CHAPTER 8 ELECTRON CONFIGURATION AND CHEMICAL PERIODICITY

END–OF–CHAPTER PROBLEMS

8.1 Elements are listed in the periodic table in an ordered, systematic way that correlates with a periodicity of their chemical and physical properties. The theoretical basis for the table in terms of atomic number and electron configuration does not allow for an "unknown element" between Sn and Sb.

8.2 Today, the elements are listed in order of increasing atomic number. This makes a difference in the sequence of elements in only a few cases, as the larger atomic number usually has the larger atomic mass. One of these exceptions is iodine, $Z = 53$, which is after tellurium, $Z = 52$, even though tellurium has a higher atomic mass.

8.3 Plan: The value should be the average of the elements above and below the one of interest.
Solution:
a) Predicted atomic mass (K) =

$$\frac{\text{Na} + \text{Rb}}{2} = \frac{22.99 + 85.47}{2} = \mathbf{54.23\ amu} \qquad \text{(actual value = 39.10 amu)}$$

b) Predicted melting point (Br_2) =

$$\frac{Cl_2 + I_2}{2} = \frac{-101.0 + 113.6}{2} = \mathbf{6.3°C} \qquad \text{(actual value = –7.2°C)}$$

8.4 The allowed values of n: positive integers: 1, 2, 3, 4,...∞
The allowed values of l: integers from 0 to $n – 1$: 0, 1, 2, ... $n – 1$
The allowed values of m_l: integers from $–l$ to 0 to $+l$: $–l$, $(–l + 1)$, ... 0, ... $(l – 1)$, $+l$
The allowed values of m_s: –1/2 or +1/2

8.5 The quantum number m_s relates to just the electron; all the others describe the orbital.

8.6 The exclusion principle states that no two electrons in the same atom may have the same four quantum numbers. Within a particular orbital, there can be only two electrons and they must have opposing spins.

8.7 In a one-electron system, all sublevels of a particular level (such as 2*s* and 2*p*) have the same energy. In many electron systems, the principal energy levels are split into sublevels of differing energies. This splitting is due to electron-electron repulsions. Be^{3+} would be more like H since both have only one 1*s* electron.

8.8 Shielding occurs when inner electrons protect or shield outer electrons from the full nuclear attractive force. The effective nuclear charge is the nuclear charge an electron actually experiences. As the number of inner electrons increases, shielding increases, and the effective nuclear charge decreases.

8.9 Penetration occurs when the probability distribution of an orbital is large near the nucleus, which results in an increase of the overall attraction of the nucleus for the electron, lowering its energy. Shielding results in lessening this effective nuclear charge on outer shell electrons, since they spend most of their time at distances farther from the nucleus and are shielded from the nuclear charge by the inner electrons. The lower the l quantum number of an orbital, the more time the electron spends penetrating near the nucleus. This results in a lower energy for a 3*p* electron than for a 3*d* electron in the same atom.

8.10 Plan: The integer in front of the letter represents the n value. The l value designates the orbital type: $l = 0 = s$ orbital; $l = 1 = p$ orbital; $l = 2 = d$ orbital; $l = 3 = f$ orbital. Remember that a p orbital set contains 3 orbitals, a d orbital set has 5 orbitals, and an f orbital set has 7 orbitals. Any one orbital can hold a maximum of 2 electrons.

Solution:
a) The l = 1 quantum number can only refer to a *p* orbital. These quantum numbers designate the 2*p* orbital set (n = 2), which hold a maximum of **6** electrons, 2 electrons in each of the three 2*p* orbitals.
b) There are five 3*d* orbitals, therefore a maximum of **10** electrons can have the 3*d* designation, 2 electrons in each of the five 3*d* orbitals.
c) There is one 4*s* orbital which holds a maximum of **2** electrons.

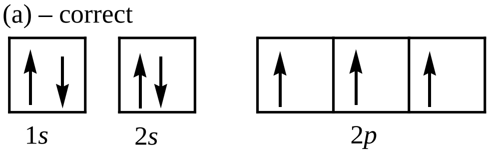
8.11 a) The l = 1 quantum number can only refer to a *p* orbital, and the m_l value of 0 specifies one particular *p* orbital, which holds a maximum of **2** electrons.
b) The 5*p* orbitals, like any *p* orbital set, can hold a maximum of **6** electrons.
c) The l = 3 quantum number can only refer to an *f* orbital. These quantum numbers designate the 4*f* orbitals, which hold a maximum of **14** electrons, 2 electrons in each of the seven 4*f* orbitals.

8.12 Plan: The integer in front of the letter represents the n value. The l value designates the orbital type: $l = 0 = s$ orbital; $l = 1 = p$ orbital; $l = 2 = d$ orbital; $l = 3 = f$ orbital. Remember that a *p* orbital set contains 3 orbitals, a *d* orbital set has 5 orbitals, and an *f* orbital set has 7 orbitals. Any one orbital can hold a maximum of 2 electrons.
Solution:
a) **6** electrons can be found in the three 4*p* orbitals, 2 in each orbital.
b) The l = 1 quantum number can only refer to a *p* orbital, and the m_l value of +1 specifies one particular *p* orbital, which holds a maximum of **2** electrons with the difference between the two electrons being in the m_s quantum number.
c) **14** electrons can be found in the 5*f* orbitals (l = 3 designates *f* orbitals; there are 7*f* orbitals in a set).

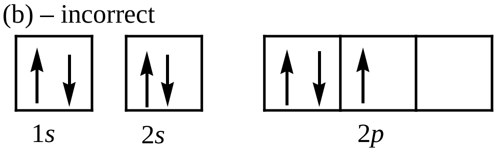
8.13 a) **Two** electrons, at most, can be found in any *s* orbital.
b) The l = 2 quantum number can only refer to a *d* orbital. These quantum numbers designate the 3*d* orbitals, which hold a maximum of **10** electrons, 2 electrons in each of the five 3*d* orbitals.
c) A maximum of **10** electrons can be found in the five *6d* orbitals.

8.14 Properties recur periodically due to similarities in electron configurations recurring periodically.
Na: $1s^22s^22p^6\underline{\mathbf{3s^1}}$
K: $1s^22s^22p^63s^23p^6\underline{\mathbf{4s^1}}$
The properties of Na and K are similar due to a similarity in their outer shell electron configuration; both have one electron in an outer shell *s* orbital.

8.15 Hund's rule states that electrons will fill empty orbitals in the same sublevel before filling half-filled orbitals. This lowest-energy arrangement has the maximum number of unpaired electrons with parallel spins. In the correct electron configuration for nitrogen shown in (a), the 2*p* orbitals each have one unpaired electron; in the incorrect configuration shown in (b), electrons were paired in one of the 2*p* orbitals while leaving one 2*p* orbital empty. The arrows in the 2*p* orbitals of configuration (a) could alternatively all point down.
(a) – correct



(b) – incorrect



8.16 Similarities in chemical behavior are reflected in similarities in the distribution of electrons in the highest energy orbitals. The periodic table may be re-created based on these similar outer electron configurations when orbital filling in is order of increasing energy.

8.17 For elements in the same group (vertical column in periodic table), the electron configuration of the outer electrons are identical except for the n value. For elements in the same period (horizontal row in periodic table), their configurations vary because each succeeding element has one additional electron. The electron configurations are similar only in the fact that the same level (principal quantum number) is the outer level.

8.18 <u>Plan:</u> Write the electron configuration for the atom or ion and find the electron for which you are writing the quantum numbers. Assume that the electron is in the ground-state configuration and that electrons fill in a p_x–p_y–p_z order. By convention, we assign the first electron to fill an orbital with an m_s value of +1/2. Also by convention, $m_l = -1$ for the p_x orbital, $m_l = 0$ for the p_y orbital, and $m_l = +1$ for the p_z orbital. Also, keep in mind the following letter orbital designation for each l value: $l = 0 = s$ orbital, $l = 1 = p$ orbital, $l = 2 = d$ orbital, and $l = 3 = f$ orbital.

<u>Solution:</u>

a) Rb: [Kr]5s^1. The outermost electron in a rubidium atom would be in a 5*s* orbital (rubidium is in Row 5, Group 1). The quantum numbers for this electron are $\boldsymbol{n = 5, l = 0, m_l = 0,}$ and $\boldsymbol{m_s = +1/2}$.

b) The S^- ion would have the configuration [Ne]3$s^2$3p^5. The electron added would go into the 3p_z orbital and is the second electron in that orbital. Quantum numbers are $\boldsymbol{n = 3, l = 1, m_l = +1,}$ and $\boldsymbol{m_s = -1/2}$.

c) Ag atoms have the configuration [Kr]5$s^1$4d^{10}. The electron lost would be from the 5*s* orbital with quantum numbers $\boldsymbol{n = 5, l = 0, m_l = 0,}$ and $\boldsymbol{m_s = +1/2}$.

d) The F atom has the configuration [He]2$s^2$2p^5. The electron gained would go into the 2p_z orbital and is the second electron in that orbital. Quantum numbers are $\boldsymbol{n = 2, l = 1, m_l = +1,}$ and $\boldsymbol{m_s = -1/2}$.

8.19 a) $\boldsymbol{n = 2; l = 0; m_l = 0; m_s = +1/2}$

b) $\boldsymbol{n = 4; l = 1; m_l = +1; m_s = -1/2}$

c) $\boldsymbol{n = 6; l = 0; m_l = 0; m_s = +1/2}$

d) $\boldsymbol{n = 2; l = 1; m_l = -1; m_s = +1/2}$

8.20 <u>Plan:</u> The atomic number gives the number of electrons and the periodic table shows the order for filling sublevels. Recall that *s* orbitals hold a maximum of 2 electrons, a *p* orbital set holds 6 electrons, a *d* orbital set holds 10 electrons, and an *f* orbital set holds 14 electrons.

<u>Solution:</u>

a) Rb: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1}$

b) Ge: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^2}$

c) Ar: $\mathbf{1s^22s^22p^63s^23p^6}$

8.21 a) Br: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^5}$

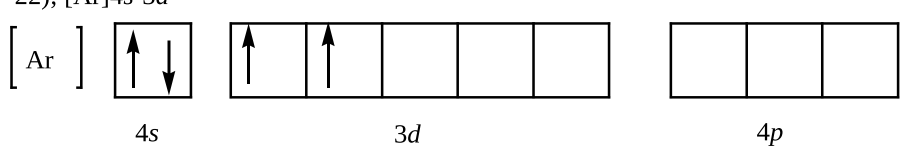
b) Mg: $\mathbf{1s^22s^22p^63s^2}$

c) Se: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^4}$

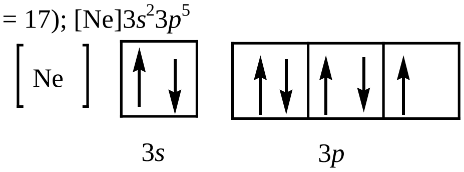
8.22 <u>Plan:</u> The atomic number gives the number of electrons and the periodic table shows the order for filling sublevels. Recall that *s* orbitals hold a maximum of 2 electrons, a *p* orbital set holds 6 electrons, a *d* orbital set holds 10 electrons, and an *f* orbital set holds 14 electrons. Valence electrons are those in the highest energy level; in transition metals, the $(n - 1)d$ electrons are also counted as valence electrons. For a condensed ground-state electron configuration, the electron configuration of the previous noble gas is shown by its element symbol in brackets, followed by the electron configuration of the energy level being filled.

<u>Solution:</u>

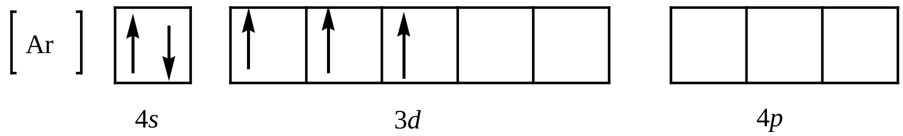
a) Ti ($Z = 22$); [Ar]4$s^2$3d^2



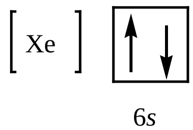
b) Cl ($Z = 17$); [Ne]3$s^2$3p^5



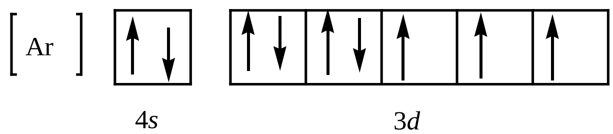
c) V (Z = 23); [Ar]$4s^2 3d^3$



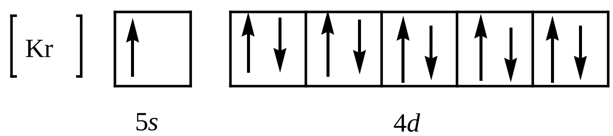
8.23 a) Ba: [Xe]$6s^2$



b) Co: [Ar]$4s^2 3d^7$



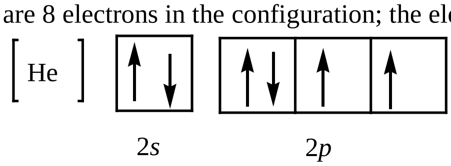
c) Ag: [Kr]$5s^1 4d^{10}$



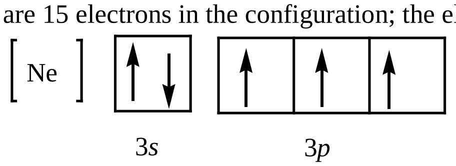
8.24 Plan: Add up all of the electrons in the electron configuration to obtain the atomic number of the element which is then used to identify the element and its position in the periodic table. When drawing the partial orbital diagram, only include electrons after those of the previous noble gas; remember to put one electron in each orbital in a set before pairing electrons.

Solution:

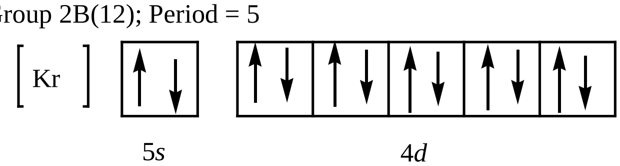
a) There are 8 electrons in the configuration; the element is O, Group 6A(16), Period 2.



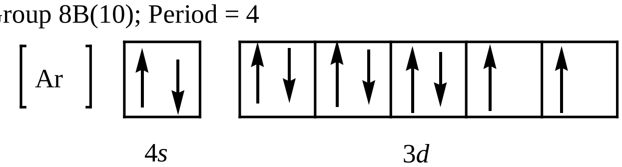
b) There are 15 electrons in the configuration; the element is P, Group 5A(15), Period 3.



8.25 a) Cd; Group 2B(12); Period = 5



b) Ni; Group 8B(10); Period = 4



8.26 Plan: Use the periodic table and the partial orbital diagram to identify the element.
Solution:
a) The orbital diagram shows the element is in Period 4 (n = 4 as outer level). The configuration is $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^1}$ or $[Ar]4s^23d^{10}4p^1$. One electron in the p level indicates the element is in Group **3A(13)**. The element is Ga.
b) The orbital diagram shows the 2s and 2p orbitals filled which would represent the last element in Period 2, Ne. The configuration is $\mathbf{1s^22s^22p^6}$ or $[He]2s^22p^6$. Filled s and p orbitals indicate Group **8A(18)**.

8.27 a) $\mathbf{[Kr]5s^14d^4}$ Nb; **5B(5)**
b) $\mathbf{[He]2s^22p^3}$ N; **5A(15)**

8.28 Plan: Inner electrons are those seen in the previous noble gas and completed transition series (d orbitals). Outer electrons are those in the highest energy level (highest n value). Valence electrons are the outer electrons for main-group elements; for transition metals, valence electrons also include electrons in the outermost d set of orbitals. It is easiest to determine the types of electrons by writing a condensed electron configuration.
Solution:
a) O (Z = 8); $[He]2s^22p^4$. There are **2** inner electrons (represented by [He]) and **6** outer electrons. The number of valence electrons (**6**) equals the outer electrons in this case.
b) Sn (Z = 50); $[Kr]5s^24d^{10}5p^2$. There are 36 (from [Kr]) + 10 (from the filled 4d set) = **46** inner electrons. There are **4** outer electrons (highest energy level is n = 5) and **4** valence electrons.
c) Ca (Z = 20); $[Ar]4s^2$. There are **2** outer electrons (the 4s electrons), **2** valence electrons, and **18** inner electrons (from [Ar]).
d) Fe (Z = 26); $[Ar]4s^23d^6$. There are **2** outer electrons (from n = 4 level), **8** valence electrons (the d orbital electrons count in this case because the sublevel is not full), and **18** inner electrons (from [Ar]).
e) Se (Z = 34); $[Ar]4s^23d^{10}4p^4$. There are **6** outer electrons (2 + 4 in the n = 4 level), **6** valence electrons (filled d sublevels count as inner electrons), and **28** inner electrons (18 from [Ar] and 10 from the filled 3d set).

8.29

	inner electrons	outer electrons	valence electrons
a) Br	**28**	**7**	**7**
b) Cs	**54**	**1**	**1**
c) Cr	**18**	**1**	**6**
d) Sr	**36**	**2**	**2**
e) F	**2**	**7**	**7**

8.30 Plan: Add up all of the electrons in the electron configuration to obtain the atomic number of the element which is then used to identify the element and its position in the periodic table.
Solution:
a) The electron configuration $[He]2s^22p^1$ has a total of 5 electrons (3 + 2 from He configuration) which is element boron with symbol **B**. Boron is in Group 3A(13). Other elements in this group are **Al, Ga, In,** and **Tl**.
b) The electrons in this element total 16, 10 from the neon configuration plus 6 from the rest of the configuration. Element 16 is sulfur, **S**, in Group 6A(16). Other elements in Group 6A(16) are **O, Se, Te,** and **Po**.
c) Electrons total 3 + 54 (from xenon) = 57. Element 57 is lanthanum, **La**, in Group 3B(3). Other elements in this group are **Sc, Y,** and **Ac**.

8.31 a) **Se**; other members **O, S, Te, Po**
b) **Hf**; other members **Ti, Zr, Rf**
c) **Mn**; other members **Tc, Re, Bh**

8.32 a) **Mg**: $\mathbf{[Ne]3s^2}$
b) **Cl**: $\mathbf{[Ne]3s^23p^5}$
c) **Mn**: $\mathbf{[Ar]4s^23d^5}$
d) **Ne**: $\mathbf{[He]2s^22p^6}$

8.33 Atomic size increases down a main group and decreases across a period. Ionization energy decreases down a main group and increases across a period. These opposite trends result because as the atom gets larger, the outer electron is further from the attraction of the positive charge of the nucleus, which is what holds the electron in the atom. It thus takes less energy (lower IE) to remove the outer electron in a larger atom than to remove the outer electron in a smaller atom. As the atomic size decreases across a period due to higher Z_{eff}, it takes more energy (higher IE) to remove the outer electron.

8.34 a) **A = silicon; B = fluorine; C = strontium; D = sulfur**
b) **F < S < Si < Sr**
c) **Sr < Si < S < F**

8.35 High IEs correspond to elements in the upper right of the periodic table, while relatively low IEs correspond to elements at the lower left of the periodic table.

8.36 a) For a given element, successive ionization energies always increase. As each successive electron is removed, the positive charge on the ion increases, which results in a stronger attraction between the leaving electron and the ion.
b) When a large jump between successive ionization energies is observed, the subsequent electron must come from a full lower energy level. Thus, by looking at a series of successive ionization energies, we can determine the number of valence electrons. For instance, the electron configuration for potassium is $[Ar]4s^1$. The first electron lost is the one from the 4*s* level. The second electron lost must come from the 3*p* level, and hence breaks into the core electrons. Thus, we see a significant jump in the amount of energy for the second ionization when compared to the first ionization.
c) There is a large increase in ionization energy from IE_3 to IE_4, suggesting that the element has 3 valence electrons. The Period 2 element would be **B**; the Period 3 element would be **Al;** and the Period 4 element would be **Ga**.

8.37 The first drop occurs because the 3*p* sublevel is higher in energy than the 3*s*, so the 3*p* electron of Al is pulled off more easily than a 3*s* electron of Mg. The second drop occurs because the $3p^4$ electron occupies the same orbital as another 3*p* electron. The resulting electron-electron repulsion raises the orbital energy and thus it is easier to remove an electron from S ($3p^4$) than P ($3p^3$).

8.38 A high, endothermic IE_1 means it is very difficult to remove the first outer electron. This value would exclude any metal, because metals lose an outer electron easily. A very negative, exothermic EA_1 suggests that this element easily gains one electron. These values indicate that the element belongs to the halogens, Group **7A(17)**, which form **–1** ions.

8.39 After an initial shrinking for the first 2 or 3 elements, the size remains relatively constant as the shielding of the 3*d* electrons just counteracts the increase in the number of protons in the nucleus so the Z_{eff} remains relatively constant.

8.40 Plan: Atomic size decreases up a main group and left to right across a period.
Solution:
a) Increasing atomic size: **K < Rb < Cs**; these three elements are all part of the same group, the alkali metals. Atomic size decreases up a main group (larger outer electron orbital), so potassium is the smallest and cesium is the largest.
b) Increasing atomic size: **O < C < Be**; these three elements are in the same period and atomic size decreases across a period (increasing effective nuclear charge), so beryllium is the largest and oxygen the smallest.
c) Increasing atomic size: **Cl < S < K**; chlorine and sulfur are in the same period so chlorine is smaller since it is further to the right in the period. Potassium is the first element in the next period so it is larger than either Cl or S.
d) Increasing atomic size: **Mg < Ca < K**; calcium is larger than magnesium because Ca is further down the alkaline earth metal group on the periodic table than Mg. Potassium is larger than calcium because K is further to the left than Ca in Period 4 of the periodic table.

8.41 a) **Pb > Sn > Ge** b) **Sr > Sn > Te** c) **Na > F > Ne** d) **Na > Mg > Be**

8.42 Plan: Ionization energy increases up a group and left to right across a period.
Solution:
a) **Ba < Sr < Ca** The "group" rule applies in this case. Ionization energy increases up a main group. Barium's outer electron receives the most shielding; therefore, it is easiest to remove and has the lowest IE.
b) **B < N < Ne** These elements have the same *n*, so the "period" rule applies. Ionization energy increases from left to right across a period. B experiences the lowest Z_{eff} and has the lowest IE. Ne has the highest IE, because it's very difficult to remove an electron from the stable noble gas configuration.
c) **Rb < Se < Br** IE decreases with increasing atomic size, so Rb (largest atom) has the smallest IE. Se has a lower IE than Br because IE increases across a period.
d) **Sn < Sb < As** IE increases up a group, so Sn and Sb will have smaller IEs than As. The "period" rule applies for ranking Sn and Sb.

8.43 a) **Li > Na > K** b) **F > C > Be** c) **Ar > Cl > Na** d) **Cl > Br > Se**

8.44 Plan: When a large jump between successive ionization energies is observed, the subsequent electron must come from a full lower energy level. Thus, by looking at a series of successive ionization energies, we can determine the number of valence electrons. The number of valence electrons identifies which group the element is in.
Solution:
The successive ionization energies show a very significant jump between the third and fourth IEs. This indicates that the element has three valence electrons. The fourth electron must come from the core electrons and thus has a very large ionization energy. The electron configuration of the Period 2 element with three valence electrons is $\mathbf{1}\boldsymbol{s}^{\mathbf{2}}\mathbf{2}\boldsymbol{s}^{\mathbf{2}}\mathbf{2}\boldsymbol{p}^{\mathbf{1}}$ which represents boron, **B**.

8.45 The successive ionization energies show a significant jump between the second and third IEs, indicating that the element has only two valence electrons. The configuration is $\mathbf{1}\boldsymbol{s}^{\mathbf{2}}\mathbf{2}\boldsymbol{s}^{\mathbf{2}}\mathbf{2}\boldsymbol{p}^{\mathbf{6}}\mathbf{3}\boldsymbol{s}^{\mathbf{2}}$**, Mg**.

8.46 Plan: For a given element, successive ionization energies always increase. As each successive electron is removed, the positive charge on the ion increases, which results in a stronger attraction between the leaving electron and the ion. A very large jump between successive ionization energies will occur when the electron to be removed comes from a full lower energy level. Examine the electron configurations of the atoms. If the IE_2 represents removing an electron from a full orbital, then the IE_2 will be very large. In addition, for atoms with the same outer electron configuration, IE_2 is larger for the smaller atom.
Solution:
a) **Na** would have the highest IE_2 because ionization of a second electron would require breaking the stable [Ne] configuration:
First ionization: Na ([Ne]$3s^1$) $\rightarrow$ Na^+ ([Ne]) + e^- (low IE)
Second ionization: Na^+ ([Ne]) $\rightarrow$ Na^{+2} ([He]$2s^22p^5$) + e^- (high IE)
b) **Na** would have the highest IE_2 because it has one valence electron and is smaller than K.
c) You might think that Sc would have the highest IE_2, because removing a second electron would require breaking the stable, filled 4*s* shell. However, **Be** has the highest IE_2 because Be's small size makes it difficult to remove a second electron.

8.47 a) **Al** b) **Sc** c) **Al**

8.48 Three of the ways that metals and nonmetals differ are: 1) metals conduct electricity, nonmetals do not; 2) when they form stable ions, metal ions tend to have a positive charge, nonmetal ions tend to have a negative charge; and 3) metal oxides are ionic and act as bases, nonmetal oxides are covalent and act as acids.

8.49 Metallic character increases down a group and decreases toward the right across a period. These trends are the same as those for atomic size and opposite those for ionization energy.

8.50 Generally, oxides of metals are basic while oxides of nonmetals are acidic. As the metallic character decreases, the oxide becomes more acidic. Thus, oxide acidity increases from left to right across a period and from bottom to top in a group.

8.51 An $(n-1)d^{10}ns^0np^0$ configuration is called a pseudo-noble gas configuration. In^{3+}: $[Kr]4d^{10}$

8.52 **3+ < 2+ < 1+ < 0 < 1– < 2– < 3–**

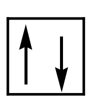
8.53 Plan: Metallic behavior decreases up a group and decreases left to right across a period.
Solution:
a) **Rb** is more metallic because it is to the left and below Ca.
b) **Ra** is more metallic because it lies below Mg in Group 2A(2).
c) **I** is more metallic because it lies below Br in Group 7A(17).

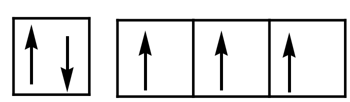
8.54 a) **S** b) **In** c) **As**

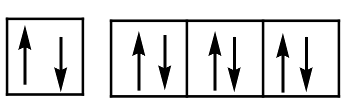
8.55 Plan: For main-group elements, the most stable ions have electron configurations identical to noble gas atoms. Write the electron configuration of the atom and then remove or add electrons until a noble gas configuration is achieved. Metals lose electrons and nonmetals gain electrons.
Solution:
a) Cl: $1s^22s^22p^63s^23p^5$; chlorine atoms are one electron short of a noble gas configuration, so a **–1** ion will form by adding an electron to have the same electron configuration as an argon atom: Cl^-, $\mathbf{1s^22s^22p^63s^23p^6}$.
b) Na: $1s^22s^22p^63s^1$; sodium atoms contain one more electron than the noble gas configuration of neon. Thus, a sodium atom loses one electron to form a **+1** ion: Na^+, $\mathbf{1s^22s^22p^6}$.
c) Ca: $1s^22s^22p^63s^23p^64s^2$; calcium atoms contain two more electrons than the noble gas configuration of argon. Thus, a calcium atom loses two electrons to form a **+2** ion: Ca^{2+}, $\mathbf{1s^22s^22p^63s^23p^6}$.

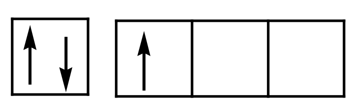
8.56 a) Rb^+: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^6}$ **+1**
b) N^{3-}: $\mathbf{1s^22s^22p^6}$ **–3**
c) Br^-: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^6}$ **–1**

8.57 Plan: To find the number of unpaired electrons look at the electron configuration expanded to include the different orientations of the orbitals, such as p_x and p_y and p_z. Remember that one electron will occupy every orbital in a set (p, d, or f) before electrons will pair in an orbital in that set. In the noble gas configurations, all electrons are paired because all orbitals are filled.
Solution:
a) Configuration of 2A(2) group elements: [noble gas]ns^2, **no unpaired electrons**. The electrons in the *ns* orbital are paired.
b) Configuration of 5A(15) group elements: [noble gas]$ns^2np_x^1np_y^1np_z^1$. **Three** unpaired electrons, one each in
p_x, p_y, and p_z.
c) Configuration of 8A(18) group elements: noble gas configuration ns^2np^6 with no half-filled orbitals, **no unpaired electrons**.
d) Configuration of 3A(13) group elements: [noble gas]ns^2np^1. There is **one** unpaired electron in one of the *p* orbitals.


a) *ns*


b) *ns* *np*


c) *ns* *np*


d) *ns* *np*

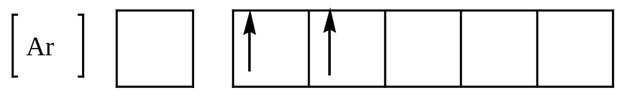
8.58 To find the number of unpaired electrons look at the electron configuration expanded to include the different orientations of the orbitals, such as p_x and p_y and p_z. In the noble gas configurations, all electrons are paired because all orbitals are filled.

a) Configuration of 4A(14) group elements: [noble gas]$ns^2np_x^1np_y^1np_z^0$. **Two** unpaired electrons.
b) Configuration of 7A(17) group elements: [noble gas]$ns^2np_x^2np_y^2np_z^1$. **One** unpaired electron.
c) Configuration of 1A(1) group elements: [noble gas]ns^1. **One** unpaired electron.
d) Configuration of 6A(16) group elements: [noble gas]$ns^2np_x^2np_y^1np_z^1$. **Two** unpaired electrons.

8.59 Plan: Substances are paramagnetic if they have unpaired electrons. Write the electron configuration of the atom and then remove the specified number of electrons. Remember that all orbitals in a *p*, *d*, or *f* set will each have one electron before electrons pair in an orbital. In the noble gas configurations, all electrons are paired because all orbitals are filled.

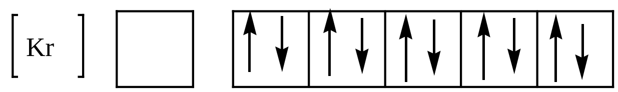
Solution:

a) V: [Ar]$4s^23d^3$; **V^{3+}: [Ar]$3d^2$** Transition metals first lose the *s* electrons in forming ions, so to form the +3 ion a vanadium atom loses two 4*s* electrons and one 3*d* electron. **Paramagnetic**



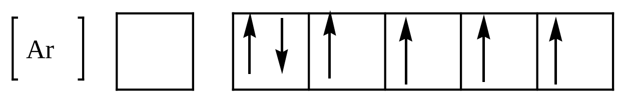
4*s* 3*d*

b) Cd: [Kr]$5s^24d^{10}$; **Cd^{2+}: [Kr]$4d^{10}$** Cadmium atoms lose two electrons from the 4*s* orbital to form the +2 ion. **Diamagnetic**



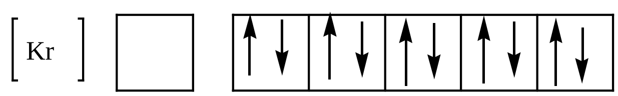
5*s* 4*d*

c) Co: [Ar]$4s^23d^7$; **Co^{3+}: [Ar]$3d^6$** Cobalt atoms lose two 4*s* electrons and one 3*d* electron to form the +3 ion. **Paramagnetic**



4*s* 3*d*

d) Ag: [Kr]$5s^14d^{10}$; **Ag^+: [Kr]$4d^{10}$** Silver atoms lose the one electron in the 5*s* orbital to form the +1 ion. **Diamagnetic**



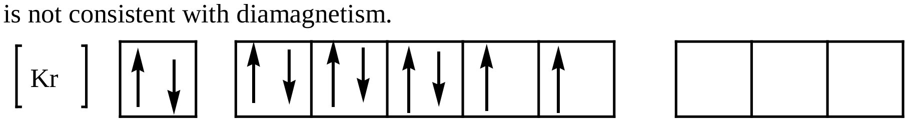
5*s* 4*d*

8.60 a) Mo^{3+}: [Kr]$4d^3$ **paramagnetic**
b) Au^+: [Xe]$4f^{14}5d^{10}$ **diamagnetic**
c) Mn^{2+}: [Ar]$3d^5$ **paramagnetic**
d) Hf^{2+}: [Xe]$4f^{14}5d^2$ **paramagnetic**

8.61 Plan: Substances are diamagnetic if they have no unpaired electrons. Draw the partial orbital diagrams, remembering that all orbitals in *d* set will each have one electron before electrons pair in an orbital.

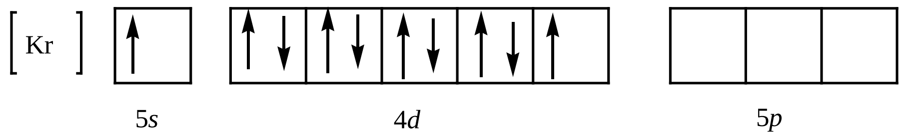
Solution:

You might first write the condensed electron configuration for Pd as [Kr]$5s^24d^8$. However, the partial orbital diagram is not consistent with diamagnetism.

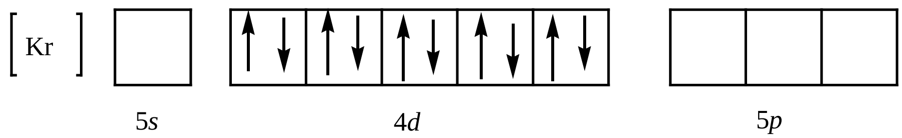


5*s* 4*d* 5*p*

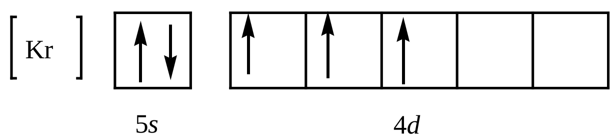
Promoting an *s* electron into the *d* sublevel (as in (c) [Kr]$5s^14d^9$) still leaves two electrons unpaired.



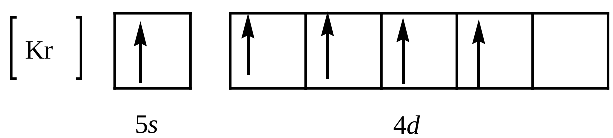
The only configuration that supports diamagnetism is **(b) [Kr]4d^{10}**.



8.62 The expected electron configuration for Group 5B(5) elements is $ns^2(n–1)\ d^3$.
Nb (expected): [Kr]5$s^2$4d^3 3 unpaired e$^–$



Nb (actual): [Kr]5$s^1$4d^4 5 unpaired e$^–$



8.63 Plan: The size of ions increases down a group. For ions that are isoelectronic (have the same electron configuration) size decreases with increasing atomic number.
Solution:
a) Increasing size: **Li^+ < Na^+ < K^+**, size increases down Group 1A(1).
b) Increasing size: **Rb^+ < $Br^–$ < $Se^{2–}$**, these three ions are isoelectronic with the same electron configuration as krypton. Size decreases with increasing atomic number in an isoelectronic series.
c) Increasing size: **$F^–$ < $O^{2–}$ < $N^{3–}$**, the three ions are isoelectronic with an electron configuration identical to neon. Size decreases with increasing atomic number in an isoelectronic series.

8.64 a) **$Se^{2–}$ > $S^{2–}$ > $O^{2–}$**, size increases down a group.
b) **$Te^{2–}$ > $I^–$ > Cs^+**, size decreases with increasing atomic number in an isoelectronic series.
c) **Cs^+ > Ba^{2+} > Sr^{2+}**, both reasons as in parts a) and b).

8.65

a) **oxygen**	b) **cesium**	c) **aluminum**	d) **carbon**	e) **rubidium**	f) **bismuth**
g) **thallium**	h) **krypton**	i) **silicon**	j) **ruthenium**	k) **vanadium**	l) **indium**
m) **scandium**	n) **manganese**	o) **lutetium**	p) **sulfur**	q) **strontium**	r) **arsenic**

8.66 Plan: Write the formula of the oxoacid. Remember that in naming oxoacids (H + polyatomic ion), the suffix of the polyatomic changes: -ate becomes -ic acid and -ite becomes -ous acid. Determine the oxidation state of the nonmetal in the oxoacid; hydrogen has an O.N. of +1 and oxygen has an O.N. of –2. Based on the oxidation state of the nonmetal, and the oxidation state of the oxide ion (–2), the formula of the nonmetal oxide may be determined. The name of the nonmetal oxide comes from the formula; remember that nonmetal compounds use prefixes to indicate the number of each type of atom in the formula.
Solution:
a) hypochlorous acid = HClO has Cl^+ so the oxide is **Cl_2O** = **dichlorine oxide** or **dichlorine monoxide**
b) chlorous acid = $HClO_2$ has Cl^{3+} so the oxide is **Cl_2O_3** = **dichlorine trioxide**
c) chloric acid = $HClO_3$ has Cl^{5+} so the oxide is **Cl_2O_5** = **dichlorine pentaoxide**
d) perchloric acid = $HClO_4$ has Cl^{7+} so the oxide is **Cl_2O_7** = **dichlorine heptaoxide**
e) sulfuric acid = H_2SO_4 has S^{6+} so the oxide is **SO_3** = **sulfur trioxide**
f) sulfurous acid = H_2SO_3 has S^{4+} so the oxide is **SO_2** = **sulfur dioxide**
g) nitric acid = HNO_3 has N^{5+} so the oxide is **N_2O_5** = **dinitrogen pentaoxide**

h) nitrous acid = HNO_2 has N^{3+} so the oxide is $\mathbf{N_2O_3}$ = **dinitrogen trioxide**
i) carbonic acid = H_2CO_3 has C^{4+} so the oxide is $\mathbf{CO_2}$ = **carbon dioxide**
j) phosphoric acid = H_3PO_4 has P^{5+} so the oxide is $\mathbf{P_2O_5}$ = **diphosphorus pentaoxide or** $\mathbf{P_4O_{10}}$ = **tetraphosphorus decaoxide**.

8.67 $$\lambda = hc/\Delta E = \frac{\left(6.626\text{x}10^{-34}\ \text{J}\bullet\text{s}\right)\left(3.00\text{x}10^{8}\ \text{m/s}\right)}{\left(2.7\ \text{eV}\right)\left(\frac{1.602\text{x}10^{-19}\ \text{J}}{1\ \text{eV}}\right)} = 4.59564\text{x}10^{-7} = 4.6\text{x}10^{-7}\ \text{m}$$

The absorption of light of this wavelength (blue) leads to the complimentary color (yellow) being seen. An electron in gold's 5*d* subshell can absorb blue light in its transition to a 6*s* subshell, giving gold its characteristic "gold" color.

8.68 Plan: Remember that isoelectronic species have the same electron configuration. Atomic radius decreases up a group and left to right across a period.
Solution:
a) A chemically unreactive Period 4 element would be Kr in Group 8A(18). Both the Sr^{2+} ion and Br^- ion are isoelectronic with Kr. Their combination results in $\mathbf{SrBr_2}$**, strontium bromide**.
b) Ar is the Period 3 noble gas. Ca^{2+} and S^{2-} are isoelectronic with Ar. The resulting compound is **CaS**, **calcium sulfide**.
c) The smallest filled *d* subshell is the 3*d* shell, so the element must be in Period 4. Zn forms the Zn^{2+} ion by losing its two *s* subshell electrons to achieve a *pseudo–noble gas* configuration ($[Ar]3d^{10}$). The smallest halogen is fluorine, whose anion is F^-. The resulting compound is $\mathbf{ZnF_2}$**, zinc fluoride**.
d) Ne is the smallest element in Period 2, but it is not ionizable. Li is the largest atom whereas F is the smallest atom in Period 2. The resulting compound is **LiF**, **lithium fluoride**.

8.69 Plan: Determine the electron configuration for iron, and then begin removing one electron at a time. Remember that all orbitals in a *d* set will each have one electron before electrons pair in an orbital, and electrons with the highest *n* value are removed first. Ions with all electrons paired are diamagnetic. Ions with at least one unpaired electron are paramagnetic. The more unpaired electrons, the greater the attraction to a magnetic field.
Solution:

Fe	$[Ar]4s^23d^6$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 4
Fe^+	$[Ar]4s^13d^6$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 5
Fe^{2+}	$[Ar]3d^6$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 4
Fe^{3+}	$[Ar]3d^5$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 5
Fe^{4+}	$[Ar]3d^4$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 4
Fe^{5+}	$[Ar]3d^3$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 3
Fe^{6+}	$[Ar]3d^2$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 2
Fe^{7+}	$[Ar]3d^1$	partially filled 3*d* = **paramagnetic**	number of unpaired electrons = 1
Fe^{8+}	[Ar]	filled orbitals = **diamagnetic**	number of unpaired electrons = 0
Fe^{9+}	$[Ne]3s^23p^5$	partially filled 3*p* = **paramagnetic**	number of unpaired electrons = 1
Fe^{10+}	$[Ne]3s^23p^4$	partially filled 3*p* = **paramagnetic**	number of unpaired electrons = 2
Fe^{11+}	$[Ne]3s^23p^3$	partially filled 3*p* = **paramagnetic**	number of unpaired electrons = 3
Fe^{12+}	$[Ne]3s^23p^2$	partially filled 3*p* = **paramagnetic**	number of unpaired electrons = 2
Fe^{13+}	$[Ne]3s^23p^1$	partially filled 3*p* = **paramagnetic**	number of unpaired electrons = 1
Fe^{14+}	$[Ne]3s^2$	filled orbitals = **diamagnetic**	number of unpaired electrons = 0

$\mathbf{Fe^+}$ and $\mathbf{Fe^{3+}}$ would both be most attracted to a magnetic field. They each have 5 unpaired electrons.

8.70 a) Rubidium atoms form **+1** ions, Rb^+; bromine atoms form **–1** ions, Br^-.
b) Rb: $[Kr]5s^1$; Rb^+: [Kr]; Rb^+ is a diamagnetic ion that is isoelectronic with Kr.
Br: $[Ar]4s^23d^{10}4p^5$; Br^-: $[Ar]4s^23d^{10}4p^6$ or [Kr]; Br^- is a diamagnetic ion that is isoelectronic with Kr.
c) Rb^+ is a smaller ion than Br^-; **B** best represents the relative ionic sizes.

8.71 a): $X^{2+} = [Kr]4d^8$; $X = [Kr]5s^24d^8$. The element is **palladium** and the oxide is **PdO**.
b): $X^{3+} = [Ar]3d^6$; $X = [Ar]4s^23d^7$. The element is **cobalt** and the oxide is $\mathbf{Co_2O_3}$.
c): $X^+ = [Kr]4d^{10}$; $X = [Kr]5s^14d^{10}$. The element is **silver** and the oxide is $\mathbf{Ag_2O}$.

d): X^{+4} = [Ar]$3d^3$; X = [Ar]$4s^2 3d^5$. The element is **manganese** and the oxide is **MnO_2**.

8.72 There is a large increase in ionization energy from IE_3 to IE_4, suggesting that the element has 3 valence electrons. The Period 2 element would be **B**; the Period 3 element would be **Al** and the Period 4 element would be **Ga**.

8.73

balloonium	=	helium
inertium	=	neon
allotropium	=	sulfur
brinium	=	sodium
canium	=	tin
fertilium	=	nitrogen
liquidium	=	bromine
utilium	=	aluminum
crimsonium	=	strontium