8 Periodic Properties of the Elements

Review Questions

8.1 A periodic property is one that is predictable based on the element's position within the periodic table.

8.2 The relative size of the sodium and potassium ions is important to nerve signal transmission. The pumps and channels within cell membranes are so sensitive that they can distinguish between the sizes of these two ions and selectively allow only one or the other to pass. The movement of ions is the basis for the transmission of nerve signals in the brain and throughout the body.

8.3 The first attempt to organize the elements according to similarities in their properties was made by the German chemist Johann Dobereiner. He grouped elements into triads; three elements with similar properties. A more complex approach was attempted by the English chemist John Newlands. He organized elements into octaves, analogous to musical notes. When arranged this way, the properties of every eighth element were similar.

8.4 The modern periodic table is credited primarily to the Russian chemist Dmitri Mendeleev. Mendeleev's table is based on the periodic law, which states that when elements are arranged in order of increasing mass, their properties recur periodically. Mendeleev arranged the elements in a table in which mass increased from left to right and elements with similar properties fell in the same columns.

8.5 Meyer proposed an organization of the known elements based on some periodic properties. Moseley listed elements according to the atomic number rather than atomic mass. This resolved the problems in Mendeleev's table where an increase in atomic mass did not correlate with similar properties.

8.6 The periodic law was based on the observations that the properties of elements recur and certain elements have similar properties. The theory that explains the existence of the periodic law is quantum-mechanical theory.

8.7 Electron spin is a fundamental property of electrons. It is more correctly expressed as saying the electron has inherent angular momentum. The value m_s is the spin quantum number. An electron with $m_s = +1/2$ has a spin opposite of an electron with $m_s = -1/2$.

8.8 In the Stern–Gerlach experiment a beam of silver atoms is split into two separate trajectories by a magnet. The spin of the electrons within the atoms creates a tiny magnetic field that interacts with the external field. One spin orientation causes the deflection of the beam in one direction, while the other orientation causes a deflection in the opposite direction. Since there were only two trajectories, the spin of the electron is quantized, that is, it can have one of two values and nothing in between.

8.9 An electron configuration shows the particular orbitals that are occupied by electrons in an atom. Some examples are H = $1s^1$, He = $1s^2$, and Li = $1s^22s^1$.

8.41 (a) The reactions of the alkali metals with halogens result in the formation of metal halides.
$2\ M(s) + X_2 \rightarrow 2\ MX(s)$

(b) Alkali metals react with water to form the dissolved alkali metal ion, the hydroxide ion, and hydrogen gas.
$2\ M(s) + 2\ H_2O(l) \rightarrow 2\ M^+(aq) + 2\ OH^-(aq) + H_2(g)$

8.42 All of the halogens are powerful oxidizing agents.

(a) The halogens react with metals to form metal halides.
$2\ M(s) + n\ X_2 \rightarrow 2\ MX_n(s)$

(b) The halogens react with hydrogen to form hydrogen halides.
$H_2(g) + X_2 \rightarrow 2\ HX(g)$

(c) The halogens react with each other to form interhalogen compounds.
e.g. $Br_2(l) + F_2(g) \rightarrow 2\ BrF(g)$

Problems by Topic

Electron Configurations

8.43 (a) Si Silicon has 14 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbital, two into the 3*s* orbital, and two into the 3*p* orbital. $1s^22s^22p^63s^23p^2$

(b) O Oxygen has 8 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, and four into the 2*p* orbital. $1s^22s^22p^4$

(c) K Potassium has 19 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbital, two into the 3*s* orbital, six into the 3*p* orbital, and one into the 4s orbital. $1s^22s^22p^63s^23p^64s^1$

(d) Ne Neon has 10 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, and six into the 2*p* orbital. $1s^22s^22p^6$

8.44 (a) C Carbon has 6 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, and two into the 2*p* orbital. $1s^22s^22p^2$

(b) P Phosphorus has 15 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbital, two into the 3s orbital, and three into the 3*p* orbital. $1s^22s^22p^63s^23p^3$

(c) Ar Argon has 18 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbital, two into the 3*s* orbital, and six into the 3*p* orbital. $1s^22s^22p^63s^23p^6$

(d) Na Sodium has 11 electrons. Distribute two of these into the 1*s* orbital, two into the 2*s* orbital, six into the 2*p* orbital, and one into the 3*s* orbital. $1s^22s^22p^63s^1$

8.45 (a) N Nitrogen has 7 electrons and has the electron configuration $1s^22s^22p^3$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 7 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that nitrogen has 3 unpaired electrons.

$\boxed{\downarrow\uparrow}$	$\boxed{\downarrow\uparrow}$	$\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$
1*s*	2*s*	2*p*

(b) F Fluorine has 9 electrons and has the electron configuration $1s^22s^22p^5$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 9 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that fluorine has 1 unpaired electron.

[↓↑] [↓↑] [↓↑|↓↑|↑]
1s 2s 2p

(c) Mg Magnesium has 12 electrons and has the electron configuration $1s^22s^22p^63s^2$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 12 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that magnesium has no unpaired electrons.

[↓↑] [↓↑] [↓↑|↓↑|↓↑] [↓↑]
1s 2s 2p 3s

(d) Al Aluminum has 13 electrons and has the electron configuration $1s^22s^22p^63s^23p^1$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 13 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that aluminum has 1 unpaired electron.

[↓↑] [↓↑] [↓↑|↓↑|↓↑] [↓↑] [↑| |]
1s 2s 2p 3s 3p

8.46 (a) S Sulfur has 16 electrons and has the electron configuration $1s^22s^22p^63s^23p^4$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 16 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that sulfur has 2 unpaired electrons.

[↓↑] [↓↑] [↓↑|↓↑|↓↑] [↓↑] [↓↑|↑|↑]
1s 2s 2p 3s 3p

(b) Ca Calcium has 20 electrons and has the electron configuration $1s^22s^22p^63s^23p^64s^2$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 20 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that nitrogen has no unpaired electrons.

[↓↑] [↓↑] [↓↑|↓↑|↓↑] [↓↑] [↓↑|↓↑|↓↑] [↓↑]
1s 2s 2p 3s 3p 4s

(c) Ne Neon has 10 electrons and has the electron configuration $1s^22s^22p^6$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 10 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that neon has no unpaired electrons.

[↓↑] [↓↑] [↓↑|↓↑|↓↑]
1s 2s 2p

(d) He Helium has 2 electrons and has the electron configuration $1s^2$. Draw a box for each orbital, putting the lowest energy orbital (1s) on the far left and proceeding to orbitals of higher energy to the right. Distribute the 2 electrons into the boxes representing the orbitals, allowing a maximum of two electrons per orbital and remembering Hund's rule. You can see from the diagram that helium has no unpaired electrons.

[↓↑]
1s

8.47 (a) P The atomic number of P is 15. The noble gas that precedes P in the periodic table is neon, so the inner electron configuration is [Ne]. Obtain the outer electron configuration by tracing the elements between Ne and P and assigning electrons to the appropriate orbitals. Begin with [Ne]. Because P is in row 3, add two 3*s* electrons. Next add three 3*p* electrons as you trace across the *p* block to P, which is in the third column of the p block.

P $[Ne]3s^23p^3$

(b) Ge The atomic number of Ge is 32. The noble gas that precedes Ge in the periodic table is argon, so the inner electron configuration is [Ar]. Obtain the outer electron configuration by tracing the elements between Ar and Ge and assigning electrons to the appropriate orbitals. Begin with [Ar]. Because Ge is in row 4, add two 4*s* electrons. Next, add ten 3*d* electrons as you trace across the *d* block. Finally add two 4*p* electrons as you trace across the *p* block to Ge, which is in the second column of the *p* block.

Ge $[Ar]4s^23d^{10}4p^2$

(c) Zr The atomic number of Zr is 40. The noble gas that precedes Zr in the periodic table is krypton, so the inner electron configuration is [Kr]. Obtain the outer electron configuration by tracing the elements between Kr and Zr and assigning electrons to the appropriate orbitals. Begin with [Kr]. Because Zr is in row 5, add two 5*s* electrons. Next, add two 4*d* electrons as you trace across the *d* block to Zr, which is in the second column.

Zr $[Kr]5s^24d^2$

(d) I The atomic number of I is 53. The noble gas that precedes I in the periodic table is krypton, so the inner electron configuration is [Kr]. Obtain the outer electron configuration by tracing the elements between Kr and I and assigning electrons to the appropriate orbitals. Begin with [Kr]. Because I is in row 5, add two 5*s* electrons. Next, add ten 4*d* electrons as you trace across the d block. Finally add five 5*p* electrons as you trace across the *p* block to I which is in the fifth column of the *p* block.

I $[Kr]5s^24d^{10}5p^5$

8.48 (a) $[Ar]\ 4s^23d^{10}4p^6$ To determine the element corresponding to the electron configuration, begin with Ar then trace across the 4*s* block, the 3*d* block, and then the 4*p* block until you come to the sixth column. The element is Kr.

(b) $[Ar]\ 4s^23d^2$ To determine the element corresponding to the electron configuration, begin with Ar then trace across the 4*s* block, and then 3*d* block until you come to the second column. The element is Ti.

(c) $[Kr]\ 5s^24d^{10}5p^2$ To determine the element corresponding to the electron configuration, begin with Kr then trace across the 5*s* block, the 4*d* block, and then the 5*p* block until you come to the second column. The element is Sn.

(d) $[Kr]\ 5s^2$ To determine the element corresponding to the electron configuration, begin with Kr then trace across the 5*s* block to the second column. The element is Sr.

8.49 (a) Li is in period 2, and the first column in the *s* block so Li has one 2*s* electron.

(b) Cu is in period 4, and the ninth column in the *d* block (n – 1) so Cu should have nine 3*d* electrons, however, it is one of our exceptions, so it has ten 3*d* electrons.

(c) Br is in period 4, and the fifth column of the *p* block, so Br has five 4*p* electrons.

(d) Zr is in period 5, and the second column of the *d* block (n – 1), so Zr has two 4*d* electrons.

8.50 (a) Mg is in period 3, and the second column of the *s* block, so Mg has two 3*s* electrons.

(b) Cr is in period 4, and the fourth column of the *d* block (n – 1), so Cr should have four 3*d* electrons, however, Cr is one of our exceptions, so it has five 3*d* electrons.

(c) Y is in period 5, and the first column of the *d* block (n – 1), so Y has one 4*d* electron.

(d) Pb is in period 6, and the second column of the *p* block, so Pb has two 6*p* electrons.

8.51 (a) In period 4, an element with five valence electrons could be V or As.

(b) In period 4, an element with four 4*p* electrons would be in the fourth column of the *p* block, and is Se.

(c) In period 4, an element with three $3d$ electrons would be in the third column of the d block (n – 1) and is V.

(d) In period 4, an element with a complete outer shell would be in the sixth column of the p block and is Kr.

8.52 (a) In period 3, an element with three valence electrons would be in the first column of the p block and is Al.

(b) In period 3, an element with four $3p$ electrons would be in the fourth column of the p block and is S.

(c) In period 3, an element with six $3p$ electrons would be in the sixth column of the p block and is Ar.

(d) In period 3, an element with two $3s$ electrons and no $3p$ electrons would be in the second column of the s block and is Mg.

Valence Electrons and Simple Chemical Behavior form the Periodic Table

8.53 (a) Ba is in column 2A, so it has two valence electrons.

(b) Cs is in column 1A, so it has one valence electron.

(c) Ni is in column 8 of the d block, so it has 10 valence electrons (8 from the d block and 2 from the s block).

(d) S is in column 6A, so it has six valence electrons.

8.54 (a) Al is in column 3A, so it has three valence electrons. Al is a metal and will tend to lose the three valence electrons to achieve the noble gas configuration of Ne.

(b) Sn is in column 4A, so it has four valence electrons. Sn is a metal and will tend to lose the valence electrons to obtain a completely filled $n = 3$ level.

(c) Br is in column 7A, so it has seven valence electrons. Br is a nonmetal and will tend to gain an electron to achieve the noble gas configuration of Kr.

(d) Se is in column 6A, so it has six valence electrons. Se is a nonmetal and will tend to gain electrons to achieve the noble gas configuration of Kr.

8.55 (a) The outer electron configuration ns^2 would belong to a reactive metal in the alkaline earth family.

(b) The outer electron configuration ns^2np^6 would belong to an unreactive nonmetal in the noble gas family.

(c) The outer electron configuration ns^2np^5 would belong to a reactive nonmetal in the halogen family.

(d) The outer electron configuration ns^2np^2 would belong to an element in the carbon family. If $n = 2$, the element is a nonmetal, if $n = 3$ or 4, the element is a metalloid, and if $n = 5$ or 6, the element is a metal.

8.56 (a) The outer electron configuration ns^2 would belong to a metal in the alkaline earth family for period n = 2 and greater. He is a noble gas with a 1s2 electron configuration.

(b) The outer electron configuration ns^2np^6 would belong to a nonmetal in the noble gas family.

(c) The outer electron configuration ns^2np^5 would belong to a nonmetal in the halogen family.

(d) The outer electron configuration ns^2np^2 would belong to an element in the carbon family. If $n = 2$, the element is a nonmetal, if $n = 3$ or 4, the element is a metalloid, and if $n = 5$ or 6, the element is a metal.

Effective Nuclear Charge and Atomic Radius

8.57 The valence electrons in nitrogen would experience a greater effective nuclear charge. Be has four protons and N has seven protons. Both atoms have two core electrons that predominately contribute to the shielding, while the valence electrons will contribute a slight shielding effect. So, Be has an effective nuclear charge of slightly more than 2+ and N has an effective nuclear charge of slightly more than 5+.

8.58 S(16) = [Ne]$3s^23p^4$ Mg(12) = [Ne]$3s^2$ Al(13) = [Ne]$3s^23p^1$ Si(14) = [Ne]$3s^23p^2$
All four atoms have the same number of core electrons that contribute to shielding. So, the effective nuclear charge will decrease with decreasing number of protons. S > Si > Al > Mg

8.59 (a) K(19) [Ar]$4s^1$ Z_{eff} = Z – core electrons = 19 – 18 = 1+

(b) Ca(20) [Ar]$4s^2$ Z_{eff} = Z – core electrons = 20 – 18 = 2+

(c) O(8) [He]$2s^22p^4$ Z_{eff} = Z – core electrons = 8 – 2 = 6+

(d) C(6) [He]$2s^22p^2$ Z_{eff} = Z – core electrons = 6 – 2 = 4+

8.60 B has an electron configuration of $1s^22s^22p^1$. To estimate the effective nuclear charge experienced by the outer electrons we need to distinguish between two different types of shielding: (1) the shielding of the outermost electrons by the core electrons and (2) the shielding of the outermost electrons by each other. The three outermost electrons in boron experience the 5+ charge of the nucleus through the shield of the two 1*s* core electrons. We can estimate that the shielding experienced by any one of the outermost electrons due to the core electrons is nearly 2. For the 2*s* electrons the shielding due to the other 2*s* electron is nearly zero. For the 2*p* electron however, we would expect that the 2s electrons would contribute some shielding because although the 2*p* orbital penetrates the 2*s* orbital to some degree most of the 2*p* orbital lies outside the 2*s* orbital. So the effective nuclear charge would be slightly greater than 3+ and the effective nuclear charge felt by the 2*s* electrons would be greater than the effective nuclear charge felt by the 2*p* electrons.

8.61 (a) Al or In — In atoms are larger than Al atoms because as you trace the path between Al and In on the periodic table you move down a column. Atomic size increases as you move down a column because the outermost electrons occupy orbitals with a higher principal quantum number that are therefore larger, resulting in a larger atom.

(b) Si or N — Si atoms are larger than N atoms because as you trace the path between N and Si on the periodic table you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.

(c) P or Pb — Pb atoms are larger than P atoms because as you trace the path between P and Pb on the periodic table you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.

(d) C or F — C atoms are larger than F atoms because as you trace the path between C and F on the periodic table you move to the right within the same period. As you move to the right across a period, the effective nuclear charge experienced by the outermost electrons increase, which results in a smaller size.

8.62 (a) Sn or Si — Sn atoms are larger than Si atoms because as you trace the path between Si and Sn on the periodic table you move down a column. Atomic size increases as you move down a column because the outermost electrons occupy orbitals with a higher principal quantum number that are therefore larger, resulting in a larger atom.

(b) Br or Ga — Ga atoms are larger than Br atoms because as you trace the path between Ga and Br on the periodic table you move to the right within the same period. As you move to the right across a period, the effective nuclear charge experienced by the outermost electrons increases, which results in a smaller size.

(c) Sn or Bi — Based on periodic trends alone, you cannot tell which atom is larger because as you trace the path between Sn and Bi you go to the right across a period (atomic size decreases) and then down a column (atomic size increases). These effects tend to oppose each other, and it is not easy to tell which will predominate.

(d) Se or Sn — Sn atoms are larger than Se atoms because as you trace the path between Se and Sn on the periodic table you move down a column (atomic size increases) and then to the left across a period (atomic size increases). These effects add together for an overall increase.

8.63 Ca, Rb, S, Si, Ge, F F is above and to the right of the other elements, so we start with F as the smallest atom. As you trace a path from F to S you move to the left (size increases) and down (size increases), next you move left from S to Si (size increases), then down to Ge (size increases), next move to the left to Ca (size increases), and then to the left and down to Rb (size increases). So, in order of increasing atomic radii F < S < Si < Ge < Ca < Rb.

8.64 Cs, Sb, S, Pb, Se Cs is below and to the left of the other elements, so we start with Cs as the largest atom. As you trace a path from Cs to Pb you move to the right in the same period (size decreases), next, going from Pb to Sb you move up a column and then to the right (size decreases), from Sb to Se you move up the column and then to the right (size decreases), and finally from Se to S you move up the column (size decreases). So, in order of decreasing radii Cs > Pb > Sb > Se > S.

Ionic Electron Configurations, Ionic Radii, Magnetic Properties, and Ionization Energy

8.65 (a) O^{2-} Begin by writing the electron configuration of the neutral atom.
O $1s^22s^22p^4$
Since this ion has a 2 – charge, add two electrons to write the electron configuration of the ion.
O^{2-} $1s^22s^22p^6$ This is isoelectronic with Ar.

(b) Br^- Begin by writing the electron configuration of the neutral atom.
Br $[Ar]4s^23d^{10}4p^5$
Since this ion has a 1 – charge, add one electron to write the electron configuration of the ion.
Br^- $[Ar]4s^23d^{10}4p^6$ This is isoelectronic with Kr.

(c) Sr^{2+} Begin by writing the electron configuration of the neutral atom.
Sr $[Kr]5s^2$
Since this ion has a 2+ charge, remove two electrons to write the electron configuration of the ion.
Sr^{2+} [Kr]

(d) Co^{3+} Begin by writing the electron configuration of the neutral atom.
Co $[Ar]4s^23d^7$
Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
Co^{3+} $[Ar]4s^03d^6$

(e) Cu^{2+} Begin by writing the electron configuration of the neutral atom. Remember, Cu is one of our exceptions.
Cu $[Ar]4s^13d^{10}$
Since this ion has a 2+ charge, remove two electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
Cu^{2+} $[Ar]4s^03d^9$

8.66 (a) Cl^- Begin by writing the electron configuration of the neutral atom.
Cl $[Ne]3s^23p^5$
Since this ion has a 1 – charge, add one electron to write the electron configuration of the ion.
Cl^- $[Ne]3s^23p^6$ This is isoelectronic with Ar.

(b) P^{3-} Begin by writing the electron configuration of the neutral atom.
P $[Ne]3s^23p^3$
Since this ion has a 3 – charge, add three electrons to write the electron configuration of the ion.
P^{3-} $[Ne]3s^23p^6$ This is isoelectronic with Ar.

(c) K^+ Begin by writing the electron configuration of the neutral atom.
K $[Ar]4s^1$
Since this ion has a 1+ charge, remove one electron to write the electron configuration of the ion.
K^+ [Ar]

(d) Mo^{3+} Begin by writing the electron configuration of the neutral atom. Remember, Mo is one of our exceptions.
Mo $[Kr]5s^14d^5$
Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 5s orbital before removing electrons from the 4*d* orbitals.
Mo^{3+} $[Kr]5s^04d^3$

(e) V^{3+} Begin by writing the electron configuration of the neutral atom.
V $[Ar]4s^23d^3$
Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
V^{3+} $[Ar]4s^03d^2$

8.67 (a) V^{5+} Begin by writing the electron configuration of the neutral atom.
V $[Ar]4s^23d^3$
Since this ion has a 5+ charge, remove five electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
V^{5+} $[Ar]4s^03d^0 = [Ne]3s^23p^6$
[Ne] |↓↑| |↓↑|↓↑|↓↑|
3*s* 3*p*
V^{5+} is diamagnetic.

(b) Cr^{3+} Begin by writing the electron configuration of the neutral atom. Remember, Cr is one of our exceptions.
Cr $[Ar]4s^13d^5$
Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
Cr^{3+} $[Ar]4s^03d^3$
[Ar] | | |↑|↑|↑| | |
4*s* 3*d*
Cr^{3+} is paramagnetic.

(c) Ni^{2+} Begin by writing the electron configuration of the neutral atom.
Ni $[Ar]4s^23d^8$
Since this ion has a 2+ charge, remove two electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4*s* orbital before removing electrons from the 3*d* orbitals.
Ni^{2+} $[Ar]4s^03d^8$
[Ar] | | |↓↑|↓↑|↓↑|↑|↑|
4*s* 3*d*
Ni^{2+} is paramagnetic.

(d) Fe^{3+} Begin by writing the electron configuration of the neutral atom.
Fe $[Ar]4s^23d^6$
Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 4s orbital before removing electrons from the 3d orbitals.

Fe^{3+} [Ar]$4s^0 3d^5$

[Ar] [ ] [↑][↑][↑][↑][↑]

4s 3d

Fe^{3+} is paramagnetic.

8.68 (a) Cd^{2+} Begin by writing the electron configuration of the neutral atom.

Cd [Kr]$5s^2 4d^{10}$

Since this ion has a 2+ charge, remove two electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 5*s* orbital before removing electrons from the 4*d* orbitals.

Cd^{2+} [Kr]$5s^0 4d^{10}$

[Kr] [ ] [↓↑][↓↑][↓↑][↓↑][↓↑]

5*s* 4*d*

Cd^{2+} is diamagnetic.

(b) Au^+ Begin by writing the electron configuration of the neutral atom. Remember Au is one of our exceptions.

Au [Xe]$6s^1 4f^{14} 5d^{10}$

Since this ion has a + charge, remove one electron to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 6*s* orbital before removing electrons from the 5*d* or 4*f* orbitals.

Au^+ [Xe]$6s^0 4f^{14} 5d^{10}$

[Xe] [ ] [↓↑][↓↑][↓↑][↓↑][↓↑][↓↑][↓↑] [↓↑][↓↑][↓↑][↓↑][↓↑]

6*s* 4*f* 5*d*

Au^+ is diamagnetic.

(c) Mo^{3+} Begin by writing the electron configuration of the neutral atom. Remember, Mo is one of our exceptions.

Mo [Kr]$5s^1 4d^5$

Since this ion has a 3+ charge, remove three electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 5*s* orbital before removing electrons from the 4*d* orbitals.

Mo^{3+} [Kr]$5s^0 4d^3$

[Kr] [ ] [↑][↑][↑][ ][ ]

5*s* 4*d*

Mo^{3+} is paramagnetic.

(d) Zr^{2+} Begin by writing the electron configuration of the neutral atom.

Zr [Kr]$5s^2 4d^2$

Since this ion has a 2+ charge, remove two electrons to write the electron configuration of the ion. Since it is a transition metal, remove the electrons from the 5*s* orbital before removing electrons from the 4*d* orbitals.

Zr^{2+} [Kr]$5s^0 4d^2$

[Kr] [ ] [↑][↑][ ][ ][ ]

5*s* 4*d*

Zr^{2+} is paramagnetic.

8.69 (a) Li or Li^+ A Li atom is larger than Li^+ because cations are smaller than the atoms from which they are formed.

(b) I^- or Cs^+ An I^- ion is larger than a Cs^+ ion because, although they are isoelectronic, I^- has two fewer protons than Cs^+, resulting in a lesser pull on the electrons and therefore a larger radius.

(c) Cr or Cr^{3+} A Cr atom is larger than Cr^{3+} because cations are smaller than the atoms from which they are formed.

(d) O or O^{2-} An O^{2-} ion is larger than an O atom because anions are larger than the atoms from which they are formed.

8.70 (a) Sr or Sr^{2+} A Sr atom is larger than Sr^{2+} because cations are smaller than the atoms from which they are formed.

(b) N or N^{3-} An N^{3-} ion is larger than an N atom because anions are larger than the atoms from which they are formed.

(c) Ni or Ni^{2+} A Ni atom is larger than Ni^{2+} because cations are smaller than the atoms from which they are formed.

(d) S^{2-} or Ca^{2+} An S^{2-} ion is larger than a Ca^{2+} ion because, although they are isoelectronic, S^{2-} has four fewer protons than Ca^{2+}, resulting in a lesser pull on the electrons and therefore a larger radius.

8.71 Since all the species are isoelectronic, the radius will depend on the number of protons in each species. The fewer the protons the larger the radius.
F: Z = 9; Ne: Z = 10; O: Z = 8; Mg: Z = 12; Na: Z = 11
So: $O^{2-} > F^{-} > Ne > Na^{+} > Mg^{2+}$

8.72 Since all the species are isoelectronic, the radius will depend on the number of protons in each species. The fewer the protons, the larger the radius.
Se: Z = 34; Kr: Z = 36; Sr: Z = 38; Rb: Z = 37; Br: Z = 35
So: $Sr^{2+} < Rb^{+} < Kr < Br^{-} < Se^{2-}$

8.73 (a) Br or Bi Br has a higher ionization energy than Bi because, as you trace the path between Br and Bi on the periodic table, you move down a column (ionization energy decreases) and then to the left across a period (ionization energy decreases). These effects sum together for an overall decrease.

(b) Na or Rb Na has a higher ionization energy than Rb because, as you trace a path between Na and Rb on the periodic table, you move down a column. Ionization energy decreases as you go down a column because of the increasing size of orbitals with increasing n.

(c) As or At Based on periodic trends alone, it is impossible to tell which has a higher ionization energy because as you trace the path between As and At you go to the right across a period (ionization energy increases) and then down a column (ionization energy decreases). These effects tend to oppose each other, and it is not obvious which will dominate.

(d) P or Sn P has a higher ionization energy than Sn because as you trace the path between P and Sn on the periodic table you move down a column (ionization energy decreases) and then to the left across a period (ionization energy decreases). These effects sum together for an overall decrease.

8.74 (a) P or I Based on periodic trends alone, it is impossible to tell which has a higher ionization energy because as you trace the path between P and I you go to the right across a period (ionization energy increases) and then down a column (ionization energy decreases). These effects tend to oppose each other, and it is not obvious which will dominate.

(b) Se or Cl Cl has a higher ionization energy than Se because as you trace the path between Cl and Se on the periodic table you move down a column (ionization energy decreases) and then to the left across a period (ionization energy decreases). These effects sum together for an overall decrease.

(c) P or Sb P has a higher ionization energy than Sb because as you trace a path between P and Sb on the periodic table you move down a column. Ionization energy decreases as you go down a column because of the increasing size of orbitals with increasing n.

(d) Ga or Ge Ge has a higher ionization energy than Ga because as you trace a path between Ga and Ge on the periodic table you move to the right within the same period. Ionization energy increases as you go to the right because of increasing effective nuclear charge.

8.75 Since ionization energy increases as you move to the right across a period and increases as you move up a column, the element with the smallest first ionization energy would be the element farthest to the left and lowest down on the periodic table. So, In has the smallest ionization energy; as you trace a path to the right and up on the periodic table, the next element reached is Si; continuing up and to the right you reach N; and then continuing to the right you reach F. So, in the order of increasing first ionization energy the elements are In $<$ Si $<$ N $<$ F.

8.76 Since ionization energy increases as you move to the right across a period and increases as you move up a column, the element with the largest first ionization energy would be the element farthest to the right and highest up on the periodic table. So, Cl has the largest ionization energy; as you trace a path to the left on the periodic table you reach Sp; as you move down a column and to the left you reach Sn; and then moving down the column you reach Pb. So, in the order of decreasing first ionization energy the elements are Cl $>$ S $>$ Sn $>$ Pb.

8.77 The jump in ionization energy occurs when you change from removing a valence electron to removing a core electron. To determine where this jump occurs you need to look at the electron configuration of the atom.

(a) Be $1s^22s^2$ The first and second ionization energies involve removing 2*s* electrons, while the third ionization energy removes a core electron, so the jump will occur between the second and third ionization energies.

(b) N $1s^22s^22p^3$ The first five ionization energies involve removing the 2*p* and 2*s* electrons, while the sixth ionization energy removes a core electron, so the jump will occur between the fifth and sixth ionization energies.

(c) O $1s^22s^22p^4$ The first six ionization energies involve removing the 2*p* and 2*s* electrons, while the seventh ionization energy removes a core electron, so the jump will occur between the sixth and seventh ionization energies.

(d) Li $1s^22s^1$ The first ionization energy involves removing a 2*s* electron, while the second ionization energy removes a core electron, so the jump will occur between the first and second ionization energies.

8.78 The jump occurs between IE_3 and IE_4, so removing the first three electrons involves removing valence electrons and the fourth electron is a core electron, so the valence electron configuration would be ns^2np^1; this puts the element in column 3A and would be Al.

Electron Affinities and Metallic Character

8.79 (a) Na or Rb Na has a more negative electron affinity than Rb. In column 1A electron affinity becomes less negative as you go down the column.

(b) B or S S has a more negative electron affinity than B. As you trace from B to S in the periodic table you move to the right, which shows the value of the electron affinity becoming more negative. Also, as you move from period 2 to period 3 the value of the electron affinity becomes more negative. Both of these trends sum together for the value of the electron affinity to become more negative.

(c) C or N C has the more negative electron affinity. As you trace from C to N across the periodic table you would normally expect N to have the more negative electron affinity. However, N has a half-filled p sublevel, which lends it extra stability, therefore it is harder to add an electron.

(d) Li or F F has the more negative electron affinity. As you trace from Li to F on the periodic table you move to the right in the period. As you go to the right across a period the value of the electron affinity generally becomes more negative.

8.80 (a) Mg or S S has the more negative electron affinity. As you trace from Mg to S on the periodic table you move to the right in the period. As you go to the right across a period the value of the electron affinity generally becomes more negative.