible orbital diagram.

   c. Impossible orbital diagram; there are two electrons in a 2*p* orbital with the same spin.

   d. Possible electron configuration.

   e. Impossible electron configuration; only two electrons are allowed in an *s* subshell.

   f. Impossible electron configuration; only six electrons are allowed in a *p* subshell.

8.2 Look at the periodic table. Start with hydrogen and go through the periods, writing down the subshells being filled, stopping with manganese (Z = 25). You obtain the following order:

Order:	1*s*	2*s*2*p*	3*s*3*p*	4*s*3*d*4*p*
Period:	first	second	third	fourth

   Now fill the subshells with electrons, remembering that you have a total of twenty five electrons to distribute. You obtain

   $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^5$,   or $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$
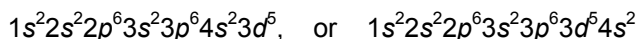
8.3 Arsenic is a main group element in Period 4, Group VA, of the periodic table. The five outer electrons should occupy the $4s$ and $4p$ subshells; the five valence electrons have the configuration $4s^2 4p^3$.

8.4 Because the sum of the $6s^2$ and $6p^2$ electrons gives four outer (valence) electrons, lead should be in Group IVA, which it is. Looking at the table, you find lead in Period 6. From its position, it would be classified as a main-group element.

8.5 The electronic configuration of phosphorus is $1s^2 2s^2 2p^6 3s^2 3p^3$. The orbital diagram is:

8.6 The radius tends to decrease across a row of the periodic table from left to right, and it tends to increase from the top of a column to the bottom. Therefore, in order of increasing radius,

Be < Mg < Na

8.7 It is more likely that (a), 1000 kJ/mol, is the ionization energy for iodine because ionization energies tend to decrease with atomic number in a group (I is below Cl in Group VIIA).

8.8 Fluorine should have a more negative electron affinity because (1) carbon has only two electrons in the p subshell, (2) the -1 fluoride ion has a stable noble-gas configuration, and (3) the electron can approach the fluorine nucleus more closely than the carbon nucleus. This follows the general trend, which is toward more negative electron affinities from left to right in any period.

■ Answers to Concept Checks

8.1 The second-period elements are those in which the $2s$ and $2p$ orbitals fill. Each orbital can hold only one electron, so all four orbitals will be filled after four electrons. Therefore, the second period will have four elements.

8.2   The *s* orbital fills in the first two elements of the period (Groups IA and IIA); then the *p* orbital starts to fill (Group IIIA). Thus, the first element is in Group IIA (Mg), and the next element is in Group IIIA (Al).

8.3   From the information given, the element must be in Group IIA. These elements have positive electron affinities and also have large third ionization energies.

8.4   A metalloid is an element near the staircase line in the periodic table (the green elements in the periodic table on the inside front cover of the book). The formula R_2O_5 suggests a period VA element. There are two metalloids in Group VA, arsenic and antimony. That this is an acidic oxide indicates this metalloid has considerable nonmetal character. So, of the two metalloids, the one nearer the top of the column, arsenic, seems most likely. This agrees with the text, which notes that arsenic(V) oxide is acidic whereas antimony(V) oxide is amphoteric.

■ Answers to Review Questions

8.1   In the original Stern-Gerlach experiment, a beam of silver atoms is directed into the field of a specially designed magnet. (The same can be done with hydrogen atoms.) The beam of atoms is split into two by the magnetic field; half are bent toward one magnetic pole face and the other half toward the other magnetic pole face. This effect shows that the atoms themselves act as magnets with a positive or a negative component as indicated by the positive or negative spin quantum numbers.

8.2   In effect, the electron acts as though it were a sphere of spinning charge (Figure 8.3). Like any circulating electric charge, it creates a magnetic field with a spin axis that has more than one possible direction relative to a magnetic field. Electron spin is subject to a quantum restriction to one of two directions corresponding to the m_s quantum numbers +1/2 and -1/2.

8.3   The Pauli exclusion principle limits the configurations of an atom by excluding configurations in which two or more electrons have the same four quantum numbers. For example, each electron in the same orbital must have different m_s values. This also implies that only two electrons occupy one orbital.

8.4   According to the principles discussed in Section 7.5, the number of orbitals in the g subshell (l = 4) is given by 2l + 1 and is thus equal to nine. Because each orbital can hold a maximum of two electrons, the *g* subshell can hold a maximum of eighteen electrons.

8.5   The orbitals, in order of increasing energy up to and including the 3*p* orbitals (but not including the 3*d* orbitals), are as follows: 1*s*, 2*s*, 2*p*, 3*s*, and 3*p* (Figure 8.7).

8.6   The noble-gas core is an inner-shell configuration corresponding to one of the noble gases. The pseudo-noble-gas core is an inner-shell configuration corresponding to one of the noble gases together with $(n-1)d^{10}$ electrons. Like the noble-gas core electrons, the d^{10} electrons are not involved in chemical reactions. The valence electron is an electron (of an atom) located outside the noble-gas core or pseudo-noble-gas core. It is an electron primarily involved in chemical reactions.

8.7   The orbital diagram for the $1s^2 2s^2 2p^4$ ground state of oxygen is
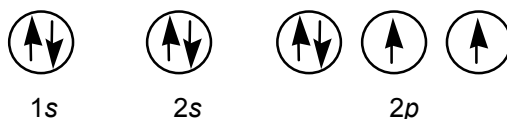
Another possible oxygen orbital diagram, but not a ground state, is
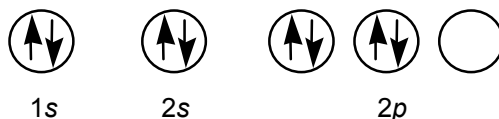
8.8   A diamagnetic substance is a substance that is not attracted by a magnetic field or is very slightly repelled by such a field. This property generally indicates the substance has only paired electrons. A paramagnetic substance is a substance that is weakly attracted by a magnetic field. This property generally indicates the substance has one or more unpaired electrons. Ground-state oxygen has two unpaired $2p$ electrons and is therefore paramagnetic.

8.9   In Groups IA and IIA, the outer s subshell is being filled: s^1 for Group IA and s^2 for Group IIA. In Groups IIIA to VIIIA, the outer p subshell is being filled: p^1 for IIIA, p^2 for IVA, p^3 for VA, p^4 for VIA, p^5 for VIIA, and p^6 for VIIIA. In the transition elements, the $(n-1)d$ subshell is being filled from d^1 to d^{10} electrons. In the lanthanides and actinides, the f subshell is being filled from f^1 to f^{14} electrons.

8.10   Mendeleev arranged the elements in increasing order of atomic weight, an arrangement that was later changed to atomic numbers. His periodic table was divided into rows (periods) and columns (groups). In his first attempt, he left spaces for what he believed to be undiscovered elements. In row five, under aluminum and above indium in Group III, he left a blank space. This Group III element he called eka-aluminum, and he predicted its properties from those of aluminum and indium. Later, the French chemist de Boisbaudran discovered this element and named it gallium.

8.11   In a plot of atomic radii versus atomic number (Figure 8.16), the major trends that emerge are the following: (1) Within each period (horizontal row), the atomic radius tends to decrease with increasing atomic number or nuclear charge. The largest atom in a period is thus the Group 1A atom, and the smallest atom in a period is thus the noble-gas atom. (2) Within each group (vertical column), the atomic radius tends to increase with the period number.

        In a plot of ionization energy versus atomic number (Figure 8.18), the major trends are (1) the ionization energy within a period increases with atomic number, and (2) the ionization energy within a group tends to decrease going down the group.

8.12   The alkaline earth element with the smallest radius is beryllium (Be).

8.13   Group VIIA (halogens) is the main group with the most negative electron affinities. Configurations with filled subshells (ground states of the noble-gas elements) would form unstable negative ions when adding one electron per atom.

8.14   The Na^+ and Mg^{2+} ions are stable because they are isoelectronic with the noble gas neon. If Na^{2+} and Mg^{3+} ions were to exist, they would be very unstable because they would not be isoelectronic with any noble-gas structure and because of the energy needed to remove an electron from an inner shell.

8.15   The elements tend to increase in metallic character from right to left in any period. They also tend to increase in metallic character down any column (group) of elements.

8.16   A basic oxide is an oxide that reacts with acids. An example is calcium oxide, CaO. An acidic oxide is an oxide that reacts with bases. An example is carbon dioxide, CO_2.

8.17   Rubidium is the alkali metal atom with a $5s^1$ configuration.

8.18   Atomic number equals 117 (protons in last known element plus those needed to reach Group VIIA).

8.19   The following elements are in Groups IIIA to VIA:

| Group IIIA	Group IVA	Group VA	Group VIA
B: metalloid	C: nonmetal	N: nonmetal	O: nonmetal
Al: metal	Si: metalloid	P: nonmetal	S: nonmetal
Ga: metal	Ge: metalloid	As: metalloid	Se: nonmetal
In: metal	Sn: metal	Sb: metalloid	Te: metalloid
Tl: metal	Pb: metal	Bi: metal	Po: metal

        Yes, each column displays the expected increasing metallic character.

8.20   The oxides of the following elements are listed as either acidic, basic, amphoteric, or
        t.g.n.i. (text gives no information):

Group IIIA	Group IVA	Group VA	Group VIA
B: acidic	C: acidic	N: acidic	O: amphoteric (H_2O)
Al: amphoteric	Si: acidic	P: acidic	S: acidic
Ga: amphoteric	Ge: t.g.n.i.	As: t.g.n.i.	Se: acidic
In: basic	Sn: amphoteric	Sb: t.g.n.i.	Te: t.g.n.i.
Tl: basic	Pb: amphoteric	Bi: t.g.n.i.	Po: t.g.n.i.

8.21 $2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$
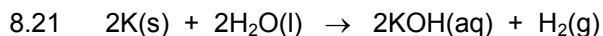
8.22   Barium should be a soft, reactive metal. Barium should form the basic oxide, BaO.
        Barium metal, for example, would be expected to react with water according to the
        equation

        $Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$
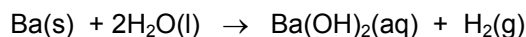
8.23   The two oxides of carbon are carbon monoxide, CO, and carbon dioxide, CO_2.

8.24 a. White phosphorus   b. Sulfur   c. Bromine   d. Sodium

■ Answers to Conceptual Problems

8.25   This statement of the Pauli principle implies there can be two electrons with the same
        spin in a given orbital. Because an electron can have either one of two spins, any
        orbital can hold a maximum of four electrons. The first six elements of the periodic
        table would have the following electronic configurations:

        (1) $1s^1$
        (2) $1s^2$
        (3) $1s^3$
        (4) $1s^4$
        (5) $1s^4 2s^1$
        (6) $1s^4 2s^2$

8.26   The first period of the periodic table would have the following allowed quantum
        numbers: $n = 1$; $l = 1$; $m_l = 0, +1, -1$; $m_s = +1/2, -1/2$. There are six different possible
        combinations. Therefore, there would be six elements in the first period.

8.27   The elements are in Group IIA (only s electrons) and IIIB (d electrons). They are also
        in Period 5. Therefore, the elements are strontium (Sr) and yttrium (Y).

8.28   The elements are in Group IIA. They must also be in Periods 4 (no d electrons) and 5 (d electrons). Therefore, the elements are calcium (Ca) and strontium (Sr).

8.29   Keeping in mind that a filled orbital is usually a stable configuration for an atom, an element in this universe with five electrons would probably lose the two electrons in the second orbital and form a cation with a charge of positive two. The other possible option is for the atom to gain seven additional electrons to fill the second orbital. However, this is unlikely given that the nuclear charge would be relatively small, and electron-electron repulsions in such an atom would be large.

8.30   Keep in mind that the ionization energy of an atom provides a measure of how strongly an electron is attracted to that atom. The electron affinity of an atom provides a measure of how strongly attracted an additional electron is to the atom. Both electron affinity and ionization energy provide information about the strength of the attraction between electrons and a particular nucleus. An element that forms an anion easily has an electron affinity much less than zero and a very large first ionization energy. Examples are the elements on the upper right of the periodic table, like fluorine, with an electron affinity of -328 kJ/mol and a first ionization energy of 1681 kJ/mol.

8.31   The elements that form oxides of the form RO_2 are in Groups IVA and VIA. However, there are no metals in Group VIA. Therefore, the elements are in Group IVA. The metalloid is germanium (Ge), and the metal is tin (Sn). GeO_2 is the acidic oxide, and SnO_2 is the amphoteric oxide.

8.32   Oxides of Groups IIIA and VA have oxides of the form R_2O_3. However, Group VA oxides can also be of the form R_2O_5, so the element is in group IIIA. It is also acidic. Therefore, it must be boron oxide, B_2O_3.

8.33 a. Elements of Group IIA and Group VIIIA do not form negative ions, and thus have electron affinities that are greater than zero.

   b. The large difference between the second and third ionization energies means this would be a Group IIA element.

   c. Luster and conductivity are properties of metals. Group IIA is the only main group that contains just metallic elements.

8.34 a. These elements would form +2 ions and have oxides with formula AO.

   b. $2\,A\ +\ O_2\ \rightarrow\ 2\,AO$

   c. The oxides with formula AO would be basic. The reaction with water would be

$$AO\ +\ H_2O\ \rightarrow\ A(OH)_2$$

# ■ Solutions to Practice Problems

Note on significant figures: If the final answer to a solution needs to be rounded off, it is given first with one nonsignificant figure, and the last significant figure is underlined. The final answer is then rounded to the correct number of significant figures. In multiple-step problems, intermediate answers are given with at least one nonsignificant figure; however, only the final answer has been rounded off.

8.35 a. Not allowed; the paired electrons in the 2*p* orbital should have opposite spins.

b. Allowed; electron configuration is $1s^22s^22p^4$.

c. Not allowed; the electrons in the 1*s* orbital must have opposite spins.

d. Not allowed; the 2*s* orbital can hold only two electrons maximum, with opposite spins.

8.36 a. Allowed; $1s^22s^12p^3$.

b. Not allowed; the two electrons in the 1*s* orbital must have opposite spins.

c. Not allowed; the 2*s* orbital can hold only two electrons maximum, with opposite spins.

d. Not allowed; the unpaired electrons in the 2*p* orbitals should have parallel spins.

8.37 a. Impossible state; the 2*p* orbitals can hold no more than six electrons.

b. Impossible state; the 3*s* orbital can hold no more than two electrons.

c. Possible state.

d. Possible state; however, the 3*p* and 4*s* orbitals should be filled before the 3*d* orbital.

8.38 a. Possible state; however, the 2*s* orbital should be filled with two electrons before the 2*p* orbitals are filled completely.

b. Impossible; the 2*p* orbitals can hold no more than six electrons.

c. Possible state; however, the 4*s* orbital should be filled before the 3*d* is filled.

d. Possible state; however, the 3*s* and 3*p* orbitals should be filled before the 3*d* orbital.

8.39 The six possible orbital diagrams for $1s^2 2p^1$ are
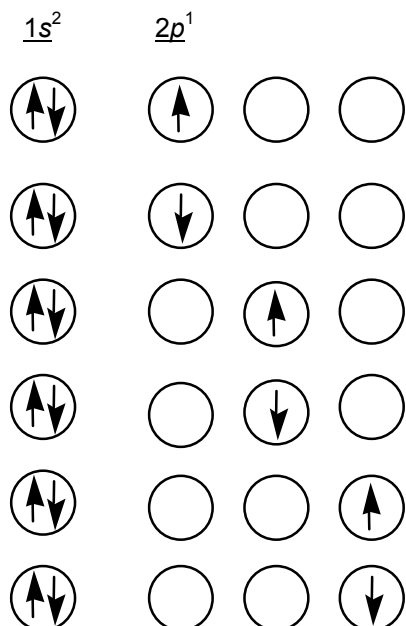
$\underline{1s^2}$ $\underline{2p^1}$

8.40 The twelve possible orbital diagrams for $1s^1 2p^1$ are

$\underline{1s^1}$ $\underline{2p^1}$ $\underline{1s^1}$ $\underline{2p^1}$ $\underline{1s^1}$ $\underline{2p^1}$
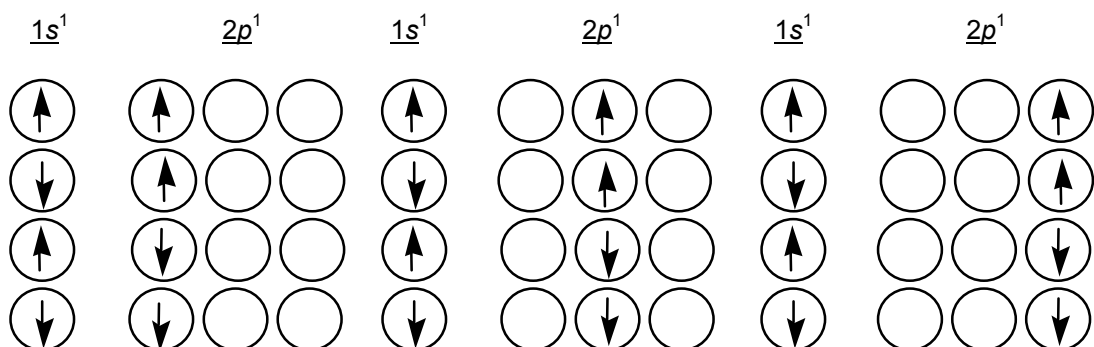
8.41 Iodine (Z = 53): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^2 5p^5$

8.42 Arsenic (Z = 33): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$

8.43 Manganese (Z = 25): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$

8.44 Nickel (Z = 28): $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

8.45   Bromine (Z = 35): $4s^2 4p^5$

8.46   Bismuth (Z = 83): $6s^2 6p^3$

8.47   Zirconium (Z = 40): $4d^2 5s^2$

8.48   Manganese (Z = 25): $3d^5 4s^2$

8.49   The highest value of n is six, so thallium (Tl) is in the sixth period. The $5d$ subshell is filled, and there is a $6p$ electron, so Tl belongs in an A group. There are three valence electrons, so Tl is in Group IIIA. It is a main-group element.

8.50   The highest value of n is six, so iridium (Ir) is in Period 6. The $5d$ subshell is not completely filled, so Ir belongs to a B group. There are nine valence electrons, so Ir is in Group VIIIB. Iridium is a d-transition element.
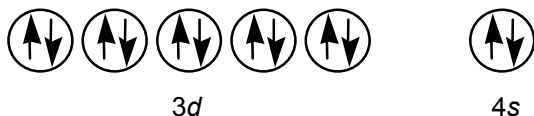
8.51   Cobalt (Z = 27): [Ar]

3d   4s
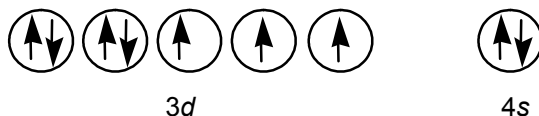
8.52   Samarium (Z = 62): [Xe]

4f   6s
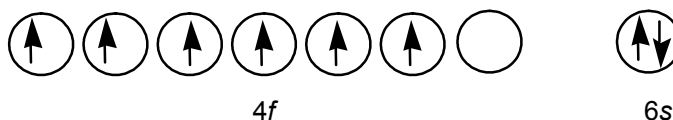
8.53   Potassium (Z = 19): [Ar]

4s

All the subshells are filled in the argon core; however, the 4s electron is unpaired, causing the ground state of the potassium atom to be a paramagnetic substance.

8.54   Zinc (Z = 30): [Ar]

3d   4s

All the subshells through the 4s subshell are filled, so the ground state of a zinc atom is a diamagnetic substance.

8.55   Atomic radius increases going down a column (group), from S to Se, and increases going from right to left in a row, from Se to As. Thus, the order by increasing atomic radius is S, Se, As.

8.56   Atomic radius increases going down a column (group), from F to Cl, and increases going from right to left in a row, from S to Cl. Thus, the order by increasing atomic radius is F, Cl, S.

8.57   Ionization energy increases going left to right in a row. Thus, the order by increasing ionization energy is Na, Al, Cl, Ar.

8.58   Ionization energy increases going up a column (group), from Ca to Mg, and increases going left to right in a row, from Mg to S. Thus, the order by increasing ionization energy is Ca, Mg, S.

8.59 a.   In general, the electron affinity becomes more negative going from left to right within a period. Thus, Br has a more negative electron affinity than As.

     b.   In general, a nonmetal has a more negative electron affinity than a metal. Thus, F has a more negative electron affinity than Li.

8.60 a.   In general, the electron affinity becomes more negative going from left to right within a period. Thus, Cl has a more negative electron affinity than S.

     b.   In general, the electron affinity of a nonmetal is more negative than that of a metal. Thus, Se has a more negative electron affinity than K.

8.61   Chlorine forms the ClO_3^- ion, so bromine should form the BrO_3^- ion, and potassium forms the K^+ ion, so lithium should be Li^+. Thus, the expected formula of lithium bromate is $LiBrO_3$.

8.62   The expected positive oxidation states of selenium are +4 and +6. The corresponding oxides have the simplest formulas of SeO_2 and SeO_3.
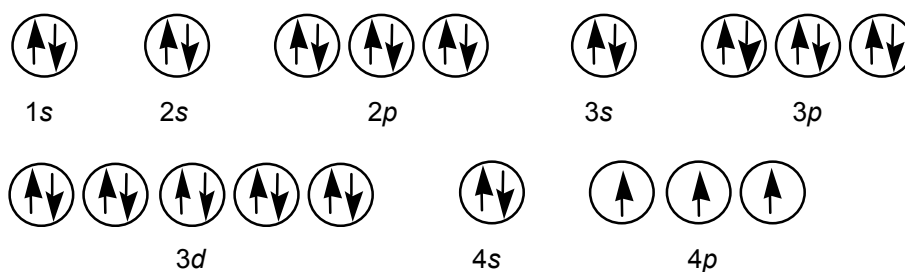
■ Solutions to General Problems

8.63   Strontium: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2$

8.64   Tin: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 5s^2 5p^2$
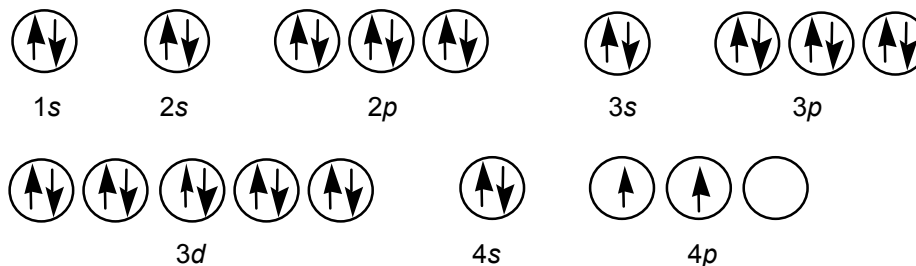
8.65   Polonium: $6s^2 6p^4$

8.66   Thallium: $6s^2 6p^1$

8.67   The orbital diagram for arsenic is:

$$\underset{1s}{\boxed{\uparrow\downarrow}} \quad \underset{2s}{\boxed{\uparrow\downarrow}} \quad \underset{2p}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}} \quad \underset{3s}{\boxed{\uparrow\downarrow}} \quad \underset{3p}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}}$$

$$\underset{3d}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}} \quad \underset{4s}{\boxed{\uparrow\downarrow}} \quad \underset{4p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}}$$

8.68   The orbital diagram for germanium is

$$\underset{1s}{\boxed{\uparrow\downarrow}} \quad \underset{2s}{\boxed{\uparrow\downarrow}} \quad \underset{2p}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}} \quad \underset{3s}{\boxed{\uparrow\downarrow}} \quad \underset{3p}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}}$$

$$\underset{3d}{\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}\boxed{\uparrow\downarrow}} \quad \underset{4s}{\boxed{\uparrow\downarrow}} \quad \underset{4p}{\boxed{\uparrow}\boxed{\uparrow}\boxed{\ }}$$

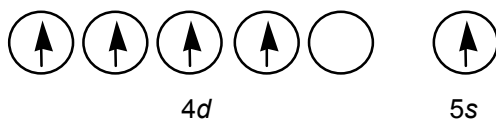8.69   For eka-lead: [Rn] $5f^{14} 6d^{10} 7s^2 7p^2$. It is a metal; the oxide is eka-PbO or eka-PbO_2.

8.70   For eka-bismuth: [Rn] $5f^{14} 6d^{10} 7s^2 7p^3$. It is a metal; the oxide is eka-Bi_2O_3 or eka-Bi_2O_5.
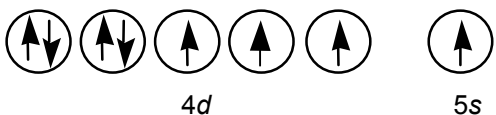
8.71   The ionization energy of Fr is ~370 kJ/mol (slightly less than that of Cs).

8.72   The ionization energy of At is <1000 kJ/mol (slightly less than that of I).

8.73   Niobium: [Kr]

4d   5s

8.74   Ruthenium: [Kr]

4d   5s

8.75 a. Cl_2 b. Na   c. Sb   d. Ar

8.76 a. F_2 b . Ba   c. Ga   d. O_2

8.77   Element with Z = 23: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$. The element is in Group VB (three of the
        five valence electrons are d electrons) and in Period 4 (largest n is 4). It is a d-block
        transition element.

8.78   Element with Z = 33: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$. The element is in Group VA (five
        valence electrons, none of them d electrons) and in Period 4 (largest n is 4). It is a
        main-group element.

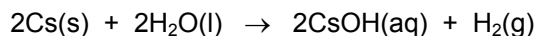# ■ Solutions to Cumulative-Skills Problems

8.79   The equation is:

$$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$$
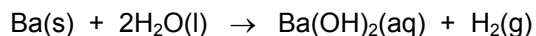
Using the equation, calculate the moles of H_2; then use the ideal gas law to convert to volume.

$$\text{mol } H_2 = 2.50 \text{ g Ba} \times \frac{1 \text{ mol Ba}}{137.33 \text{ g Ba}} \times \frac{1 \text{ mol } H_2}{1 \text{ mol Ba}} = 0.018\underline{2}04 \text{ mol}$$

$$V = \frac{nRT}{P} = \frac{(0.018204 \text{ mol})(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(294.2 \text{ K})}{(748/760) \text{ atm}}$$

$$= 0.44\underline{6}51 \text{ L} \quad (447 \text{ mL})$$

8.80   The equation is:

$$2Cs(s) + 2H_2O(l) \rightarrow 2CsOH(aq) + H_2(g)$$

Use the ideal gas law to calculate moles of H_2 from 48.1 mL (0.0481 L). Then, use the equation to convert moles of H_2 to moles and mass of Cs.

$$n = \frac{PV}{RT} = \frac{\left[(768/760) \text{ atm}\right](0.0481 \text{ L})}{(0.082057 \text{ L} \cdot \text{atm/K} \cdot \text{mol})(292.2 \text{ K})} = 0.00202\underline{7} \text{ mol } H_2$$

$$\text{mol Cs} = 0.002027 \text{ mol } H_2 \times \frac{2 \text{ mol Cs}}{1 \text{ mol } H_2} = 0.004\underline{0}54 \text{ mol Cs}$$

$$\text{mass Cs} = 0.004054 \text{ mol Cs} \times \frac{132.9 \text{ g Cs}}{1 \text{ mol Cs}} = 0.53\underline{8}7 = 0.539 \text{ g Cs}$$

8.81   Radium is in Group IIA; hence, the radium cation is Ra^{2+}, and its oxide is RaO. Use the atomic weights to calculate the percentage of Ra in RaO.

$$\text{Percent Ra} = \frac{226 \text{ amu Ra}}{226 \text{ amu Ra} + 16.00 \text{ amu O}} \times 100\% = 93.\underline{3}8 = 93.4\% \text{ Ra}$$

8.82   Tellurium is in Group VIA; hence, its anion is Te^{2-}, and its hydrogen compound is H_2Te.
        Use the atomic weights to calculate the percentage of Te in H_2Te.

$$Percent\ Te\ =\ \frac{127.60\ amu\ Te}{127.60\ amu\ Te\ +\ (2)(1.00794\ amu\ H)}\ x\ 100\%$$

$$=\ 98.44\underline{4}7\ =\ 98.445\%\ Te$$

8.83   Convert 5.00 mg (0.00500 g) Na to moles of Na; then convert to energy using the first
        ionization energy of 496 kJ/mol Na.

$$mol\ Na\ =\ 0.00500\ g\ Na\ x\ \frac{1\ mol\ Na}{22.99\ g\ Na}\ =\ 2.1\underline{7}4\ x\ 10^{-4}\ mol\ Na$$

$$2.1\underline{7}4\ x\ 10^{-4}\ mol\ Na\ x\ \frac{496\ kJ}{1\ mol\ Na}\ =\ 0.10\underline{7}8\ =\ 0.108\ kJ\ =\ 108\ J$$

8.84   Convert 2.65 mg (0.00265 g) Cl atoms to moles of Cl(g); then convert to energy using
        the electron affinity of 349 kJ/mol Cl(g).

$$mol\ Cl\ =\ 0.00265\ g\ Cl\ x\ \frac{1\ mol\ Cl}{35.45\ g\ Cl}\ =\ 7.4\underline{7}53\ x\ 10^{-5}\ mol\ Cl$$

$$7.4\underline{7}53\ x\ 10^{-5}\ mol\ Cl\ x\ \frac{-349\ kJ}{1\ mol\ Cl}\ =\ -2.6\underline{0}8\ x\ 10^{-2}\ =\ -2.61\ x\ 10^{-2}\ kJ$$

8.85   Use the Bohr formula, where $n_f = \infty$ and $n_i = 1$.

$$\Delta E\ =\ -R_H \left[\frac{1}{\infty^2}\ -\ \frac{1}{1^2} \right]\ =\ -R_H[-1]\ =\ R_H\ =\ \frac{2.179\ x\ 10^{-18}\ J}{1\ H\ atom}$$
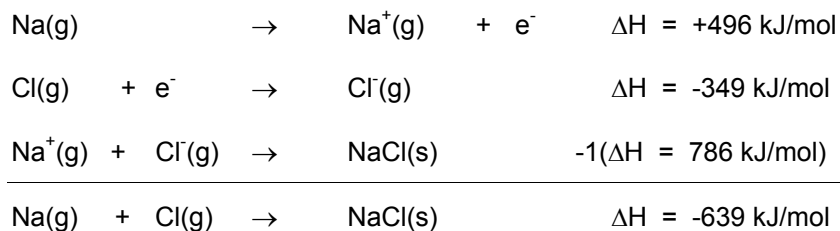
$$I.E.\ =\ \frac{2.179\ x\ 10^{-18}\ J}{1\ H\ atom}\ x\ \frac{6.022\ x\ 10^{23}\ H\ atoms}{1\ mol\ H}\ =\ \frac{1.31\underline{2}19\ x\ 10^6\ J}{1\ mol\ H}$$

$$=\ 1.312\ x\ 10^3\ kJ/mol\ H$$

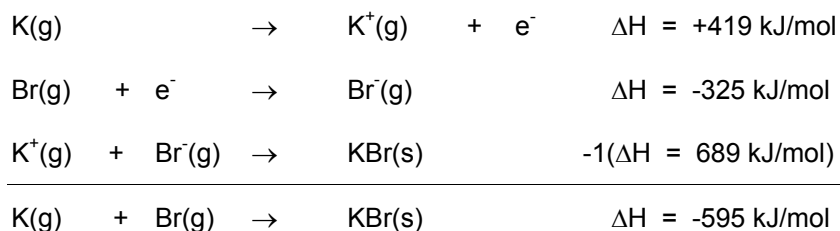8.86   Use the Bohr formula, where $n_f = \infty$, $n_i = 1$, and $Z = 2$.

$$\Delta E = -R_H \left[\frac{2^2}{\infty^2} - \frac{2^2}{1^2} \right] = -R_H[-4] = 4R_H = \frac{4 \times 2.179 \times 10^{-18} \text{ J}}{1 \text{ H atom}}$$

$$\text{I.E.} = \frac{4 \times 2.179 \times 10^{-18} \text{ J}}{1 \text{ He}^+ \text{ ion}} \times \frac{6.022 \times 10^{23} \text{ He}^+ \text{ ions}}{1 \text{ mol He}^+} \times \frac{5.2\underline{4}87 \times 10^6 \text{ J}}{1 \text{ mol He}^+}$$

$$= 5.25 \times 10^3 \text{ kJ/mol He}^+$$

8.87   Add the three equations after reversing the equation for the lattice energy and its ΔH:

Na(g)		$\rightarrow$	Na$^+$(g)	+ e$^-$	ΔH = +496 kJ/mol
Cl(g)	+ e$^-$	$\rightarrow$	Cl$^-$(g)		ΔH = -349 kJ/mol
Na$^+$(g) + Cl$^-$(g)		$\rightarrow$	NaCl(s)		-1(ΔH = 786 kJ/mol)
Na(g) + Cl(g)		$\rightarrow$	NaCl(s)		ΔH = -639 kJ/mol

8.88   Add the three equations after reversing the equation for the lattice energy and its ΔH:

K(g)		$\rightarrow$	K$^+$(g)	+ e$^-$	ΔH = +419 kJ/mol
Br(g)	+ e$^-$	$\rightarrow$	Br$^-$(g)		ΔH = -325 kJ/mol
K$^+$(g) + Br$^-$(g)		$\rightarrow$	KBr(s)		-1(ΔH = 689 kJ/mol)
K(g) + Br(g)		$\rightarrow$	KBr(s)		ΔH = -595 kJ/mol