8.20 (a) This configuration has only 8 electrons, while the F atom possesses 9 electrons.
(b) This is an excited-state configuration, because the 2*s* orbital is not full.
(c) This is Pauli-forbidden: *s* orbitals can hold no more than two electrons.
(d) This configuration uses a non-existent orbital: there is no 1*p* orbital.

8.24 Use the filling rules to determine the correct configuration, being aware of exceptions.
N (7 electrons): $1s^2\ 2s^2\ 2p^3$;
Ti (22 electrons): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$;
As (33 electrons): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$;
Xe (54 electrons): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6$.

8.26 The ground state for F is $1s^2\ 2s^2\ 2p^5$. Excited states with no electron having $\boldsymbol{n} > 2$ can be formed by moving electrons out of the 1*s* and/or 2*s* orbital and placing them in the 2*p* orbital. There are only two ways to do this: $1s^1\ 2s^2\ 2p^6$; and $1s^2\ 2s^1\ 2p^6$.

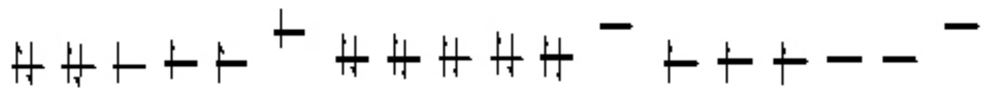
8.56 To determine the correct configuration of a cation, first determine the configuration of the neutral atom, then remove the appropriate number of electrons, removing valence *s* electrons before valence *d* electrons:
Co (27 electrons): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$;
Co^{3+} (remove 3 electrons from Co): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$;
Quantum numbers: $n = 3$; $l = 2$; $m_l = -2, -1, 0, 1, 2$
$m_s = +1/2$ or $-1/2$
Since there are six electrons, Hund's Rule suggests 5 up and 1 down spin.
For example:

n	l	m_l	m_2
3	2	-2	+1/2
3	2	-1	+1/2
3	2	0	+1/2
3	2	+1	+1/2
3	2	+2	+1/2
3	2	0	-1/2

8.58 Total electron spin depends on the number of unpaired electrons. In partially filled shells, each unpaired electron contributes 1/2 to the total spin.
Gd: [Xe] $6s^2\ 5d^1\ 4f^7$, each 4*f* orbital is half filled, so net spin = (7+1)(1/2) = 4;
Sr: [Kr] $5s^2$, all orbitals are filled, so net spin = 0;
Ag^+: [Kr] $4d^{10}$ all orbitals are filled, so net spin = 0.

8.68 Remember that in the transition metal cations, the ***ns*** orbital is less stable than (***n***-1)*d*:

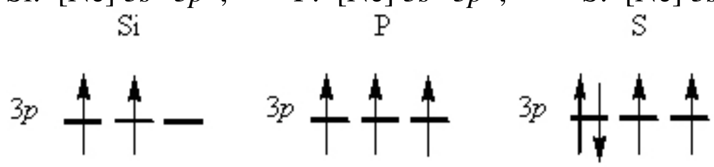
Ir^{+} is $6s^1 5d^7$ Cd^{2+} is $5s^0 4d^{10}$ V^{2+} is $4s^0 3d^3$



(bottom orbitals are *d*, top orbitals are *s*)

8.74 Unpaired electrons occur only among the valence orbitals. Construct the configuration of the neutral atoms:

Si: [Ne] $3s^2\ 3p^2$; P: [Ne] $3s^2\ 3p^3$; S: [Ne] $3s^2\ 3p^4$;



(P has the most unpaired electrons)

8.28 Size increases as the *effective* nuclear charge decreases. As you go down a group (column), although the nuclear charge increases, an extra "shell" of electrons is added that shields the outer electrons from the nucleus. Hence the *effective* nuclear charge decreases and so Cl is larger than F. As you go across a period (row), the nuclear charge is increasing but additional electrons do not effectively shield. Hence, the *effective* nuclear charge increases and the atoms get smaller as you go to the right in a row. Thus Cl is smaller than S is smaller than P. Overall:

F < Cl < S < P.

8.30 The value of IE_2 is less than twice the value of IE_1, indicating that the first two electrons removed are both valence electrons. The value of IE_3 is nearly four times the value of IE_2, indicating that the third electron removed is a core electron rather than a valence electron. The positive electron affinity indicates that an added electron must occupy a new orbital. These facts are consistent with a column 2 element with s^2 valence configuration. (The element is Ba).

8.60 Ionization energy decreases with increasing value of ***n*** and, for the same value of ***n***, increases with increasing *Z* (increasing Coulombic attraction). Cl, Ar, and Ar^+ have ***n*** = 3 for the electron being removed, while Br has ***n*** = 4, so Br has the lowest *IE* of this set. Among the others, the second *IE* of Ar is larger than its first, and Ar has larger *Z* than Cl:

$Ar^+ > Ar > Cl > Br$.

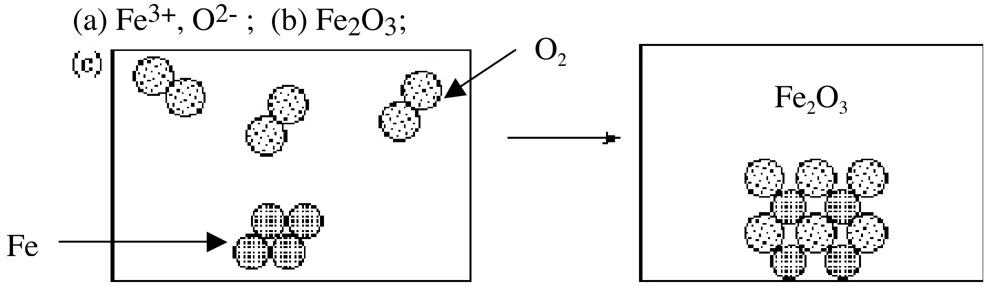
8.90 (a) Successive ionization energies for any element increase because there is less effective screening as electrons are removed. The fourth ionization energy is much larger than any of the other three because the electron that must be removed comes from the 2*p* rather than an ***n*** = 3 orbital. (b) As electrons are removed, the reduction in screening results in tighter binding and a smaller ion. Thus, $Al^{4+} < Al^{3+} < Al^{2+} < Al^{+}$; (c) Electron affinity is the energy released when an electron is added, so it is the reverse of ionization energy, and Al^{4+} has the largest (most negative) value.

2.42 Elements in Groups 1 and 2 easily lose one and two electron(s), respectively. Those in Groups 16 and 17 easily gain two and one electron(s), respectively.
(a) Cs^{+}; (b) Sr^{2+}; and (c) I^{-}.

2.44 Cations and anions combine to form neutral compounds. In any neutral compound, the total amount of positive charge must match the total amount of negative charge. The compounds are: CsI and SrI_2.

2.64 Charge must be conserved, and every chemical compound must be electrically neutral. Three oxygen atoms gain two electrons each for every two iron atoms that lose three electrons each.

(a) Fe^{3+}, O^{2-} ; (b) Fe_2O_3;
(c)



8.42 First identify the chemical eqn for the formation of 1 $CaBr_2$(s):

$$Ca(s) + Br_2(l) \rightarrow CaBr_2(s)$$

Now determine the steps needed for this reaction. Remember that Ca will form a +2 cation and Br will form a –1 anion.

(1)	$Ca(s) \rightarrow Ca(g)$	ΔE = energy of vaporization = 178 kJ/mol
(2)	$Ca(g) \rightarrow Ca^{+}(g) + e^{-}$	$\Delta E = IE(1)$ = 589.8 kJ/mol
(3)	$Ca^{+}(g) \rightarrow Ca^{2+}(g) + e^{-}$	$\Delta E = IE(2)$ = 1145 kJ/mol
(4)	$Br_2(l) \rightarrow Br_2(g)$	ΔE =energy of vaporization = 30.9 kJ/mol
(5)	$Br_2(g) \rightarrow 2Br(g)$	ΔE = bond energy = 224 kJ/mol
(6)	$2[Br(g) + e^{-} \rightarrow Br^{-}(g)]$	$\Delta E = 2EA$ = 2(-324.6 kJ/mol) =

= -649.2 kJ/mol

(7) $2\ Br^{-}(g) + Ca^{2+}(g) \rightarrow CaBr_2(s)$ ΔE = lattice energy = -2176 kJ/mol

Summing up all of these energies gives the overall energy change for the formation of calcium bromide:

(178 + 589.8 + 1145 + 30.9+ 224 + -649.2 + - 2176) kJ/mol =-658 kJ/mol

8.44 A Born-Haber diagram should show the energy of vaporization of the metal, first and second ionization energy of the metal, vaporization energy of the liquid, energy to break the molecular bonds, electron affinity of anions, and the lattice energy that brings the ions of the salt together.