CHAPTER 7 ELECTRON CONFIGURATION AND CHEMICAL PERIODICITY

END–OF–CHAPTER PROBLEMS

7.1 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.

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

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

b) Predicted melting point (Br_2) =

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

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

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

7.11 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.

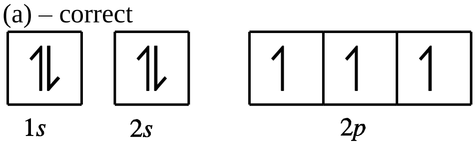
7.13 Plan: The integer in front of the letter represents the n value. The l value designates the orbital type: $l = 0 = s$ orbital; $l = 1 = p$ orbital; $l = 2 = d$ orbital; $l = 3 = f$ orbital. Remember that a p orbital set contains 3 orbitals, a d orbital set has 5 orbitals, and an f orbital set has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

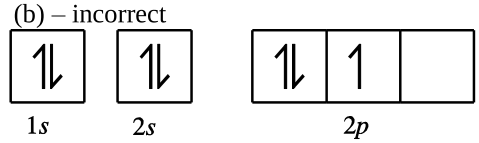
c) **14** electrons can be found in the 5*f* orbitals ($l = 3$ designates f orbitals; there are 7*f* orbitals in a set).

7.16 Hund's rule states that electrons will fill empty orbitals in the same subshell 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



7.18 For elements in the same group (vertical column in periodic table), the electron configuration of the valence electrons is 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 valence level.

7.20 The total electron capacity for an energy level is $\mathbf{2n^2}$, 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: $4s^24p^64d^{10}4f^{14}$.

7.21 <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. 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 could be $\mathbf{n = 5, l = 0, m_l = 0,}$ and $\mathbf{m_s = +1/2}$ **or -1/2**.
b) The S^- ion would have the configuration $[Ne]3s^23p^5$. The electron added would go into a different orbital than the first electron paired electron and would be the second electron in the orbital it enters. Quantum numbers could be $\mathbf{n = 3, l = 1, m_l = +1, 0}$ **or -1 (depending on where the 4th electron has been placed),** and $\mathbf{m_s = -1/2}$ **or +1/2 depending on the spin of the electron already in the orbital that the added electron enters**.
c) Ag atoms have the configuration $[Kr]5s^14d^{10}$. The electron lost would be from the 5*s* orbital with quantum numbers $\mathbf{n = 5, l = 0, m_l = 0,}$ and $\mathbf{m_s = +1/2}$ **or -1/2, depending on which electron was lost**.
d) The F atom has the configuration $[He]2s^22p^5$. The electron gained would go into the only 2p orbital with a single electron and would be the second electron in that orbital. Quantum numbers could be $\mathbf{n = 2, l = 1, m_l = +1,}$ and $\mathbf{m_s = -1/2}$ **or +1/2 depending on the spin of the electron already in that orbital**.

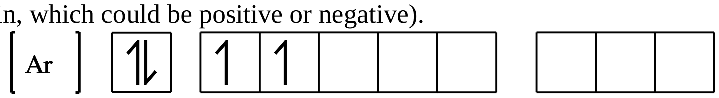
7.23 <u>Plan:</u> The atomic number gives the number of electrons and the periodic table shows the order for filling subshells. Recall that *s* orbitals hold a maximum of 2 electrons, a *p* orbital set holds 6 electrons, a *d* orbital set holds 10 electrons, and an *f* orbital set holds 14 electrons.
<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: $\mathbf{1s^22s^22p^63s^23p^6}$

7.25 <u>Plan:</u> The atomic number gives the number of electrons and the periodic table shows the order for filling subshells. 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) Cl: $\mathbf{1s^22s^22p^63s^23p^5}$
b) Si: $\mathbf{1s^22s^22p^63s^23p^2}$
c) Sr: $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2}$

7.27 Plan: The atomic number gives the number of electrons and the periodic table shows the order for filling subshells. 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.

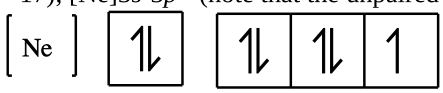
Solution:

a) Ti ($Z = 22$); [Ar]$4s^2 3d^2$ (note that the 2 3d electrons can be placed in any of the 5 boxes as long as they have the same spin, which could be positive or negative).



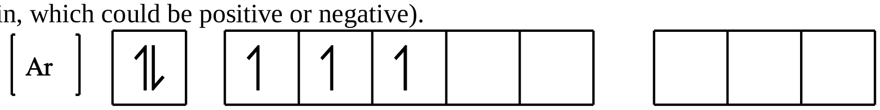
$4s$ $3d$ $4p$

b) Cl ($Z = 17$); [Ne]$3s^2 3p^5$ (note that the unpaired electron could have positive or negative spin).



$3s$ $3p$

c) V ($Z = 23$); [Ar]$4s^2 3d^3$ (note that the 3 3d electrons can be placed in any of the 5 boxes as long as they have the same spin, which could be positive or negative).

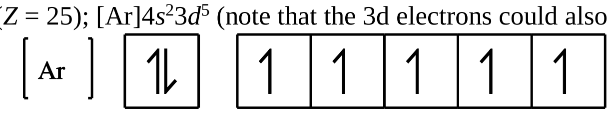


$4s$ $3d$ $4p$

7.29 Plan: The atomic number gives the number of electrons and the periodic table shows the order for filling subshells. 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.

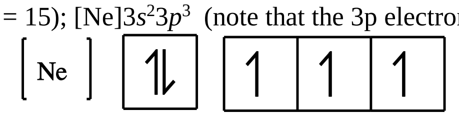
Solution:

a) Mn ($Z = 25$); [Ar]$4s^2 3d^5$ (note that the 3d electrons could also ALL have negative spin).



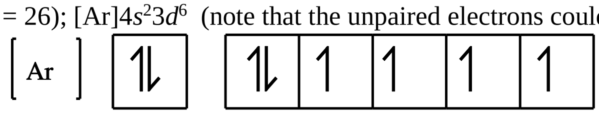
$4s$ $3d$

b) P ($Z = 15$); [Ne]$3s^2 3p^3$ (note that the 3p electrons could also ALL have negative spin).



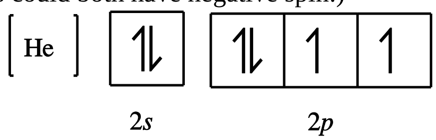
$3s$ $3p$

c) Fe ($Z = 26$); [Ar]$4s^2 3d^6$ (note that the unpaired electrons could also all have negative spins).

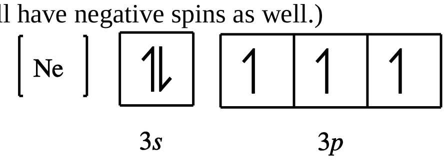


$4s$ $3d$

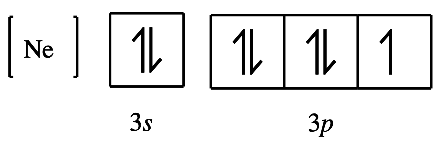
7.31 Plan: Add up all of the electrons in the electron configuration to obtain the atomic number of the element which is then used to identify the element and its position in the periodic table. When drawing the partial orbital diagram, only include electrons after those of the previous noble gas; remember 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 16, Period 2. (Note that the unpaired 2p electrons could both have negative spin.)



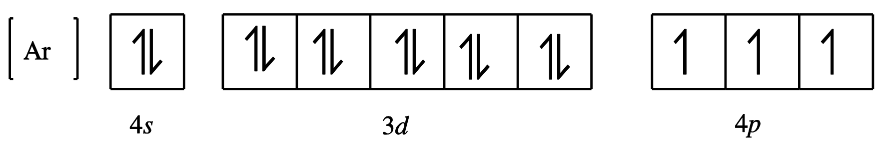
b) There are 15 electrons in the configuration; the element is P, Group 15, Period 3. (Note that the 3p electrons could all have negative spins as well.)



7.33 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 17 electrons in the configuration; the element is Cl; Group 17; Period 3. Note that the single electron could have negative spin.



b) There are 33 electrons in the configuration; the element is As; Group 15; Period 4. Note that the 4p electrons could all have negative spin.



7.35 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 **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 **18**.

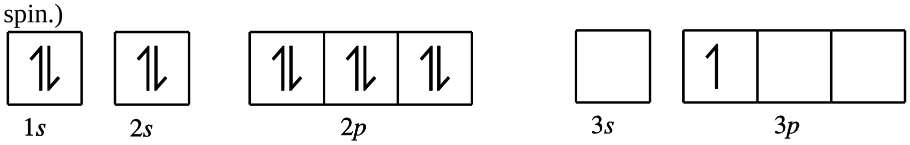
7.37 Plan: Core electrons are those seen in the previous noble gas and completed transition series (d orbitals). Valence electrons are those in the highest energy level (highest n value). For transition metals, valence electrons also include electrons in the outermost d set of orbitals. It is easiest to determine the types of electrons by writing a condensed electron configuration.

Solution:
a) O (Z = 8); [He]$2s^22p^4$. There are **2** core electrons (represented by [He]) and 4 valence electrons.
b) Sn (Z = 50); [Kr]$5s^24d^{10}5p^2$. There are 36 (from [Kr]) + 10 (from the filled 4d set) = **46** core electrons. The highest energy level is n = 5 so there are **4** valence electrons.
c) Ca (Z = 20); [Ar]$4s^2$. There are **2** valence electrons (the 4s electrons), and **18** core electrons (from [Ar]).
d) Fe (Z = 26); [Ar]$4s^23d^6$. There are **8** valence electrons (2 from n = 4 level and the d orbital electrons count in this case because the subshell is not full) , and **18** core electrons (from [Ar]).
e) Se (Z = 34); [Ar]$4s^23d^{10}4p^4$. There are (2 + 4 in the n = 4 level), **6** valence electrons (filled d subshells count as core electrons), and **28** core electrons (18 from [Ar] and 10 from the filled 3d set).

7.39 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 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 Group16. Other elements in Group 16 are **O, Se, Te,** and **Po**.
c) Electrons total 3 + 54 (from xenon) = 57. Element 57 is lanthanum, **La**, in Group 3. Other elements in this group are **Sc, Y,** and **Ac**.

7.41 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]$2s^22p^2$ has a total of 6 electrons (4 + 2 from He configuration) which is element carbon with symbol **C**; other Group 14 elements include **Si, Ge, Sn,** and **Pb**.
b) Electrons total 5 + 18 (from argon) = 23 which is **vanadium**; other Group 5 elements include **Nb, Ta,** and **Db**.
c) The electrons in this element total 15, 10 from the neon configuration plus 5 from the rest of the configuration. Element 15 is **phosphorus**; other Group 15 elements include **N, As, Sb,** and **Bi**.

7.43 Plan: Write the ground-state electron configuration of sodium; for the excited state, move the outermost electron to the next orbital.
Solution:
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 $1s^22s^22p^63p^1$. (Note that the electron in 3p can be in any of the 3 boxes and have positive or negative spin.)



7.50 A high, endothermic IE_1 means it is very difficult to remove the first valence electron. This value would exclude any metal, because metals lose a valence 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 17, which form **–1** ions.

7.53 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 valence 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.

7.55 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 valence 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.

7.57 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 $\mathbf{1s^22s^22p^1}$ which represents boron, **B**.

7.58The 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 $\mathbf{1s^22s^22p^63s^2}$**, Mg**.

7.59 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 valence 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.

7.61 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.

7.62 Metallic character decreases up 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.

7.63 Plan: Write the electron configurations for the two elements. Remember that these elements lose electrons to achieve pseudo-noble gas configurations.
Solution:
The two largest elements in Group 14, 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]$5s^24d^{10}$) + 2 e^-
Pb ([Xe]$6s^25d^{10}6p^2$) $\rightarrow$ Pb^{2+} ([Xe]$6s^25d^{10}$) + 2 e^-
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]$4d^{10}$) + 4 e^-
Pb ([Xe]$6s^25d^{10}6p^2$) $\rightarrow$ Pb^{4+} ([Xe]$5d^{10}$) + 4 e^-
Possible ions for tin and lead have **+2** and **+4** charges.

7.67 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 2.
c) **I** is more metallic because it lies below Br in Group 17.

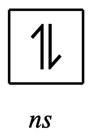
7.69 Plan: Metallic behavior decreases up a group and decreases left to right across a period.
Solution:
a) **As** should be less metallic than antimony because it lies above Sb in the same group of the periodic table.
b) **P** should be less metallic because it lies to the right of silicon in the same 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.

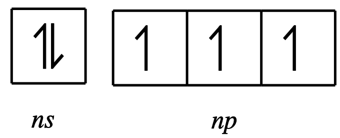
7.71 Plan: For main-group elements, the most stable ions have electron configurations identical to noble gas atoms. Write the electron configuration of the atom and then remove or add electrons until a noble gas configuration is achieved. Metals lose electrons and nonmetals gain electrons.
Solution:
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+}, $\mathbf{1s^22s^22p^63s^23p^6}$.

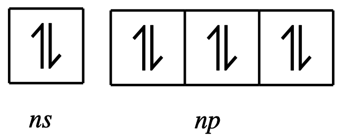
7.73 Plan: For main-group elements, the most stable ions have electron configurations identical to noble gas atoms. Write the electron configuration of the atom and then remove or add electrons until a noble gas configuration is achieved. Metals lose electrons and nonmetals gain electrons.
Solution:
a) Al: $1s^22s^22p^63s^23p^1$; aluminum atoms contain three more electrons than the noble gas configuration of Ne. Thus, an aluminum atom loses its 3 valence shell electrons to form a **+3** ion: Al^{3+}, $\mathbf{1s^22s^22p^6}$.
b) S: $1s^22s^22p^63s^23p^4$; sulfur atoms are two electrons short of the noble gas configuration of argon. Thus, a sulfur atom gains two electrons to form a **–2** ion: S^{2-}, $\mathbf{1s^22s^22p^63s^23p^6}$.
c) Sr: $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^2$; strontium atoms contain two more electrons than the noble gas configuration of krypton. Thus, a strontium atom loses two electrons to form a **+2** ion: Sr^{2+}, $\mathbf{1s^22s^22p^63s^23p^64s^23d^{10}4p^6}$.

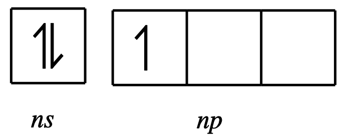
7.75 Plan: To find the number of unpaired electrons look at the electron configuration expanded to include the different orientations of the orbitals, such as p_x and p_y and p_z. Remember that one electron will occupy every orbital in a set (*p*, *d*, or *f*) before electrons will pair in an orbital in that set. In the noble gas configurations, all electrons are paired because all orbitals are filled.

Solution:
a) Configuration of 2 group elements: [noble gas]ns^2, **no unpaired electrons**. The electrons in the *ns* orbital are paired.
b) Configuration of 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. Spins can be all positive or all negative.
c) Configuration of 18 group elements: noble gas configuration ns^2np^6 with no half-filled orbitals, **no unpaired electrons**.
d) Configuration of 13 group elements: [noble gas]ns^2np^1. There is **one** unpaired electron in one of the *p* orbitals. The unpaired electron can be in any of the p orbitals and can have positive or negative spin.

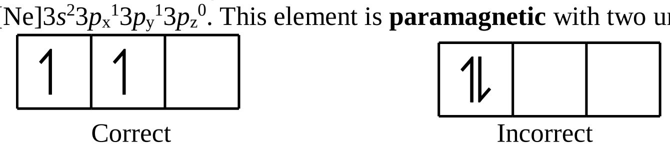
a) 

b) 

c) 

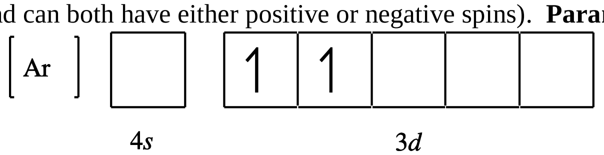
d) 

7.77 Plan: Substances are paramagnetic if they have unpaired electrons. 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 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) Ga (Z = 31) = [Ar]$4s^23d^{10}4p^1$. The *s* and *d* subshells are filled, so all electrons are paired. The lone *p* electron is unpaired, so this element is **paramagnetic**.
b) Si (Z = 14) = [Ne]$3s^23p_x^13p_y^13p_z^0$. This element is **paramagnetic** with two unpaired electrons.

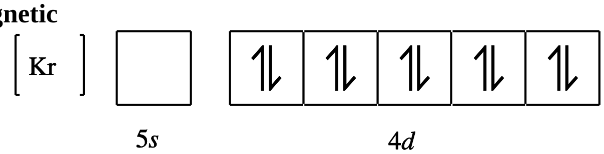


c) Be (Z = 4) = [He]$2s^2$. The two *s* electrons are paired so Be is **not paramagnetic**.
d) Te (Z = 52) = [Kr]$5s^24d^{10}5p_x^25p_y^15p_z^1$ is **paramagnetic** with two unpaired electrons in the *5p* set.

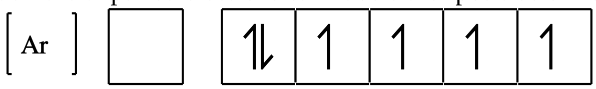
7.79 Plan: Substances are paramagnetic if they have unpaired electrons. Write the electron configuration of the atom and then remove the specified number of electrons. Remember that all orbitals in a *p*, *d*, or *f* set will each have one electron before electrons pair in an orbital. In the noble gas 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. Note that the two d electrons can be in any of the boxes and can both have either positive or negative spins). **Paramagnetic**



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

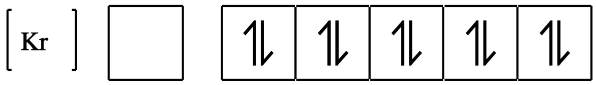


c) Co: [Ar]$4s^2 3d^7$; **Co^{3+}: [Ar]$3d^6$** Cobalt atoms lose two 4*s* electrons and one 3*d* electron to form the +3 ion. Note that all the unpaired electrons can have either positive or negative spins. **Paramagnetic**



4*s* 3*d*

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

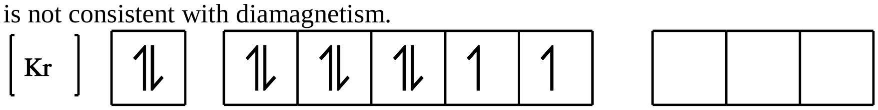


5*s* 4*d*

7.81 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.

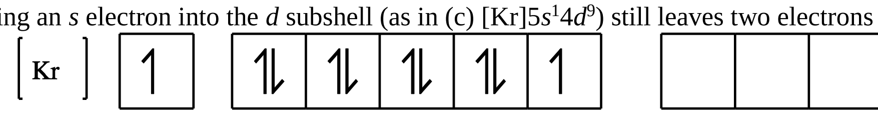
Solution:

You might first write the condensed electron configuration for Pd as [Kr]$5s^2 4d^8$. However, the partial orbital diagram is not consistent with diamagnetism.



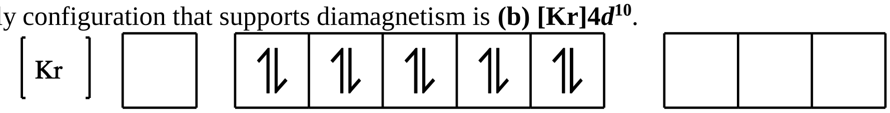
5*s* 4*d* 5*p*

Promoting an *s* electron into the *d* subshell (as in (c) [Kr]$5s^1 4d^9$) still leaves two electrons unpaired.



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

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



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

7.83 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 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.

7.86 Plan: Write the electron configuration for each atom. Remove the specified number of electrons, given by the positive ionic charge, to write the configuration for the ions. Remember that electrons with the highest *n* value are removed first.

Solution:

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

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

7.89 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 18. 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**.

7.90 Plan: Recall Hess's law: the enthalpy change of an overall process is the sum of the enthalpy changes of its individual steps. Both the ionization energies and the electron affinities of the elements are needed.
Solution:
a) F: ionization energy = 1681 kJ/mol electron affinity = –328 kJ/mol

$$F(g) \rightarrow F^{+}(g) + e^{-} \qquad \Delta E = 1681 \text{ kJ/mol}$$
$$F(g) + e^{-} \rightarrow F^{-}(g) \qquad \Delta E = -328 \text{ kJ/mol}$$

Reverse the electron affinity reaction to give: $F^{-}(g) \rightarrow F(g) + e^{-}$ $\Delta E = +328$ kJ/mol
Summing the ionization energy reaction with the reversed electron affinity reaction (Hess's law):

$$\cancel{F(g)} \rightarrow F^{+}(g) + e^{-} \qquad \Delta E = 1681 \text{ kJ/mol}$$
$$F^{-}(g) \rightarrow \cancel{F(g)} + e^{-} \qquad \Delta E = +328 \text{ kJ/mol}$$
$$F^{-}(g) \rightarrow F^{+}(g) + 2\,e^{-} \qquad \Delta E = \mathbf{2009\ kJ/mol}$$

b) Na: ionization energy = 496 kJ/mol electron affinity = –52.9 kJ/mol

$$Na(g) \rightarrow Na^{+}(g) + e^{-} \qquad \Delta E = 496 \text{ kJ/mol}$$
$$Na(g) + e^{-} \rightarrow Na^{-}(g) \qquad \Delta E = -52.9 \text{ kJ/mol}$$

Reverse the ionization reaction to give: $Na^{+}(g) + e^{-} \rightarrow Na(g)$ $\Delta E = -496$ kJ/mol
Summing the electron affinity reaction with the reversed ionization reaction (Hess's law):

$$\cancel{Na(g)} + e^{-} \rightarrow Na^{-}(g) \qquad \Delta E = -52.9 \text{ kJ/mol}$$
$$Na^{+}(g) + e^{-} \rightarrow \cancel{Na(g)} \qquad \Delta E = -496 \text{ kJ/mol}$$
$$Na^{+}(g) + 2\,e^{-} \rightarrow Na^{-}(g) \qquad \Delta E = -548.9 \text{ kJ/mol} = \mathbf{-549\ kJ/mol}$$

9.91 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

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