CHAPTER SIX
Electron Configuration

6.1 **Electron configuration** of an atom is how the electrons are distributed among the various atomic orbitals. The Pauli exclusion principle states that no two electrons in an atom can have the same four quantum numbers. Also involved in electron configuration is Hund's rule, which states that the most stable arrangement of electrons in subshells is the one with the greatest number of parallel spins.

6.2 The symbol $4d^6$ gives in chemical shorthand the electron configuration of one of the partially filled shells in an element. 4 shows what energy level (quantum number n) the subshell is in. "d" tells one what type of subshell it is (s, p, d, or f; defined by the quantum number l). 6 tells one how many electrons are in the subshell.

6.3 **Diamagnetic** substances are substances that are slightly repelled by a magnet, whereas **paramagnetic** substances are those that are attracted by a magnet. The lithium atom is an example of a paramagnetic atom, while a helium atom in the ground state is diamagnetic. When it is said that electrons are paired it means that a pair of electrons in the same orbital are antiparallel to each other, or have an opposite spin.

6.4 The term "shielding of electrons" refers to a 2*s* or 2*p* electron being "shielded" from the attractive force of the nucleus by the 1s electrons. The important consequence of the shielding effect is that it reduces the electrostatic attraction between protons in the nucleus and the electron in the 2*s* or 2*p* orbital; this would take place in atom such as Lithium.

6.5 **Transition metals** either have completely filled d subshells or readily give rise to cations that have incompletely filled d subshells. An example of a transition metal is copper. **Lanthanides,** or rare earth series, have incompletely filled 4*f* subshells or readily give rise to cations that have incompletely filled 4*f* subshells. An example of a lanthanide is cerium. The last row of elements are the **actinide series**. Most of these elements are not found in nature but have been synthesized. An example of an actinide is thorium.

6.6 Most transition metals, as previously noted, have incompletely filled *d* subshells or readily give rise to cations that have incompletely filled *d* subshells. In the first transition metal series additional electrons are placed in the 3*d* orbitals, according to Hund's rule. However, copper and chromium are two irregularities to this pattern, which have a different electron configuration than what would be expected. The reason for these irregularities is that a slightly greater stability is associated with the half-filled and completely filled subshells. Electrons in the same subshell have equal energy but different spatial distributions. Consequently, their shielding of one another is relatively small, and the electrons are more strongly attracted by the nucleus when they have the $3d^5$ configuration.

6.7 A **noble gas core** is the electron configuration of the noble gas element that most nearly precedes the element being considered. The electron configuration of a xenon core is as follows: $[Kr]5s^2 4d^{10} 5p6$

6.8 The statement is correct by definition. Pauli's exclusion principle states that no two electrons can share the same set of four quantum numbers.

6.9 **(a)** is wrong because the magnetic quantum number m_l can have only whole number values.

(c) is wrong because the maximum value of the angular momentum quantum number l is $n - 1$.

(e) is wrong because the electron spin quantum number m_s can have only half–integral values.

6.10 For aluminum, there are not enough electrons in the 2*p* subshell. (The 2*p* subshell holds six electrons.) The number of electrons (13) is correct. The electron configuration should be $1s^2 2s^2 2p^6 3s^2 3p^1$. The configuration shown might be an excited state of an aluminum atom.

For boron, there are too many electrons. (Boron only has five electrons.) The electron configuration should be $1s^2 2s^2 2p^1$. What would be the electric charge of a boron ion with the electron arrangement given in the problem?

For fluorine, there are also too many electrons. (Fluorine only has nine electrons.) The configuration shown is that of the F^- ion. The correct electron configuration is $1s^2 2s^2 2p^5$.

6.11 Since the atomic number is odd, it is mathematically impossible for all the electrons to be paired. There must be at least one that is unpaired. The element would be paramagnetic.

6.12 You should write the electron configurations for each of these elements to answer this question. In some cases, an orbital diagram may be helpful.

B: [He]$2s^2 2p^1$ (1 unpaired electron)
P: [Ne]$3s^2 3p^3$ (3 unpaired electrons)
Mn: [Ar]$4s^2 3d^5$ (5 unpaired electrons)
Kr: (0 unpaired electrons)
Cd: [Kr]$5s^2 4d^{10}$ (0 unpaired electrons)
Pb: [Xe]$6s^2 4f^{14} 5d^{10} 6p^2$ (2 unpaired electrons)
Ne: (0 unpaired electrons, Why?)
Sc: [Ar]$4s^2 3d^1$ (1 unpaired electron)
Se: [Ar]$4s^2 3d^{10} 4p^4$ (2 unpaired electrons)
Fe: [Ar]$4s^2 3d^6$ (4 unpaired electrons)
I: [Kr]$5s^2 4d^{10} 5p^5$ (1 unpaired electron)

6.13

B:	$1s^2 2s^2 2p^1$	As:	[Ar]$4s^2 3d^{10} 4p^3$
V:	[Ar]$4s^2 3d^3$	I:	[Kr]$5s^2 4d^{10} 5p^5$
Ni:	[Ar]$4s^2 3d^8$	Au:	[Xe]$6s^1 4f^{14} 5d^{10}$

What is the meaning of "[Ar]"? of "[Kr]"? of "[Xe]"?

6.14 *Step 1:* Germanium (Ge) has 32 electrons. These electrons need to be placed in atomic orbitals.

Step 2: The noble gas element that most nearly proceeds Se is Ar. Therefore, the *noble gas core* is [Ar]. This core accounts for 18 electrons.

Step 3: See Figure 7.24 of your text to check the order of filling subshells past the Ar noble gas core. You should find that the order of filling is 4*s*, 3*d*, then 4*p*. There are 14 remaining electrons to distribute among these orbitals.

The 4*s* orbital can hold *2* electrons. Each of the five 3*d* orbitals can hold *2* electrons for a total of *10* electrons. This leaves *2* electrons to fill the 4*p* orbitals.

The electrons configuration for Ge is:

$$[Ar]\ 4s^2 3d^{10} 4p^2$$

You should follow the same reasoning for the remaining atoms.

Fe: [Ar]$4s^2 3d^6$
W: [Xe]$6s^2 4f^{14} 5d^4$
Zn: [Ar]$4s^2 3d^{10}$
Tl: [Xe]$6s^2 4f^{14} 5d^{10} 6p^1$
Ni: [Ar]$4s^2 3d^8$

6.15 There are a total of twelve electrons:

Orbital	n	l	m_l	m_s
1s	1	0	0	$+\frac{1}{2}$
1s	1	0	0	$-\frac{1}{2}$
2s	2	0	0	$+\frac{1}{2}$
2s	2	0	0	$-\frac{1}{2}$
2p	2	1	1	$+\frac{1}{2}$
2p	2	1	1	$-\frac{1}{2}$
2p	2	1	0	$+\frac{1}{2}$
2p	2	1	0	$-\frac{1}{2}$
2p	2	1	−1	$+\frac{1}{2}$
2p	2	1	−1	$-\frac{1}{2}$
3s	3	0	0	$+\frac{1}{2}$
3s	3	0	0	$-\frac{1}{2}$

The element is magnesium.

6.16 $\underset{3s^2}{\underline{\uparrow\downarrow}}$ $\underset{3p^3}{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}$ $\underset{3s^2}{\underline{\uparrow\downarrow}}$ $\underset{3p^4}{\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}}$ $\underset{3s^2}{\underline{\uparrow\downarrow}}$ $\underset{3p^5}{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}}$

S^+ (5 valence electrons)
3 unpaired electrons

S (6 valence electrons)
2 unpaired electrons

S^- (7 valence electrons)
1 unpaired electron

S^+ has the most unpaired electrons

6.17 The excited atoms are still neutral, so the total number of electrons is the same as the atomic number of the element.

(a) He (2 electrons), $1s^2$
(b) N (7 electrons), $1s^2 2s^2 2p^3$
(c) Na (11 electrons), $1s^2 2s^2 2p^6 3s^1$
(d) As (33 electrons), $[Ar]4s^2 3d^{10} 4p^3$
(e) Cl (17 electrons), $[Ne]3s^2 3p^5$

6.18 Applying the Pauli exclusion principle and Hund's rule:

(a) $\underset{1s^2}{\underline{\uparrow\downarrow}}$ $\underset{2s^2}{\underline{\uparrow\downarrow}}$ $\underset{2p^5}{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}}$

(b) [Ne] $\underset{3s^2}{\underline{\uparrow\downarrow}}$ $\underset{3p^3}{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}$

(c) [Ar] $\underline{\uparrow\downarrow}$ $\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}$

$4s^2$ $3d^7$