Chapter 7

Electron Structure of the Atom

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Chapter 7 Topics

1. Electromagnetic radiation
2. The Bohr model of the atom
3. The modern model of the atom
4. Periodicity of electron configurations
5. Valence electrons for the main-group elements
6. Electron configurations for ions
7. Periodic properties of atoms

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Light

- What is light, and what does it have to do with the electrons in an atom?

Figure 7.2

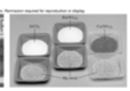


Figure 7.1

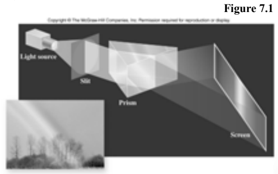
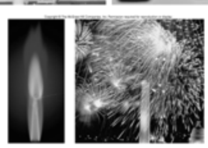


Figure 7.3



fireworks

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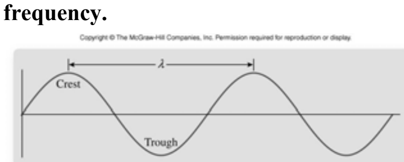
7.1 Electromagnetic Radiation and Energy

- Light is electromagnetic radiation.
- All forms of electromagnetic radiation travel through space as waves that move at the speed of light: 3.0×10^8 m/s.
- Light is characterized by wavelength and frequency.

Light waves

Figure 7.5



Waves

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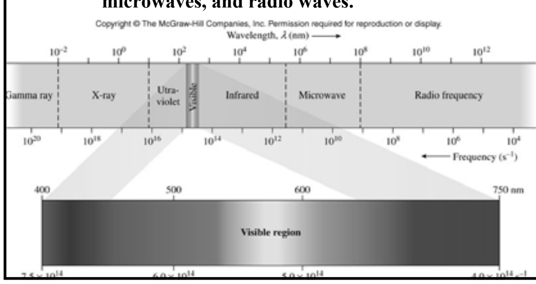
Electromagnetic Spectrum

- Forms of electromagnetic radiation include:
  - gamma rays, X-rays, ultraviolet, visible, infrared, microwaves, and radio waves.



Spectrum

Figure 7.6

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Wavelength and Frequency

- As wavelength increases, frequency decreases.
- Frequency is in units of cycles per second, 1/s or s^{-1}.

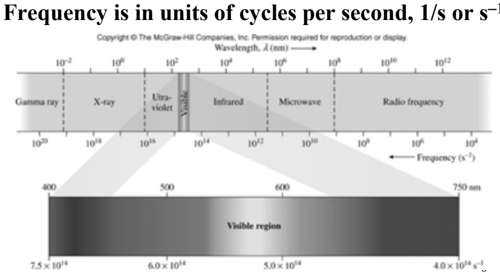
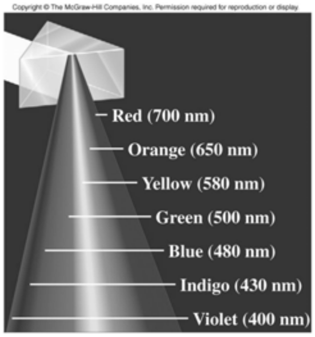


Figure 7.6

Wavelengths of Visible Light

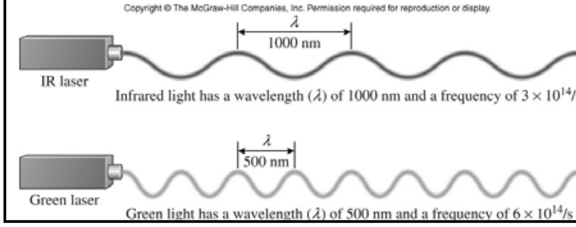


Margin Figure p. 241

ROY G. BIV

Wavelength (λ), Frequency (ν), and Energy (E_{photon})

Figure 7.7



Calculating Wavelength, Frequency, and Photon Energy

- What is the <u>frequency</u> and photon energy of the a laser light used for eye surgery if its wavelength is 710 nm? (watch your units)
- What category does this light fall into?

Atomic Spectra

- When visible light passes through a prism, its components separate into a spectrum.
- White light, such as sun light or light from a regular light bulb, gives a <u>continuous spectrum</u>:

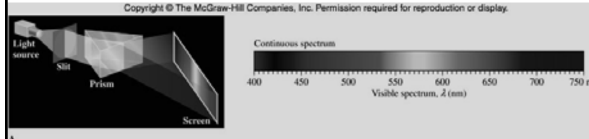
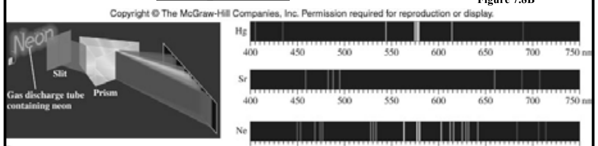


Figure 7.8A

Atomic Spectra

- Colored light, such as that from a neon sign, gives only specific colors in what is called a <u>line spectrum</u>:

Figure 7.8B



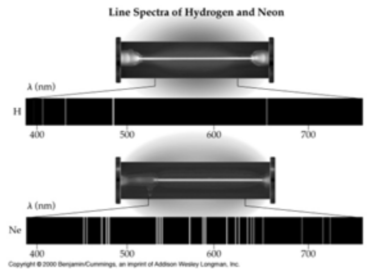
Why do only certain lines appear and what causes these specific colors of light to be emitted? <u>line spectrum</u>

Group Work

- The hydrogen line spectrum contains four colored lines in the visible region.
  a) Which color has the longest wavelength?
  b) Which color has the highest frequency?
  c) Which color has the largest photon energy?

7.2 The Bohr Model of the Hydrogen Atom

- The hydrogen line spectrum contains four visible lines:



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Bohr's Model of the Hydrogen Atom

- Niels Bohr, a student of Rutherford, studied the line spectra of the hydrogen atom to try to understand the electrons in the nuclear model of the atom.
- Bohr came up with a planetary model, in which electrons orbit the nucleus in circular pathways.
- His model was based on the idea that electrons and their energies are quantized: they can have only certain values.

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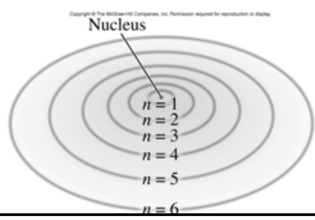
Bohr's Planetary Model

- The electron in the H atom can occupy a specific orbit. Its energy increases as its distance from the nucleus increases.



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Electron transitions to lower energy levels result in the emission of light

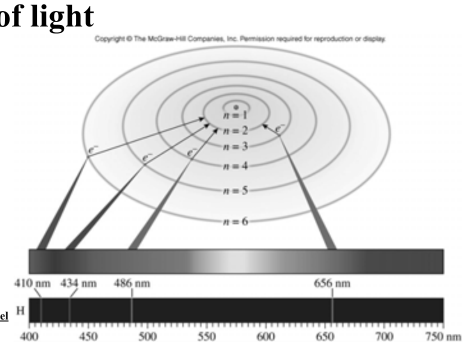

Figure 7.10

Bohr model

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Bohr Model

- When the electron moves from a higher-energy orbit to a lower-energy orbit, it loses energy equal to the difference in energy between the orbits:

Energy change

$$\Delta E = E_{\text{final}} - E_{\text{initial}}$$

- The energy is released as a photon:

$$E_{\text{photon}} = |\Delta E|$$

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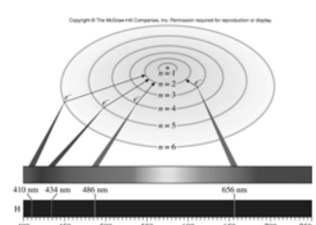
Electron Transitions and Light

- Which transition results in the emission of light with the greatest photon energy?
- Longest wavelength?
- Highest frequency?



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Deficiencies of the Bohr Model

- It explains only the hydrogen atom and its spectrum.
- A more complex model was needed to describes all atoms in general.
- However, the Bohr model did give insight into the quantized behavior of the atom that was later better understood.

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7.3 The Modern Model of the Atom

Modern model

- In the modern model of the atom, electrons are described as existing in orbitals.
- Orbitals are 3-dimensional regions in space where an electron is likely to be found.

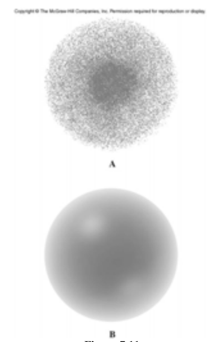

Figure 7.11

Probability Map for lowest-energy state of the electron in a H atom:

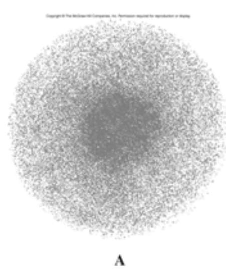

Figure 7.11

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A typical representation of an *s*-orbital:

Figure 7.11

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Principal Energy Levels

- Orbitals of similar size exist in the same principal energy level (n=1, 2, 3 …).

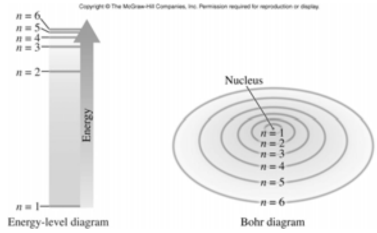

Figure 7.12

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Orbitals

- There are four general types of orbitals:
  - ***s***, ***p***, ***d***, and ***f***

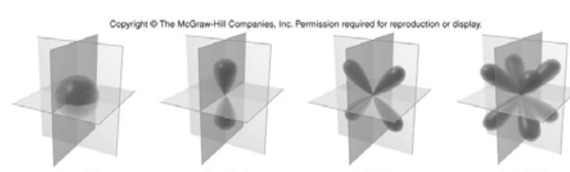

Figure 7.13

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As the principal energy level increases, the orbital extends further from the nucleus.

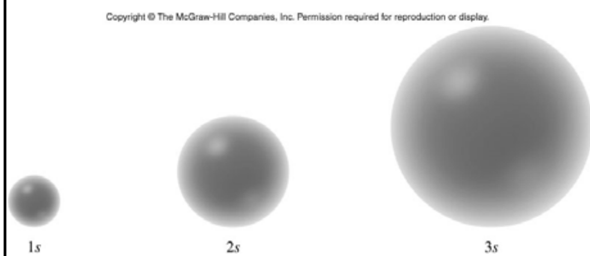

Figure 7.14

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p Orbitals

- **There are three *p* orbitals in a sublevel. They lie perpendicular to one another.**

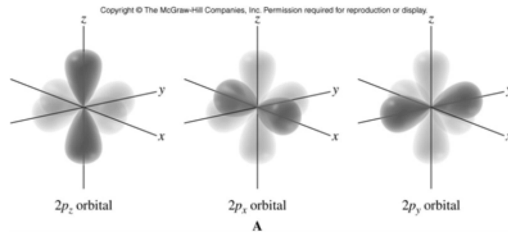

Figure 7.15A

p Orbitals

- ***p*-orbitals overlap with their centers at the nucleus.**

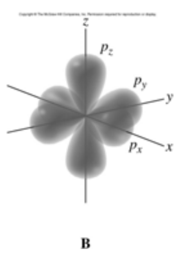

Figure 7.15B

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d Orbitals

- **There are five *d* orbitals in a sublevel.**

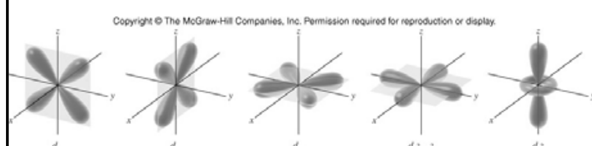

Figure 7.16

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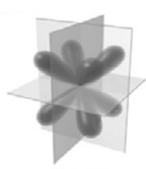
f Orbitals

- **How many *f* orbitals in a sublevel?**



An *f* orbital

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Energies of Sublevels

- **The first sublevel (*n*=1):**
  - **a single *s* orbital**
- **The second sublevel (*n*=2):**
  - **2*s* orbital and the 2*p* sublevel (three 2*p* orbitals)**
- **The third sublevel (*n*=3):**
  - **3*s* orbital, the 3*p* sublevel (three 3*p* orbitals), and the 3*d* sublevel (five 3*d* orbitals)**
- **The fourth sublevel (*n*=4):**
  - **4*s* orbital, the 4*p* sublevel (three 4*p* orbitals), and the 4*d* sublevel (five 4*d* orbitals), and seven 4*d* orbitals.**

Quantum numbers

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Orbital Diagram for the Hydrogen Atom

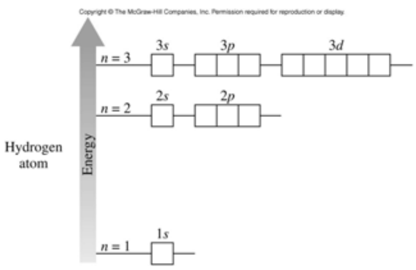


Figure 7.17

Each box represents an orbital

Orbital Diagram for a Multielectron Atom

- <u>Sublevels</u> within a principal energy level split so that
  - *s* < *p* < d < *f*

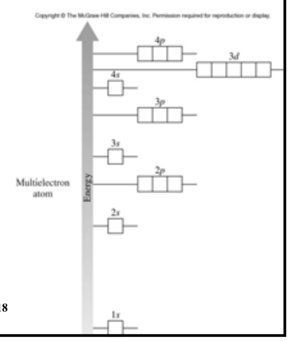


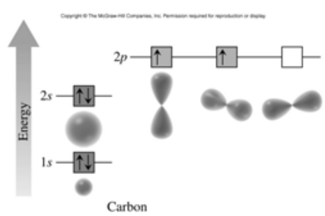
Figure 7.18

Arrangement of Electrons

- How are electrons arranged in an atom?
- We will consider the ground state – the most stable arrangement of lowest energy.

Orbital Diagram for Carbon

- What are some rules for arranging electrons in orbitals and sublevels?



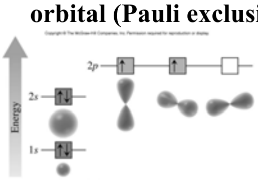
Electrons are represented by up and down arrows.

Figure p. 251

Carbon

Orbital Diagram for Carbon

- Electrons occupy the lowest-energy orbitals first (aufbau principle).
- No more than two electrons occupy each orbital (Pauli exclusion principle).



Two electrons in the same orbital must have opposite spins, represented by the up and down arrows.

Figure p. 251

Helium's Orbital Diagram

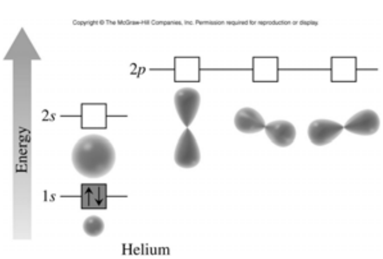


Figure p. 251

Lithium's Orbital Diagram

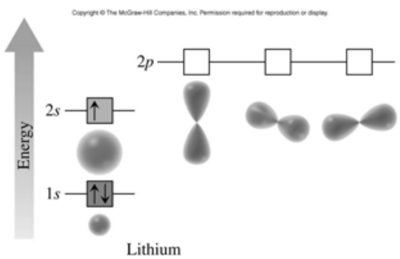


Figure p. 251

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Boron's Orbital Diagram

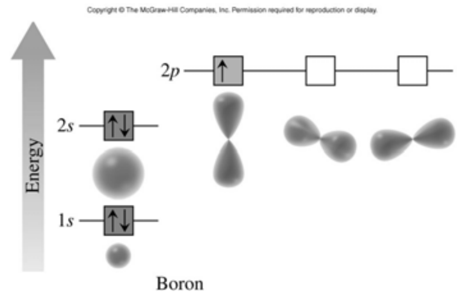

Figure p. 252

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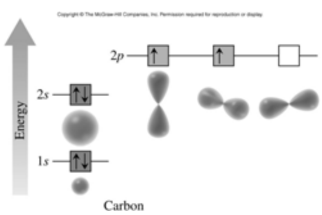
Carbon's Orbital Diagram

- **Why do the electrons in the *p* sublevel occupy separate orbitals?**



It takes a little bit of energy to pair up electrons, so single electrons occupy different orbitals of the same energy (Hund's Rule)

Figure p. 251

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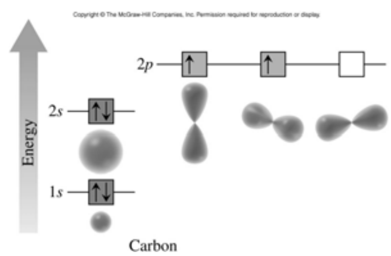
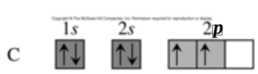


Figure p. 251

Figure p. 252

Electron Configurations

- **An electron configuration is a short-hand notation for representing the number of electrons in each sublevel.**



- **The electron configuration for carbon is:**

$1s^2 2s^2 2p^2$

Check configurations

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Electron Configurations

For elements in periods 1 and 2:

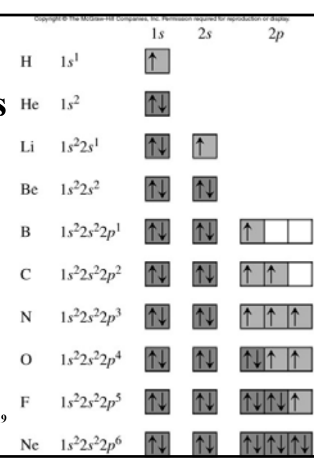
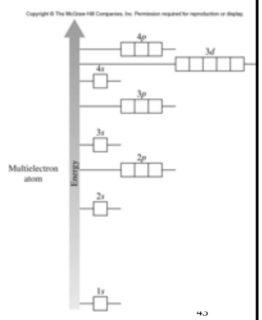


Figure 7.19

Electron Configurations

- Use an orbital diagram to write the electron configuration for silicon:



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7.4 Periodicity of Electron Configurations

- Consider the alkali metals:

Li	$1s^2 2s^1$
Na	$1s^2 2s^2 2p^6 3s^1$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$

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Periodicity of Electron Configurations

- Consider the alkaline earth metals:

Be	$1s^2 2s^2$
Mg	$1s^2 2s^2 2p^6 3s^2$
K	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Rb	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$

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Periodicity of Electron Configurations

- Consider a few of the halogens:

F	$1s^2 2s^2 2p^5$
Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$
Br	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

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Periodicity of Electron Configurations

- Consider some of the noble gases:

Ne	$1s^2 2s^2 2p^6$
Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$
Kr	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

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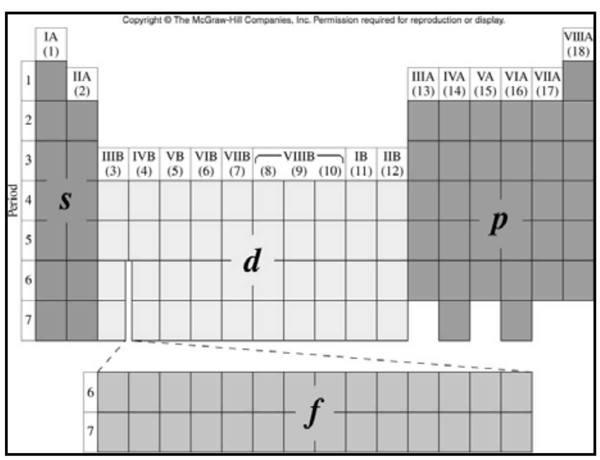

Periodicity of Electron Configurations

- Notice that the number of columns in the *s*, *p*, *d*, and *f* blocks is the same as the number of electrons allowed in each sublevel.

aufbau

- This allows us to use the periodic table to write electron configurations without the aid of an orbital diagram.

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Write the electron configuration for phosphorus:

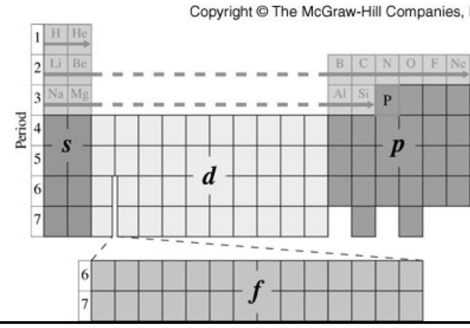

Figure 7.20

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Write the electron configuration for manganese:

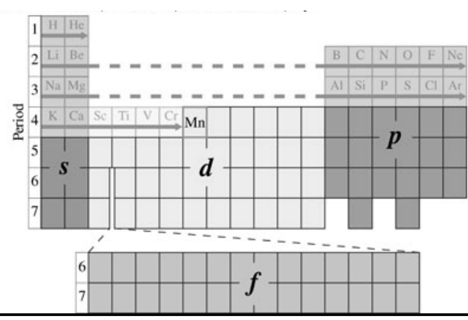

Figure 7.22

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Periodicity of Electron Configurations

- The principal energy level number, the number that comes before the sublevel letter designation, is the same as the period number for the *s* and *p* sublevels.
- For the *d* sublevels, the principal energy level number is one less than the periodic number. Why?

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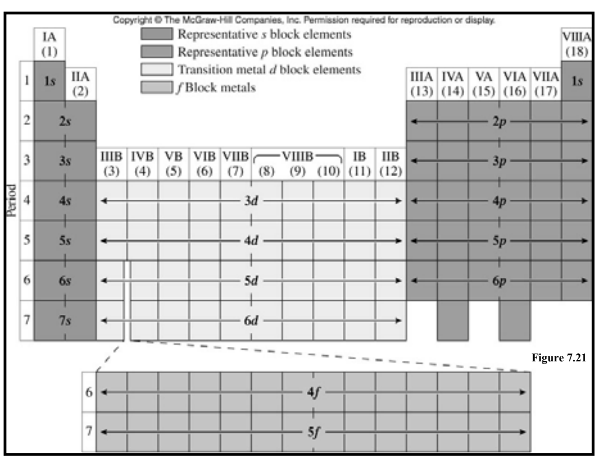

Figure 7.21

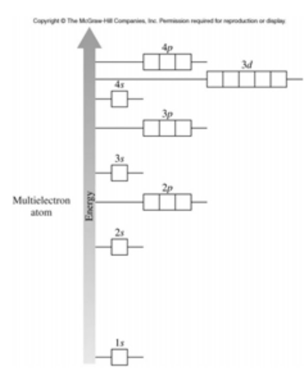

Figure 7.18

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Abbreviated Electron Configurations

- Abbreviated electron configurations are often used for elements with many electrons.
- Notice that iron's electron configuration starts out with argon's electron configuration, but ends differently:

Fe $1s^22s^22p^63s^23p^64s^23d^6$

Ar $1s^22s^22p^63s^23p^6$

- We use the symbol [Ar] to represent argon's electron configuration:

Fe [Ar] $4s^23d^6$

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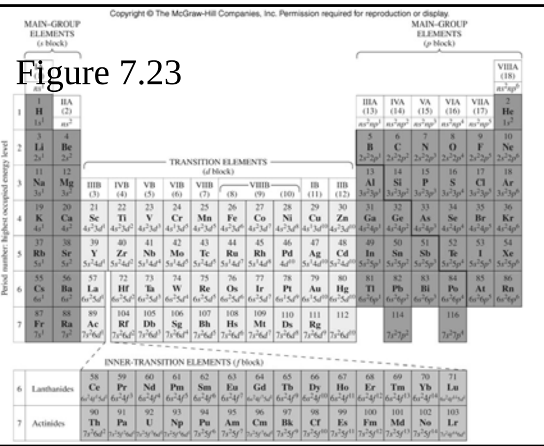

Figure 7.23

7.6 Valence Electrons for the Main-Group Elements

- The last filled principal energy level is called the valence level, or valence shell.
- The valence level contains electrons that are highest in energy and occupy orbitals that extend further from the nucleus than those in the lower levels.
- Valence electrons occupy orbitals in the valence level. All the other electrons are called core electrons, or inner electrons.

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Valence Electrons

- How many valence electrons in bromine?

Br $1s^22s^22p^63s^23p^64s^23d^{10}4p^5$

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Elements in the same group have the same number of valence electrons

- Determine the number of valence electrons in each of the following:

a) F
b) Li
c) Na
d) C
e) Si
f) Pb

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7.6 Electron Configurations for Ions

- In atoms, the number of electrons is equal to the number of protons, which is the atomic number.
- In ions, the number of electrons does not equal the atomic number. We must add or subtract electrons, depending on whether the ion is an anion or cation.
- Write the electron configuration for the potassium ion and the oxide ion.

Ion config

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7.7 Periodic Properties of Atoms

- **Electron configurations are related to the following properties:**
  - **Relative reactivity**
  - **Ionization Energy (tendency to lose electrons)**
  - **Atomic radii (atomic size)**

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Chemical Reactivity

- **From the activity series we see that the most active metals are those in Groups IA and IIA.**
- **The more active the metal, the more easily it loses electrons to form ions.**

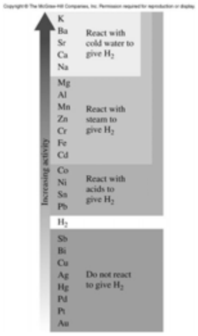

Figure 7.24

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Chemical Reactivity

- **The alkali metals form +1 ions since they all have 1 valence electron.**
  - **Oxides of the alkali metals:**

    Li_2O, Na_2O, K_2O, Rb_2O, Cs_2O
- **The alkaline earth metals form +2 ions since they all have 2 valence electrons.**
  - **Oxides of the alkaline earth metals:**

    BeO, MgO, CaO, SrO, BaO

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Periodic Properties

- **Many properties vary in a regular fashion moving down a group. One example is the reactivity of the alkali metals in water:**

Alkali Metals + Water

Figure 7.25

Li Na K Rb Cs

$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$

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Ionization Energy

- Ionization energy (*IE*) is a measure of the energy required to remove a valence electron from an atom.
- $Mg(g) \rightarrow Mg^{+}(g) + 1e^{-}$ $IE = 738$ kJ/mol

Ionization Energy 1

Ionization Energy 2

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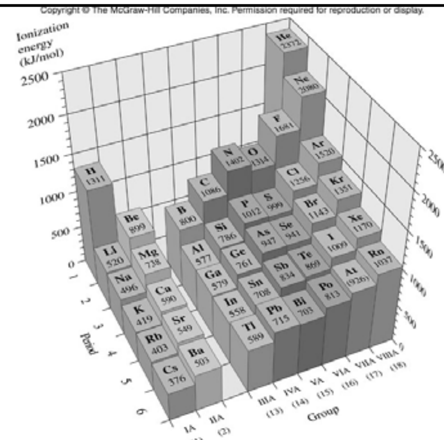

IE

Figure 7.26

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Ionization Energies

- Ionization energy increases up a group because the valence electrons become increasingly closer to the nucleus.

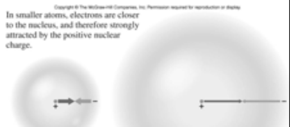


Figure 7.27

- Ionization energy increases from left to right in a period because the valence electrons are increasingly held more tightly by the nucleus, which is increasing in the # of protons.

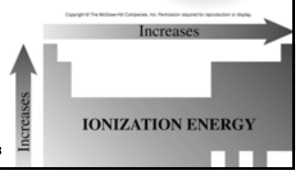


Figure 7.28

Ionization Energies

- Which has the greater ionization energy?
  a) Mg or Ca
  b) O or F
  c) S or F

Periodic Properties

- Explain the changes in reactivity of these metals in terms of ionization energy:

•Alkali Metals + Water

Figure 7.25

Li Na K Rb Cs

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$

Successive Ionization Energies

- 2nd Ionization energy (IE_2)
  $Mg^{+}(g) \rightarrow Mg^{2+}(g) + 1e^{-}$ $IE_2 = 1451$ kJ/mol

- 3rd Ionization energy (IE_3)
  $Mg^{2+}(g) \rightarrow Mg^{3+}(g) + 1e^{-}$ $IE_3 = 7733$ kJ/mol

Ionization Energy 1

Ionization Energy 2

Successive Ionization Energies

- What do these values tell us about the stability of core (inner) electrons?

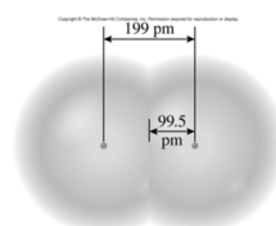
TABLE 7.1 Successive Ionization Energies for Period 2 Elements

Element	IE_1	IE_2	IE_3	IE_4	IE_5	IE_6	IE_7	IE_8	IE_9	IE_{10}
Li	520	7298	11,815							
Be	899	1757	14,849	21,006		Core electrons				
B	800	2427	3660	25,026	32,827					
C	1086	2353	4620	6222	37,830	47,277				
N	1402	2856	4582	7475	9445	53,266	64,360			
O	1314	3388	5300	7469	10,989	13,326	71,334	84,078		
F	1681	3374	6050	8408	11,023	15,164	17,868	92,038	106,434	
Ne	2080	3952	6122	9370	12,178	15,238	19,999	23,069	115,379	131,431

Note: Ionization energies given in kJ/mol.

Atomic Size

- Atomic size is usually reported as atomic radius, one-half the distance between the centers of two bonded atoms.



Cl—Cl

Figure 7.29

Atomic Radii of the Main-Group Elements

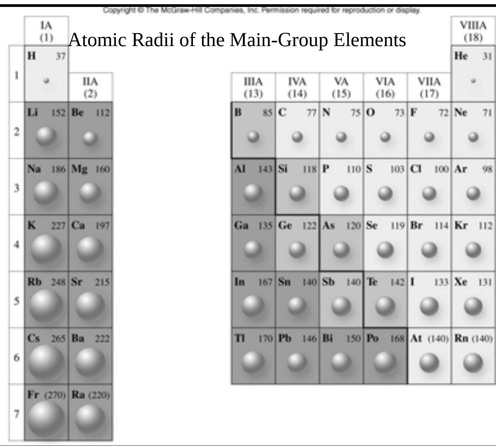


Figure 7.30

Atomic Radii

- <u>Atomic size</u> increases down a group because valence electrons are in orbitals that extend further from the nucleus.
- <u>Atomic size</u> decreases from left to right because electrons are held closer to the nucleus be the increasingly greater charge of the nucleus.

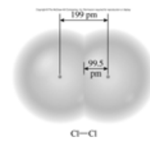


Figure 7.29

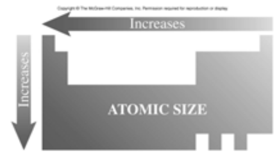


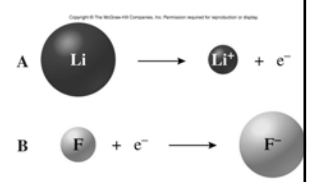
Figure 7.30

Atomic Radii

- **Which has the largest atomic radii?**
  a) Mg or Ca
  b) O or F
  c) S or F

Sizes of Ions

- Cations are smaller than their neutral atoms
- Anions are larger than their neutral atoms.



<u>Ionic radii</u>

Ionic Radii for the Main-Group Ions