CHAPTER 8 ELECTRON CONFIGURATION AND CHEMICAL PERIODICITY

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 AM(K) =

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

b) Predicted MP(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 a) Predicted BP(HBr) =

$$\frac{\text{HCl} + \text{HI}}{2} = \frac{-84.9 + (-35.4)}{2} = -60.15 = \mathbf{-60.2°C} \qquad \text{(actual value = -67.0°C)}$$

b) Predicted BP(AsH_3) =

$$\frac{PH_3 + SbH_3}{2} = \frac{-87.4 + (-17.1)}{2} = -52.25 = \mathbf{-52.2°C} \qquad \text{(actual value = -67.0°C)}$$

8.5 The allowed vales of n: 1, 2, 3, 4...∞
The allowed values of l: 0, 1, 2, ... n-1
The allowed values of m_l: $-l$, $(-l + 1)$, ... 0, ... $(l - 1)$, $+l$
The allowed values of m_s: ±1/2

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

8.7 Within an atom, no two electrons may have the same four quantum numbers. Within a particular orbital, there can be only two electrons and they must have paired spins.

8.8 In a one-electron system, all sublevels of a particular level 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.

8.9 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, the effective nuclear charge decreases.

8.10 The penetration effect 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.

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

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

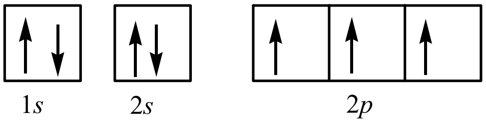
8.13 a) **6** electrons can be found in the three 4*p* orbitals, 2 in each orbital.
b) **2** electrons can have the 3 quantum numbers given for one of the 3*p* orbitals with the difference between the two being in the m_s quantum number.
c) **14** electrons can be found in the 5*f* orbitals ($l = 3$ designates *f* orbitals).

8.14 a) **2** 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.
c) A maximum of **10** electrons can be found in the *d* orbitals.

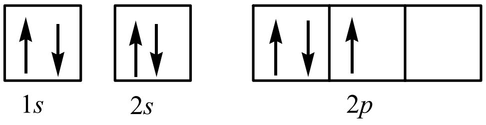
8.15 Properties recur periodically due to similarities in electron configurations recurring periodically.
Na: $1s^2\ 2s^2\ 2p^6 \mid \underline{3s^1}$
K: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6 \mid \underline{4s^1}$
The properties of Na and K are similar due to a similarity in their outer shell electron configuration; both have one electron in an outer shell *s* orbital.

8.16 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. The correct electron configuration for nitrogen is shown in (a), which is contrasted to an incorrect configuration shown in (b). The arrows in the 2*p* orbitals of configuration (a) could alternatively all point down.

(a) - correct



(b) - incorrect



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

8.18 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.

8.19 Outer electrons are the same as valence electrons for the main-group elements. The *d* electrons are often included among the valence electrons for transition elements.

8.20 The total electron capacity for an energy level is $2n^2$, so the $n = 4$ energy level holds a maximum of $2(4^2) = 32$ electrons. A filled $n = 4$ energy level would have the following configuration: $4s^24p^64d^{10}4f^{14}$.

8.21 Plan: 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 an $m_s = +1/2$. 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.

Solution:

a) The outermost electron in a rubidium atom would be in a $5s$ orbital. The quantum numbers for this electron are **$n = 5, l = 0, m_l = 0,$** and **$m_s = +1/2$.**

b) The S^- ion would have the configuration $[Ne]3s^23p^5$. The electron added would go into the $3p_z$ orbital. Quantum numbers are **$n = 3, l = 1, m_l = +1,$** and **$m_s = -1/2$**.

c) Ag atoms have the configuration $[Kr]5s^14d^{10}$. The electron lost would be from the $5s$ orbital with quantum numbers **$n = 5, l = 0, m_l = 0,$** and **$m_s = +1/2$**.

d) The F atom has the configuration $[He]2s^22p^5$. The electron gained would go into the $2p_z$ orbital. Quantum numbers are **$n = 2, l = 1, m_l = +1,$** and **$m_s = -1/2$**.

8.22 a) **$n = 2; l = 0; m_l = 0; m_s = +1/2$**
b) **$n = 4; l = 1; m_l = 1; m_s = -1/2$**
c) **$n = 6; l = 0; m_l = 0; m_s = +1/2$**
d) **$n = 2; l = 1; m_l = -1; m_s = +1/2$**

8.23 a) Rb: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^1$**
b) Ge: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^2$**
c) Ar: **$1s^22s^22p^63s^23p^6$**

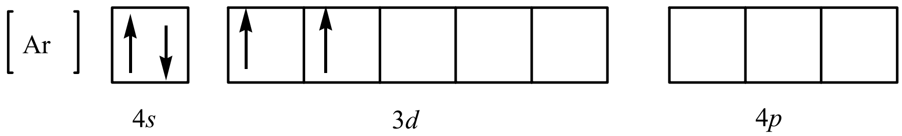
8.24 a) Br: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^5$**
b) Mg: **$1s^22s^22p^63s^2$**
c) Se: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^4$**

8.25 a) Cl: **$1s^22s^22p^63s^23p^5$**
b) Si: **$1s^22s^22p^63s^23p^2$**
c) Sr: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$**

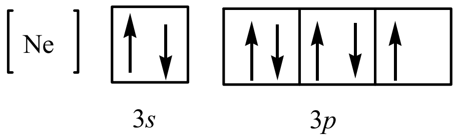
8.26 a) S: **$1s^22s^22p^63s^23p^4$**
b) Kr: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$**
c) Cs: **$1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^1$**

8.27 Valence electrons are those electrons beyond the previous noble gas configuration and unfilled *d* and *f* sublevels.

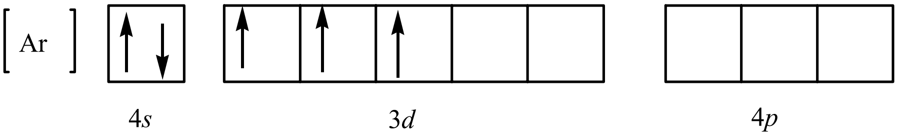
a) Ti ($Z = 22$); $[Ar]4s^23d^2$



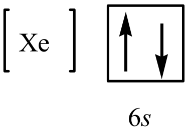
b) Cl ($Z = 17$); $[Ne]3s^23p^5$



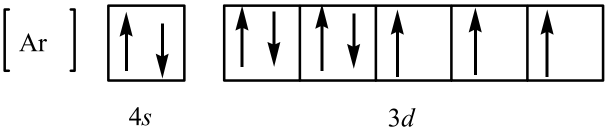
c) V ($Z = 23$); $[\mathrm{Ar}]4s^2 3d^3$



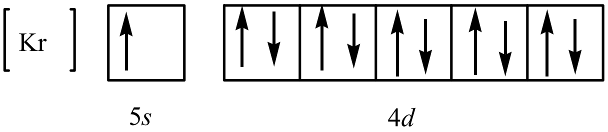
8.28 a) Ba: $[\mathrm{Xe}]6s^2$



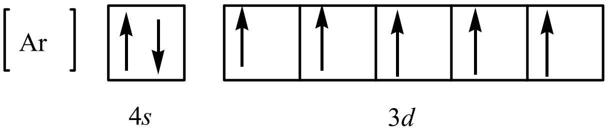
b) Co: $[\mathrm{Ar}]4s^2 3d^7$



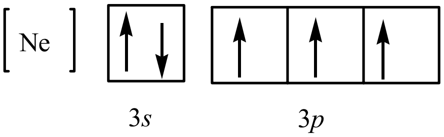
c) Ag: $[\mathrm{Kr}]5s^1 4d^{10}$



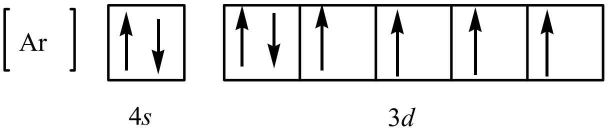
8.29 a) Mn ($Z = 25$); $[\mathrm{Ar}]4s^2 3d^5$



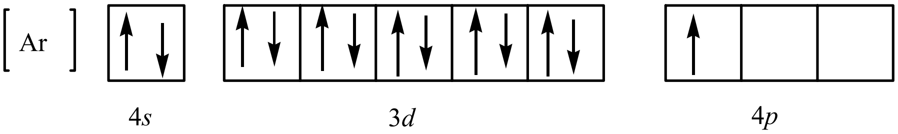
b) P ($Z = 15$); $[\mathrm{Ne}]3s^2 3p^3$



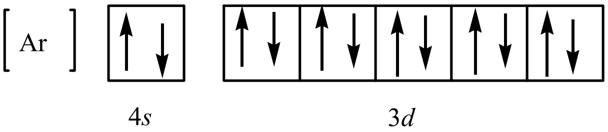
c) Fe ($Z = 26$); $[\mathrm{Ar}]4s^2 3d^6$



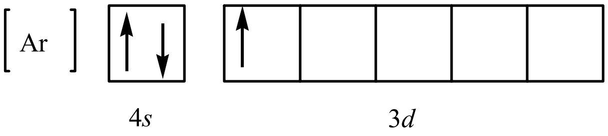
8.30 a) Ga: $[\mathrm{Ar}]4s^2 3d^{10} 4p^1$



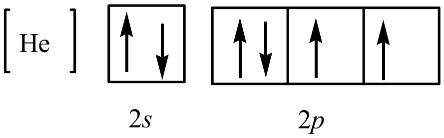
b) Zn: $[\mathrm{Ar}]4s^2 3d^{10}$



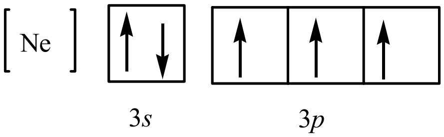
c) Sc: [Ar]$4s^2 3d^1$



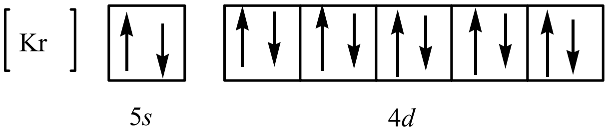
8.31 a) Element = O, Group 6A(16), period 2



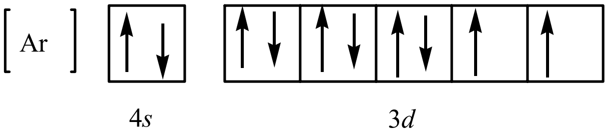
b) Element = P, Group 5A(15), period 3



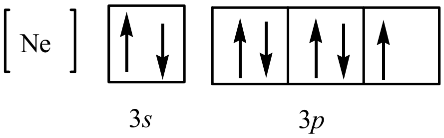
8.32 a) Cd; Group 2B (12); period = 5



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

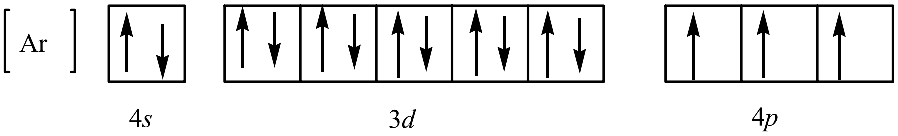


8.33 a)



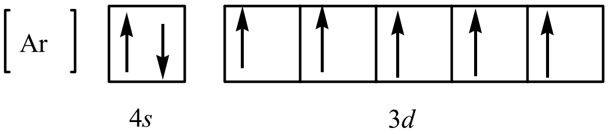
Cl; Group 7A(17); period 3

b)

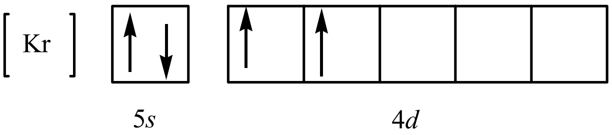


As; Group 5A(15); period 4

8.34 a) Mn; Group 7B (7); period = 4



b) Zr; Group 4B (4); period = 5



8.35 a) The orbital diagram shows the element is in period 4 (n = 4 as outer level). The configuration is **$1s^22s^22p^63s^23p^64s^23d^{10}4p^1$** 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 **$1s^22s^22p^6$** or [He]$2s^22p^6$. Filled s and p orbitals indicate the group **8A(18)**.

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

8.37 It is easiest to determine the types of electrons by writing a condensed electron configuration.
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 d level) = **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, **2** valence electrons, and **18** inner electrons.
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.
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** ((18 + 10) or (34 - 6)) inner electrons.

8.38

	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**

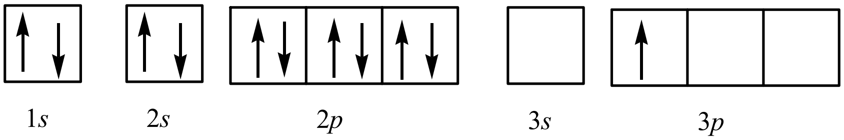
8.39 a) The electron configuration [He]$2s^22p^1$ has a total of 5 electrons (3 + 2 from He configuration) which is element boron with symbol **B**. Boron is in group 3A(13). Other elements in this group are **Al, Ga, In,** and **Tl**.
b) The electrons in this element total 16, 10 from the neon configuration 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**.

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

8.41 a) **Carbon**; other Group 4A(14) elements include **Si, Ge, Sn,** and **Pb**.
b) **Vanadium**; other Group 5B(5) elements include **Nb, Ta,** and **Db**.
c) **Phosphorus**; other Group 5A(15) elements include **N, As, Sb,** and **Bi**.

8.42 a) **Ge**; other members **C, Si, Sn, Pb**
b) **Co**; other members **Rh, Ir, Mt**
c) **Tc**; other members **Mn, Re, Bh**

8.43 The ground state configuration of Na is $1s^22s^22p^63s^1$. Upon excitation, the $3s^1$ electron is promoted to the $3p$ level, with configuration $1s^22s^22p^63p^1$.



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

8.45 The size of an atom may be defined in terms of how closely it lies to a neighboring atom. The metallic radius is one-half the distance between nuclei of adjacent atoms in a crystal of the element (typically metals). The covalent radius is one-half the distance between nuclei of identical covalently bonded atoms in molecules.

8.46 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.

8.47 High IE's correspond to elements in the upper right of the periodic table, while relatively low IE's correspond to elements at the lower left of the periodic table.

8.48 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.

When a large jump between successive ionization energies is observed, the subsequent electron must come from a 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 $4s$ level. The second electron lost must come from the $3p$ 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.

8.49 The first drop occurs because the $3p$ sublevel is higher in energy than the $3s$, so the $3p$ electron of Al is pulled off more easily than a 3s electron of Mg. The second drop occurs because the 3p4 electron occupies the same orbital as another 3p electron. The resulting electron-electron repulsion raises the orbital energy and thus it is easier to remove an electron from S (3p4) than P (3p3).

8.50 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.

8.51 This can be seen in the electron affinity values for oxygen. EA_1 for oxygen is negative because energy is released when an electron is added to the neutral atom due to its attraction to the atom's nuclear charge. The EA_2 for oxygen is positive. The second electron affinity is always positive (greater energy) because it requires energy to add a (negative) electron to a (negative) anion.

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

8.53 a) Increasing atomic size: **K < Rb < Cs**, these three elements are all part of the same group, the alkali metals. Atomic size increases down 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.

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

8.55 a) **Ba < Sr < Ca** The "group" rule applies in this case. 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. 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 decreases down a group, so Sn and Sb will have smaller IE's than As. The "period" rule applies for ranking Sn and Sb.

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

8.57 The successive ionization energies show a significant jump between the first and second IE's and between the third and fourth IE's. This indicates that 1) the first electron removed occupies a different orbital than the second electron removed, 2) the second and third electrons occupy the same orbital, 3) the third and fourth electrons occupy different orbitals, and 4) the fourth and fifth electrons occupy the same orbital. The electron configurations for period 2 elements range from $1s^22s^1$ for lithium to $1s^22s^22p^6$ for neon. The three different orbitals are $1s$, $2s$, and $2p$. The first electron is removed from the $2p$ orbital and the second from the $2s$ orbital in order for there to be a significant jump in the ionization energy between the two. The third electron is removed from the $2s$ orbital while the fourth is removed from the $1s$ orbital since IE_4 is much greater than IE_3. The configuration is $\mathbf{1s^22s^22p^1}$ which represents the five electrons in boron, **B**.

8.58 See problem 8.57 for an explanation. **Mg**: $\mathbf{1s^22s^22p^63s^2}$

8.59 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 $4s$ shell. However, **Be** has the highest IE_2 because Be's small size makes it difficult to remove a second electron.

8.60 a) **Al** b) **Sc** c) **Al**

8.61 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. How many other differences are there?

8.62 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.

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

8.64 Larger elements in group 4A, Sn and Pb, have atomic electron configurations that look like $ns^2(n - 1)d^{10}np^2$. Both of these elements are metals so they will form positive ions. To reach the noble gas configuration of xenon the atoms would have to lose 14 electrons, which is not likely. Instead the atoms lose either 2 or 4 electrons to attain a stable configuration with either the ns and $(n - 1)d$ filled for the 2+ ion or the $(n - 1)d$ orbital filled for the 4+ ion.
The Sn^{2+} and Pb^{2+} ions form by losing the two p electrons:
$Sn([Kr]5s^24d^{10}5p^2) \rightarrow Sn^{2+}([Kr]5s^24d^{10}) + 2e^-$
$Pb([Xe]6s^25d^{10}6p^2) \rightarrow Pb^{2+}([Xe]6s^25d^{10}) + 2e^-$
The Sn^{4+} and Pb^{4+} ions form by losing the two p and two s electrons:
$Sn([Kr]5s^24d^{10}5p^2) \rightarrow Sn^{4+}([Kr]4d^{10}) + 4e^-$
$Pb([Xe]6s^25d^{10}6p^2) \rightarrow Pb^{4+}([Xe]5d^{10}) + 4e^-$
Possible ions for tin and lead have **+2** and **+4** charges. The +2 ions form by loss of the outermost two p electrons, while the +4 ions form by loss of these and the outermost two s electrons.

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

8.66 Paramagnetism is the tendency of a species with unpaired electrons to be attracted by an external magnetic field. The degree of paramagnetism increases with the number of unpaired electrons, so that the number of unpaired electrons may be determined by magnetic measurements. A substance with no unpaired electrons is weakly repelled from a magnetic field and is said to be diamagnetic.
Cu^{1+}: $[Ar]3d^{10}$ thus, diamagnetic
Cu^{2+}: $[Ar]3d^9$ thus, paramagnetic, with one unpaired electron

8.67 **3+ < 2+ < 1+ < 0 < 1- < 2- < 3-**

8.68 Metallic behavior increases down a group and decreases across a period.
a) **Rb** is more metallic because it is to the left and below Ca.
b) **Ra** is more metallic because it lies below Mg.
c) **I** is more metallic because it lies below Br.

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

8.70 a) **As** should be less metallic than antimony because it lies above Sb in the same group of the periodic table.
b) **P** should be less metallic because it lies to the right of silicon in the same period of the periodic table.
c) **Be** should be less metallic since it lies above and to the right of sodium on the periodic table.

8.71 a) **Rn** b) **Te** c) **Se**

8.72 The reaction of a *nonmetal* oxide in water produces an *acidic* solution. An example of a Group 6A(16) nonmetal oxide is $SO_2(g)$: $SO_2(g) + H_2O(g) \rightarrow H_2SO_3(aq)$

8.73 Basic solution
$SrO(s) + H_2O(l) \rightarrow Sr(OH)_2(aq)$

8.74 For main group elements, the most stable ions have electron configurations identical to noble gas atoms.
a) Cl^-, **$1s^22s^22p^63s^23p^6$**, chlorine atoms are one electron short of the noble gas configuration, so a **-1** ion will form by adding an electron to have the same electron configuration as an argon atom.
b) Na^+, **$1s^22s^22p^6$**, 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.
c) Ca^{2+}, **$1s^22s^22p^63s^23p^6$**, 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.

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

8.76 a) Elements in the main group form ions that attain a filled outer level, or noble gas configuration. In this case, Al achieves the noble gas configuration of Ne ($1s^22s^22p^6$) by losing its 3 outer shell electrons to form **Al^{3+}**.
Al ($1s^22s^22p^63s^23p^1$) → Al^{3+} (**$1s^22s^22p^6$**) + 3 e^-
b) Sulfur achieves the noble gas configuration of Ar by gaining 2 electrons to form **S^{2-}**.
S ($1s^22s^22p^63s^23p^4$) + 2e^- → S^{2-} (**$1s^22s^22p^63s^23p^6$**)
c) Strontium achieves the noble gas configuration of Kr by losing 2 electrons to form **Sr^{2+}**.
Sr ($1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$) → Sr^{2+} (**$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$**) + 2 e^-

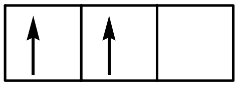
8.77 a) **P^{3-}** **$1s^22s^22p^63s^23p^6$**
b) **Mg^{2+}** **$1s^22s^22p^6$**
c) **Se^{2-}** **$1s^22s^22p^63s^23p^64s^23d^{10}4p^6$**

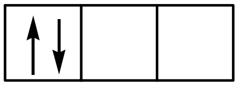
8.78 To find the number of unpaired electrons look at the electron configuration expanded to include the different orientations of the orbitals, such as p_x and p_y and p_z. In the noble gas configurations, all electrons are paired because all orbitals are filled.
a) Configuration of 2A(2) group elements: [noble gas]ns^2, **no unpaired electrons**.
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 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.

8.79 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.

8.80 Substances are paramagnetic if they have unpaired electrons. This problem is more challenging if you have difficulty picturing the orbital diagram from the electron configuration. Obviously an odd number of electrons will necessitate that at least one electron is unpaired, so odd electron species are paramagnetic. However, a substance with an even number of electrons can also be paramagnetic, because even numbers of electrons do not guarantee all electrons are paired.
a) Ga ($Z = 31$) = [Ar]$4s^23d^{10}4p^1$. The s and d sublevels are filled, so all electrons are paired. The lone p electron is unpaired, so this element is **paramagnetic**.

b) Si ($Z = 14$) = [Ne]$3s^2 3p^2$. This element is **paramagnetic** with two unpaired electrons because the *p* subshell looks like this, not this.





c) Be ($Z = 4$) = [He]$2s^2$, so it is **not paramagnetic**.

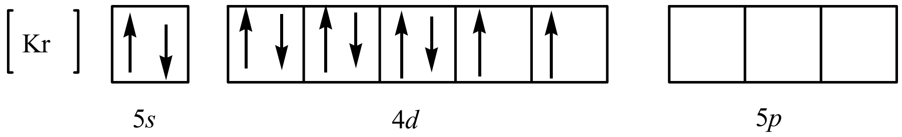
d) Te ($Z = 52$) = [Kr]$5s^2 4d^{10} 5p^4$ is **paramagnetic** with two unpaired electrons.

8.81 a) Ti^{2+}: [Ar]$3d^2$ **paramagnetic**
b) Zn^{2+}: [Ar]$3d^{10}$ **diamagnetic**
c) Ca^{2+}: [Ar] **diamagnetic**
d) Sn^{2+}: [Kr]$5s^2 4d^{10}$ **diamagnetic**

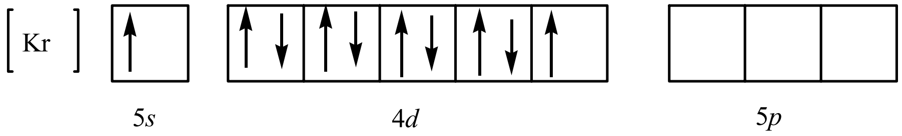
8.82 a) V^{3+}, [Ar]$3d^2$ Transition metals first lose the 4*s* electrons in forming ions, so to form the +3 ion a vanadium atom of configuration [Ar]$4s^2 3d^3$ loses two 4*s* electrons and one 3*d* electron. **Paramagnetic**
b) Cd^{2+}, [Kr]$4d^{10}$ Cadmium atoms lose two electrons from the 4*s* orbital to form the +2 ion. **Diamagnetic**
c) Co^{3+}, [Ar]$3d^7$ Cobalt atoms lose two 4*s* electrons and one 3*d* electron to form the +3 ion. **Paramagnetic**
d) Ag^+, [Kr]$4d^{10}$ Silver atoms lose the one electron in the 5*s* orbital to form the +1 ion. **Diamagnetic**

8.83 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**

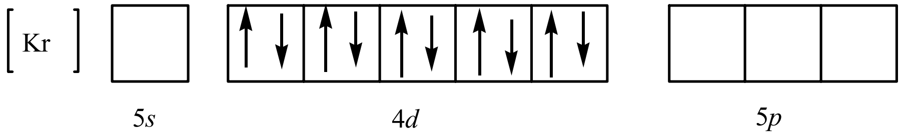
8.84 For palladium to be diamagnetic, all of its electrons must be paired. You might first write the condensed electron configuration for Pd as [Kr]$5s^2 4d^8$. However, the partial orbital diagram is not consistent with diamagnetism.



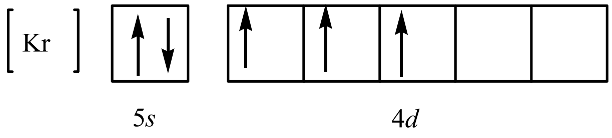
Promoting an *s* electron into the *d* sublevel (as in (c)) still leaves two electrons unpaired.



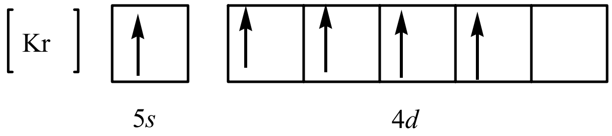
The only configuration that supports diamagnetism is **(b) [Kr]$4d^{10}$**.



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



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



8.86 The size of ions increases down a group. For ions that are isoelectronic (have the same electron configuration) size decreases with increasing atomic number.

a) Increasing size: $\mathbf{Li^+ < Na^+ < K^+}$, size increases down group 1A.

b) Increasing size: $\mathbf{Rb^+ < Br^- < Se^{2-}}$, these three ions are isoelectronic with the same electron configuration as krypton. Size decreases with atomic number in an isoelectronic series.

c) Increasing size: $\mathbf{F^- < O^{2-} < N^{3-}}$, the three ions are isoelectronic with an electron configuration identical to neon. Size decreases with atomic number in an isoelectronic series.

8.87 a) $\mathbf{Se^{2-} > S^{2-} > O^{2-}}$, size increases down a group.

b) $\mathbf{Te^{2-} > I^- > Cs^+}$, size decreases with atomic number in an isoelectronic series.

c) $\mathbf{Cs^+ > Ba^{2+} > Sr^{2+}}$, both reasons as in (a) and (b).

8.88 Because of its $1s^1$ electron configuration, it is generally placed above lithium in the periodic table. It may form a positive ion typical of alkali metals but because of its larger ionization energy, its tendency to form a positive ion is much less than other alkali metals. Hydrogen is also placed above the halogen group 7A (17) because it may pick up an electron to form the hydride ion, H^-, which has a noble gas configuration.

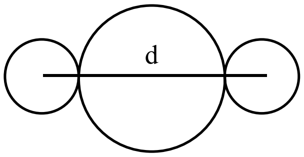
8.89
a) **oxygen**
b) **cesium**
c) **aluminum**
d) **carbon**
e) **rubidium**
f) **bismuth**
g) **thallium**
h) **krypton**
i) **silicon**
j) **ruthenium**
k) **vanadium**
l) **indium**
m) **scandium**
n) **manganese**
o) **lutetium**
p) **sulfur**
q) **strontium**
r) **arsenic**

8.90 Ce: $[\mathrm{Xe}]6s^2 4f^1 5d^1$ Eu: $[\mathrm{Xe}]6s^2 4f^7$

$\mathrm{Ce^{4+}}$: $[\mathrm{Xe}]$ $\mathrm{Eu^{2+}}$: $[\mathrm{Xe}]4f^7$

$\mathrm{Ce^{4+}}$ has a noble-gas configuration and $\mathrm{Eu^{2+}}$ has a half-filled *f* subshell.

8.91



$$d = 566 \text{ pm}$$

$$r_{\mathrm{Na^+}} = (56.4\% / 100\%)\, r_{\mathrm{Cl^-}}$$

$$r_{\mathrm{Na^+}} = 0.564\, r_{\mathrm{Cl^-}}$$

$$d = 2\left(r_{\mathrm{Cl^-}} + r_{\mathrm{Na^+}}\right)$$

$$566 \text{ pm} = 2\left(r_{\mathrm{Cl^-}} + 0.564 r_{\mathrm{Cl^-}}\right)$$

$$566 \text{ pm} = 3.128\, r_{\mathrm{Cl^-}}$$

$$r_{\mathrm{Cl^-}} = 180.946 = \mathbf{181\ pm}$$

$$r_{\mathrm{Na^+}} = 0.564(180.946 \text{ pm}) = 102.0535 = \mathbf{102\ pm}$$

8.92 <u>Plan:</u> Determine the oxidation state of the nonmetal. 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.

<u>Solution:</u>

a) hypochlorous acid = HClO has $\mathrm{Cl^+}$ so the oxide is $\mathbf{Cl_2O}$ = **dichlorine oxide** or **dichlorine monoxide**

b) chlorous acid = $\mathrm{HClO_2}$ has $\mathrm{Cl^{3+}}$ so the oxide is $\mathbf{Cl_2O_3}$ = **dichlorine trioxide**

c) chloric acid = $\mathrm{HClO_3}$ has $\mathrm{Cl^{5+}}$ so the oxide is $\mathbf{Cl_2O_5}$ = **dichlorine pentaoxide**

d) perchloric acid = $\mathrm{HClO_4}$ has $\mathrm{Cl^{7+}}$ so the oxide is $\mathbf{Cl_2O_7}$ = **dichlorine heptaoxide**

e) sulfuric acid = $\mathrm{H_2SO_4}$ has $\mathrm{S^{6+}}$ so the oxide is $\mathbf{SO_3}$ = **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 **N_2O_3 = dinitrogen trioxide**
i) carbonic acid = H_2CO_3 has C^{4+} so the oxide is **CO_2 = carbon dioxide**
j) phosphoric acid = H_3PO_4 has P^{5+} so the oxide is **P_2O_5 = diphosphorus pentaoxide**

8.93 a) The distance, *d*, between the electron and the nucleus would be smaller for He than for H, so IE would be larger.
b) You would expect that IE_1 (He) to be approximately 2 IE_1 (H), since $Z_{He} = 2\ Z_H$. However, since the **effective** nuclear charge of He is < 2 due to shielding, IE_1 (He) < 2 IE_1(H).

8.94
$$\lambda = hc / \Delta E = \frac{\left(6.626 \text{ x } 10^{-34} \text{ J} \bullet \text{s}\right)\left(3.00 \text{ x } 10^{8} \text{ m/s}\right)}{\left(2.7 \text{ eV}\right)\left(\frac{1.602 \text{ x } 10^{-19} \text{ J}}{1 \text{ eV}}\right)} = 4.59564 \text{ x } 10^{-7} = 4.6 \text{ x } 10^{-7} \text{ m}$$

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.

8.95 a) A chemically unreactive Period 4 element would be Kr. Both the Sr^{2+} ion and Br^- ion are isoelectronic with Kr. Their combination results in **$SrBr_2$, 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 **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**.

8.96 a) The small size of the Li^+ ion means that it has a high charge density, similar to that for a Mg^{2+} ion. This ability for both to deform nearby polarizable electron clouds results in similar chemical properties.
b) The large energy needed to remove two electrons from the very small Be atom results in the inability to form discrete Be^{2+} ions. Thus, its compounds are partially covalent in character, which means that its oxide is amphoteric like aluminum and unlike the basic oxides of the other Group 2A(2) metals.

8.97 Both the ionization energies and the electron affinities of the elements are needed.
a) F: ionization energy = 1681 kJ/mol electron affinity = -328 kJ/mol
$F(g) \rightarrow F^+(g) + e^-$ ΔH = 1681 kJ/mol
$F(g) + e^- \rightarrow F^-(g)$ ΔH = -328 kJ/mol
Reverse the electron affinity reaction to give: $F^-(g) \rightarrow F(g) + e^- \Delta H$ = +328 kJ/mol
Summing the ionization energy reaction with the reversed electron affinity reaction (Hess's Law):
$F^-(g) \rightarrow F^+(g) + 2\ e^-$ ΔH = +328 kJ/mol + 1681 kJ/mol = **2009 kJ/mol**
b) Na: ionization energy = 496 kJ/mol electron affinity = -52.9 kJ/mol
$Na(g) \rightarrow Na^+(g) + e^-$ ΔH = 496 kJ/mol
$Na(g) + e^- \rightarrow Na^-(g)$ ΔH = -52.9 kJ/mol
Reverse the ionization reaction to give: $Na^+(g) + e^- \rightarrow Na(g)$ ΔH = -496 kJ/mol
Summing the electron affinity reaction with the reversed ionization reaction (Hess's Law):
$Na^+(g) + 2\ e^- \rightarrow Na^-(g)$ ΔH = -496 kJ/mol + -52.9 kJ/mol = -548.9 = **-549 kJ/mol**

8.98 Plan: List the respective ionization energies for the two elements. Look for the places where there are breaks in the steady increase in the values.

Solution: Ionization energies are reported in MJ/mol.

Carbon $IE_1 = 1.09$ $IE_2 = 2.35$ $IE_3 = 4.62$ $IE_4 = 6.22$ $IE_5 = 37.83$

Oxygen $IE_1 = 1.31$ $IE_2 = 3.39$ $IE_3 = 5.30$ $IE_4 = 7.47$ $IE_5 = 10.98$

Oxygen shows a gradual increase in ionization energies as electrons are being removed from a shell. Carbon begins a trend similar to that of oxygen. However, there is a big jump between IE_4 and IE_5. This big change occurs because the removal of all the valence electrons (at IE_4). The value for IE_5 is much larger since this corresponds to the removal of the first core electron.

8.99 Plan: Determine the electron configuration for iron, and then begin removing one electron at a time. Filled subshells give diamagnetic contributions. Any partially filled subshells give a paramagnetic contribution. The more unpaired electrons, the greater the attraction to a magnetic field.

Solution:

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

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

8.100 a) Ga $[Ar]4s^2\ 3d^{10}\ 4p^1$ Ge $[Ar]4s^2\ 3d^{10}\ 4p^2$

In each case, an electron is removed from a 4*p* orbital. Because the Zeff of Ge is greater, IE, of Ge will be higher.

b) Ga^+ $[Ar]4s^2\ 3d^{10}$ Ge^+ $[Ar]4s^2\ 3d^{10}\ 4p^1$

IE_2 would be higher for Ga^+, because an electron must be removed from a 4*s* orbital. For Ge^+ the electron is removed from the 4*p* orbital.

c) Ga^{2+} $[Ar]4s^1\ 3d^{10}$ Ge^{2+} $[Ar]4s^2\ 3d^{10}$

In each case, an electron is removed from a 4*s* orbital. Because the Z_{eff} of Ge^{2+} is greater, IE_3 of Ge^{2+} will be higher.

d) Ga^{3+} $[Ar]3d^{10}$ Ge^{2+} $[Ar]4s^1\ 3d^{10}$

IE_4 would be higher for Ga^{3+}, because an electron must be removed from a 3*d* orbital. For Ge^{3+} the electron is removed from the 4*s* orbital.

8.101 a) The energy required would be 520 kJ/mol (Li)

$$\lambda = hc/E = \frac{\left(6.626 \text{ x } 10^{-34} \text{ J} \bullet \text{s}\right)\left(3.00 \text{ x } 10^{8} \text{ m/s}\right)}{\left(520 \frac{\text{kJ}}{\text{mol}}\right)\left(\frac{\text{mol}}{6.022 \text{ x } 10^{23}}\right)}\left(\frac{1 \text{ kJ}}{10^3 \text{ J}}\right) = 2.3020 \text{ x } 10^{-7} = \mathbf{2.3 \text{ x } 10^{-7} \text{ m}}$$

b) The energy required would be 899 kJ/mol (Be)

$$\lambda = hc/E = \frac{\left(6.626 \text{ x } 10^{-34} \text{ J} \bullet \text{s}\right)\left(3.00 \text{ x } 10^{8} \text{ m/s}\right)}{\left(899 \frac{\text{kJ}}{\text{mol}}\right)\left(\frac{\text{mol}}{6.022 \text{ x } 10^{23}}\right)}\left(\frac{1 \text{ kJ}}{10^3 \text{ J}}\right) = 1.3315 \text{ x } 10^{-7} = \mathbf{1.33 \text{ x } 10^{-7} \text{ m}}$$

c) This includes all elements with lower ionization energies. These are **B, Al, Ga, In, Tl, Si, Ge, Sn, Pb, Sb, Bi, Te,** and **Po**.

d) Both absorb in the ultraviolet.

8.102 All of the odd numbered elements are paramagnetic. Some of the even numbered elements are paramagnetic when electrons are placed in degenerate orbitals. The even numbered elements are He, Be, C, O, Ne, Mg, Si, S, and Ar. The paramagnetic members of this group are: C, O, Si, and S.

8.103 We know that the maximum number of orbitals per energy level, *n*, is n^2. The $n = 4$ level contains 16 orbitals (s=1, p=3, d=5, f=7; total = 16). The $n = 5$ level must contain 25 orbitals so an additional sublevel is needed to be consistent with the n^2 rule. The additional sublevel, called "*g*," contains 9 orbitals (s=1, p=3, d=5, f=7, g=9; total = 25).

$_{126}$E: [Rn]$7s^2 5f^{14} 6d^{10} 7p^6 8s^2 5g^6$

$n = 8$, $l = 0 \rightarrow 8s$; the 8*s* subshell has one orbital.

8.104 Plan: Convert the wavelength to the energy using $E = \text{hc} / \lambda$.

Solution:

Indium:

$$E = \text{hc} / \lambda = \frac{\left(6.626 \text{ x } 10^{-34} \text{ J} \bullet \text{s}\right)\left(2.99792 \text{ x } 10^{8} \text{ m/s}\right)}{\left(451.1 \text{ nm}\right)\left(\dfrac{10^{-9} \text{ m}}{1 \text{ nm}}\right)} = 4.4035065 \text{ x } 10^{-19} = \mathbf{4.404 \text{ x } 10^{-19} \text{ J}}$$

Thallium:

$$E = \text{hc} / \lambda = \frac{\left(6.626 \text{ x } 10^{-34} \text{ J} \bullet \text{s}\right)\left(2.99792 \text{ x } 10^{8} \text{ m/s}\right)}{\left(535.0 \text{ nm}\right)\left(\dfrac{10^{-9} \text{ m}}{1 \text{ nm}}\right)} = 3.7129379 \text{ x } 10^{-19} = \mathbf{3.713 \text{ x } 10^{-19} \text{ J}}$$

8.105 a) $FeO(s) + 2\ HCl(aq) \rightarrow FeCl_2(aq) + H_2O(l)$

b) $Cr_2O_3(s) + 6\ HNO_3(aq) \rightarrow 2\ Cr(NO_3)_2(aq) + 3\ H_2O(l)$

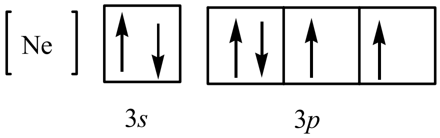
c) $Co_2O_3(s) + 3\ H_2SO_4(aq) \rightarrow Co_2(SO_4)_3(aq) + 3\ H_2O(l)$

d) $Cu_2O(s) + 2\ HBr(aq) \rightarrow 2\ CuBr(s) + H_2O(l)$

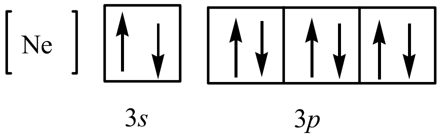
e) $3\ MnO(s) + 2\ H_3PO_4(aq) \rightarrow Mn_3(PO_4)_2(s) + 3\ H_2O(l)$

8.106 The sudden jump in energy between IE_6 and IE_7 makes this an element with 6 valence electrons.

E: [Ne]$3s^2 3p^4$



E^{2-}: [Ne]$3s^2 3p^6$



8.107 a) The first student chooses the lower bunk in the bedroom on the first floor.

b) There are seven students on the top bunk when the seventeenth student chooses.

c) The twenty-first student chooses a bottom bunk on the third floor in the largest room.

d) There are fourteen students in bottom bunks when the twenty-fifth student chooses.

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