Chapter 8:
Periodic Properties of the Elements

8.3 Electron Configurations

- An atom’s electron configuration is the distribution of its electrons into various atomic orbitals.

Rules for assigning electrons to orbitals:

1. **Pauli Exclusion Principle:** No two electrons in an atom can have the same set of four quantum numbers.
    - Thus there can be up to **two electrons** in any atomic orbital (with different spins).
2. **Aufbau (building up) Rule:** Assign electrons to atomic orbitals in order of increasing energy.

Energies of Atomic Orbitals

- The energy of an orbital can be measured by how easy it is to remove an electron from the orbital.
- According to **Coulomb's Law**, the attraction between the nucleus and an electron is stronger the closer the electron is to the nucleus.
- In lower energy orbitals, the electrons tend to be located closer to the nucleus and are harder to remove.
- In higher energy orbitals, the electrons tend to be further from the nucleus.

Energies of Atomic Orbitals

- The energy of an atomic orbital depends on:
  1. The shell (n)
  2. The subshell: $s < p < d < f$ (due to decreasing **penetration** of the core)

Figure 8.3

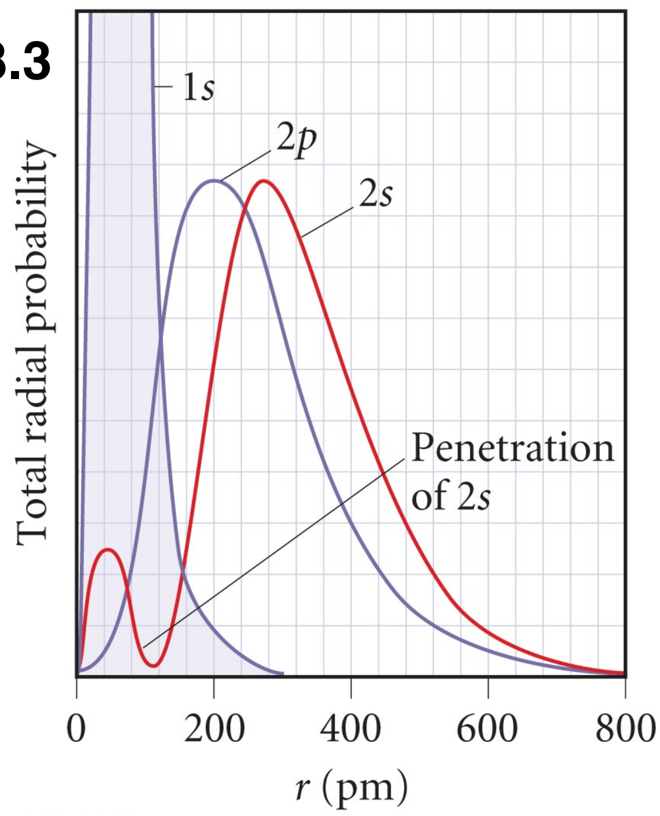


Figure 8.4

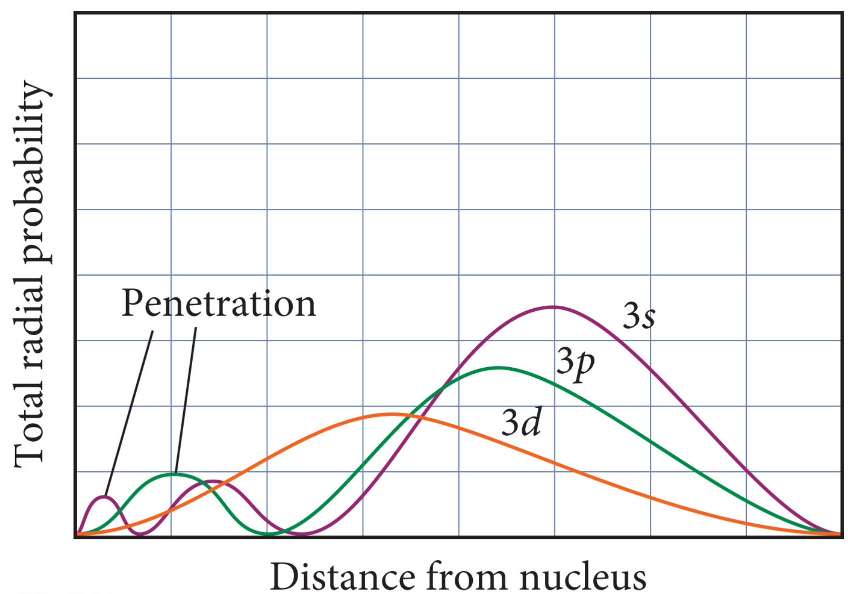
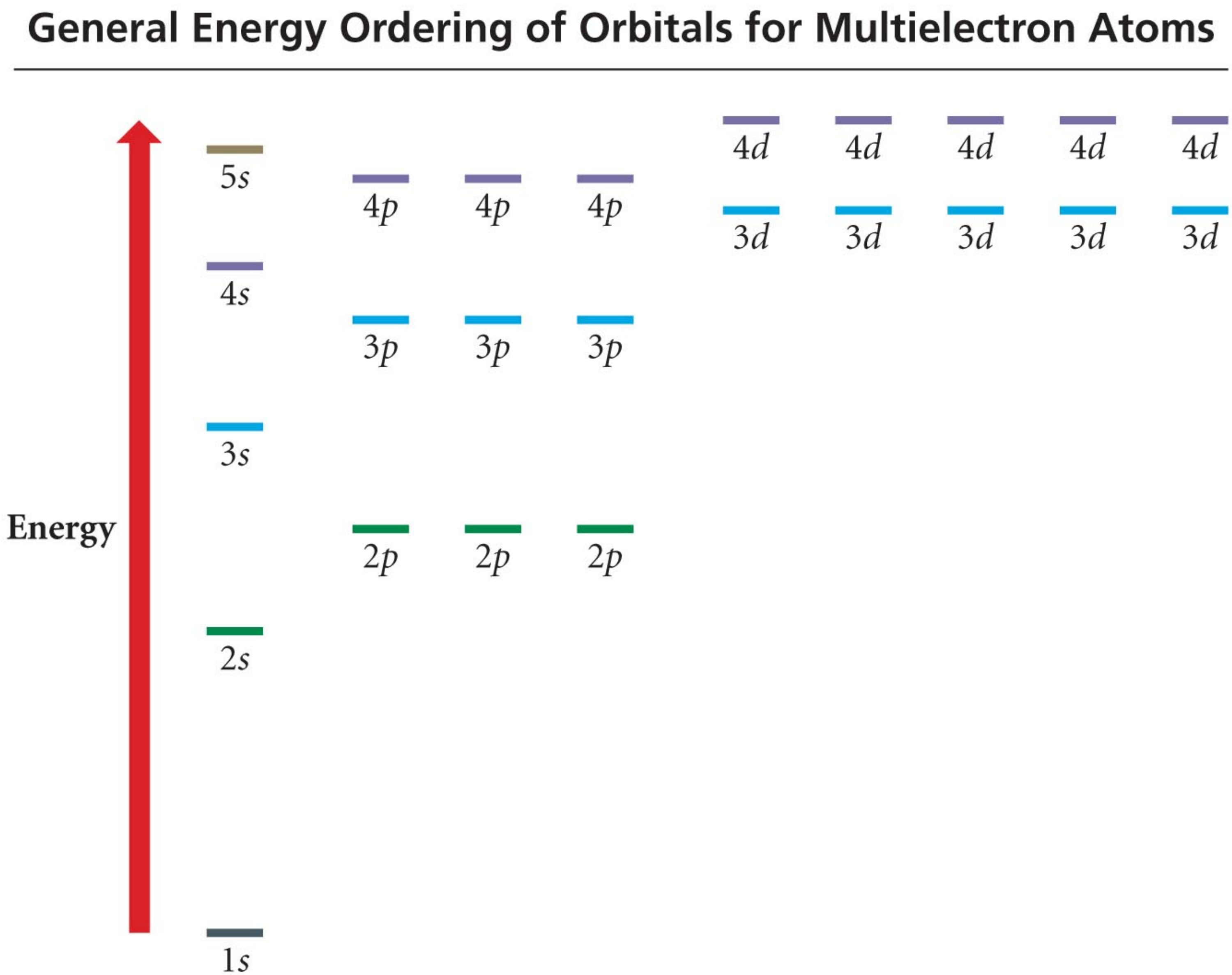


Figure 8.5


General Energy Ordering of Orbitals for Multielectron Atoms
Energy
5s
4s
3s
2s
1s
4p
4p
4p
3p
3p
3p
2p
2p
2p
4d
4d
4d
4d
4d
3d
3d
3d
3d
3d

Orbital Diagrams:

Writing Electron Configurations:

- H:
- He:
- Li:
- B:
- C:

Hund's Rule: For atomic orbitals with the same energy, assign electrons to empty orbitals before pairing up.

Paramagnetic substances contain at least one unpaired electron, and are attracted to a magnetic field.

Diamagnetic substances have all their electrons paired up (weakly <u>repelled</u> by a magnetic field)

Electron Configurations for 2nd Period (Li → Ne)

	1*s*	**2*s***	**2*p***			
Li	↑↓	↑				$1s^22s^1$
Be	↑↓	↑↓				$1s^22s^2$
B	↑↓	↑↓	↑			$1s^22s^22p^1$
C	↑↓	↑↓	↑	↑		$1s^22s^22p^2$
N	↑↓	↑↓	↑	↑	↑	$1s^22s^22p^3$
O	↑↓	↑↓	↑↓	↑	↑	$1s^22s^22p^4$
F	↑↓	↑↓	↑↓	↑↓	↑	$1s^22s^22p^5$
Ne	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^22s^22p^6$

Abbreviated form of electron configurations:

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The Electron Configurations of Transition Metals

21 **Sc** $4s^2 3d^1$	22 **Ti** $4s^2 3d^2$	23 **V** $4s^2 3d^3$	24 **Cr** $4s^1 3d^5$	25 **Mn** $4s^2 3d^5$	26 **Fe** $4s^2 3d^6$	27 **Co** $4s^2 3d^7$	28 **Ni** $4s^2 3d^8$	29 **Cu** $4s^1 3d^{10}$	30 **Zn** $4s^2 3d^{10}$
39 **Y** $5s^2 4d^1$	40 **Zr** $5s^2 4d^2$	41 **Nb** $5s^1 4d^4$	42 **Mo** $5s^1 4d^5$	43 **Tc** $5s^2 4d^5$	44 **Ru** $5s^1 4d^7$	45 **Rh** $5s^1 4d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^1 4d^{10}$	48 **Cd** $5s^2 4d^{10}$

Exceptions:

Example: Write the ground state electron configuration for **Fe**. Is it paramagnetic? If so, how many unpaired electrons are there in the configuration?

Fe: $\overbrace{\square\ \square\ \square\ \square\ \square}^{3d}$

Valence Electrons

- For the main group elements, notice that elements in the same group of the periodic table have the same configuration in their outer shell.
- The outer electrons are called **valence electrons**.

Figure 8.6 **Outer Electron Configurations of Elements 1–18**

1A	2A	3A	4A	5A	6A	7A	8A
1 **H** $1s^1$							2 **He** $1s^2$
3 **Li** $2s^1$	4 **Be** $2s^2$	5 **B** $2s^22p^1$	6 **C** $2s^22p^2$	7 **N** $2s^22p^3$	8 **O** $2s^22p^4$	9 **F** $2s^22p^5$	10 **Ne** $2s^22p^6$
11 **Na** $3s^1$	12 **Mg** $3s^2$	13 **Al** $3s^23p^1$	14 **Si** $3s^23p^2$	15 **P** $3s^23p^3$	16 **S** $3s^23p^4$	17 **Cl** $3s^23p^5$	18 **Ar** $3s^23p^6$

- The number of valence electrons for a main group element is the **same as its “A” group number**.

8.7 The Electron Configuration of Ions

- When an electron is removed from an atom to give a positively charged ion (cation) the highest energy electron is removed:

  e.g. Na: $1s^2 2s^2 2p^6 3s^1 \rightarrow$ Na^+: $1s^2 2s^2 2p^6 =$ [Ne]

- Na^+ is **isoelectronic** with Ne (same number and configuration of electrons).

- When an electron is added to give an anion, follow the Aufbau rule:

  e.g. F: $1s^2 2s^2 2p^5 \rightarrow$ F^-: $1s^2 2s^2 2p^6$ or [Ne]

- F^- is also isoelectronic with Ne.

The Electron Configuration of Ions
(contd.)

- The main group elements tend to form ions that have the same configuration as a noble gas.
- **Transition metals** possess filled ns (n = 4, 5, 6) and partially filled $(n–1)d$ orbitals [configuration: $ns^2(n\text{-}1)d^x$]
- When transition metals form cations, electrons are removed from the **orbitals with the highest *n*-value first**.
- So the transition metals will lose electrons from the the s orbital with highest n before losing d electrons.

Transition Metal Cations

- Example: What is the electron configuration of the Fe^{3+} ion?

  Fe: [Ar]$4s^2 3d^6$

  Fe^{3+}: [Ar]$4s^2 3d^3$ *or* [Ar]$4s^0 3d^5$?