Section 11.1

Atoms and Energy

Objectives

1. To describe Rutherford's model of the atom
2. To explore the nature of electromagnetic radiation
3. To see how atoms emit light

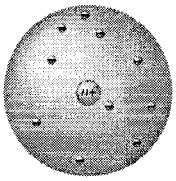
Section 11.1

Atoms and Energy

A. Rutherford's Atom

Nuclear Model of the Atom

- The atom has a small dense nucleus which
  - is positively charged
  - contains protons (+1 charge)
  - contains neutrons (no charge)
- The remainder of the atom
  - Is mostly empty space
  - Contains electrons (–1 charge)

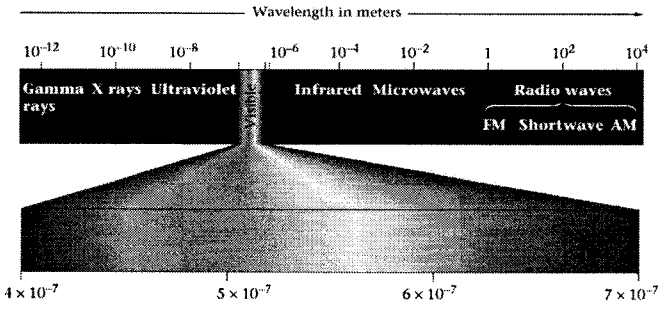


Section 11.1

Atoms and Energy

B. Energy and Light

- Electromagnetic radiation

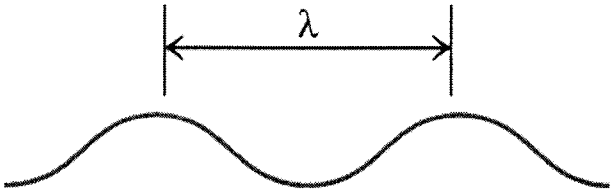


Section 11.1

Atoms and Energy

B. Energy and Light

- How are the types of light different?
  - Wavelength, λ



Section 11.1

Atoms and Energy

B. Energy and Light

- How are the types of light different?
  - Frequency, u
  - Speed, v

Section 11.1

Atoms and Energy

B. Energy and Light

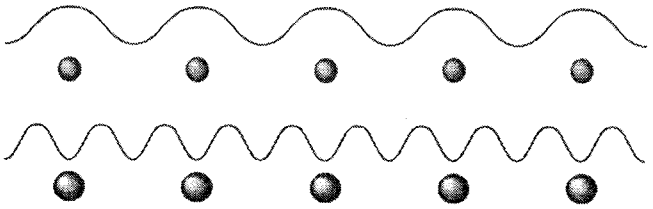
- Dual nature of light
  - Wave
  - Photon – packet of energy

Light as a wave

Light as a stream of photons (packets of energy)

B. Energy and Light

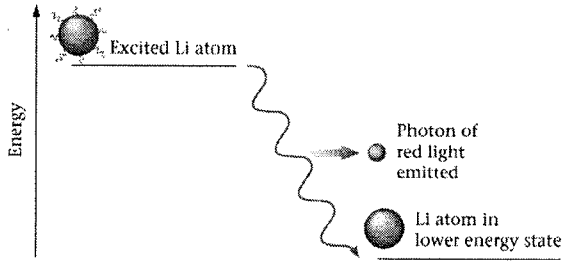
- Different wavelengths carry different amounts of energy.



Section 11.1
Atoms and Energy

C. Emission of Energy by Atoms

- Atoms can give off light.
  - They first must receive energy and become excited.
  - The energy is released in the form of a photon.



Section 11.2

The Hydrogen Atom

Objectives

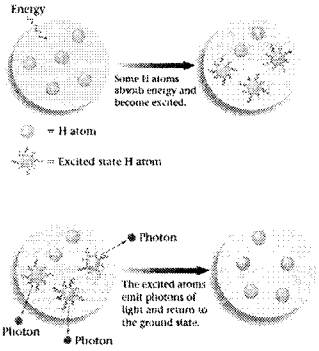
1. To understand how the emission spectrum of hydrogen demonstrates the quantized nature of energy
2. To learn about Bohr's model of the hydrogen atom
3. To understand how the electron's position is represented in the wave mechanical model

Section 11.2

The Hydrogen Atom

A. The Energy Levels of Hydrogen

- Atomic states
  - Excited state – atom with excess energy
  - Ground state – atom in the lowest possible state
- When an H atom absorbs energy from an outside source it enters an excited state.

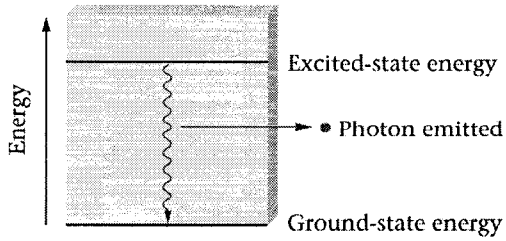


Section 11.2

The Hydrogen Atom

A. The Energy Levels of Hydrogen

- Energy level diagram



- Energy in the photon corresponds to the energy used by the atom to get to the excited state.

Section 11.2

The Hydrogen Atom

A. The Energy Levels of Hydrogen

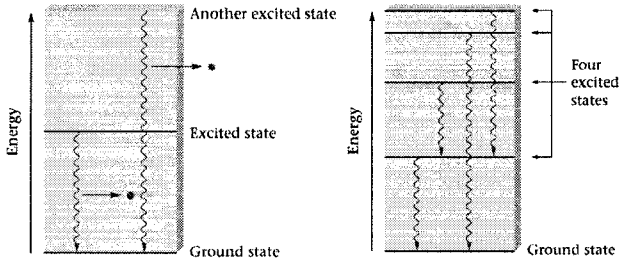
- Only certain types of photons are produced when H atoms release energy. Why?

Section 11.2

The Hydrogen Atom

A. The Energy Levels of Hydrogen

- Quantized Energy Levels
  - Since only certain energy changes occur the H atom must contain discrete energy levels.

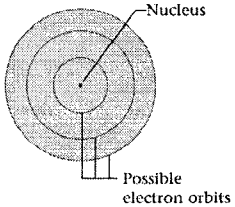


Section 11.2

The Hydrogen Atom

B. The Bohr Model of the Atom

- Bohr's model of the atom
  - Quantized energy levels
  - Electron moves in a circular orbit
  - Electron jumps between levels by absorbing or emitting photon of a particular wavelength



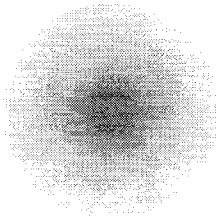
B. The Bohr Model of the Atom

- Bohr' s model of the atom was incorrect.
  - Electron does not move in a circular orbit.

Section 11.2
The Hydrogen Atom

C. The Wave Mechanical Model of the Atom

- Orbitals
  - Nothing like orbits
  - Probability of finding the electron within a certain space



Section 11.3

Atomic Orbitals

Objectives

1. To learn about the shapes of the *s*, *p* and *d* orbitals
2. To review the energy levels and orbitals of the wave mechanical model of the atom
3. To learn about electron spin

Section 11.3

Atomic Orbitals

A. The Hydrogen Orbitals

- Orbitals do not have sharp boundaries.

(a) (b)

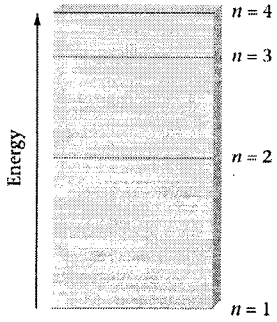
Section 11.3

Atomic Orbitals

A. The Hydrogen Orbitals

Hydrogen Energy Levels

- Hydrogen has discrete energy levels.
  - Called principal energy levels
  - Labeled with whole numbers



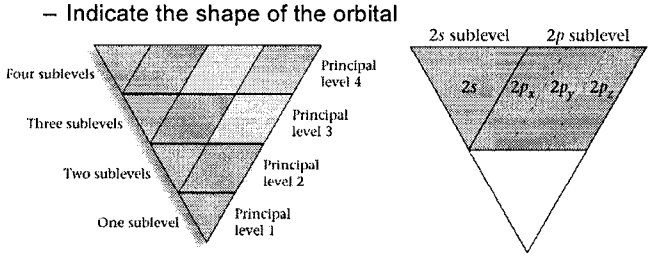
Section 11.3

Atomic Orbitals

A. The Hydrogen Orbitals

Hydrogen Energy Levels

- Each principal energy level is divided into sublevels.
  - Labeled with numbers and letters
  - Indicate the shape of the orbital



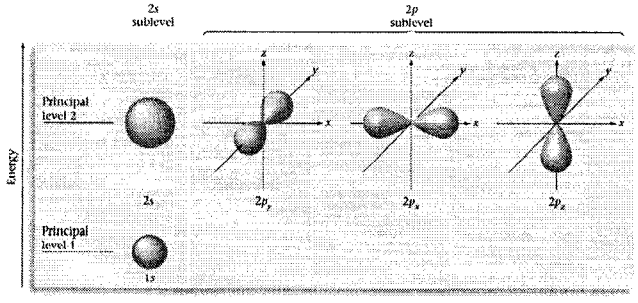
Section 11.3

Atomic Orbitals

A. The Hydrogen Orbitals

Hydrogen Energy Levels

- The *s* and *p* types of sublevel



Section 11.3

Atomic Orbitals

A. The Hydrogen Orbitals

Hydrogen Orbitals

- Why does an H atom have so many orbitals and only 1 electron?
  - An orbital is a potential space for an electron.
  - Atoms can have many potential orbitals.

B. The Wave Mechanical Model: Further Development

Atoms Beyond Hydrogen

- The Bohr model was discarded because it does not apply to all atoms.
- Atoms beyond hydrogen have an equal number of protons and electrons.
  - Need one more property to determine how the electrons are arranged
  - Spin – electron spins like a top

Section 11.3
Atomic Orbitals

B. The Wave Mechanical Model: Further Development

Atoms Beyond Hydrogen

- Pauli Exclusion Principle - an atomic orbital can hold a maximum of 2 electrons and those 2 electrons must have opposite spins

Section 11.4

Electron Configurations and Atomic Properties

Objectives

1. To understand how the principal energy levels fill with electrons in atoms beyond hydrogen
2. To learn about valence electrons and core electrons
3. To learn about the electron configurations of atoms with Z < 18
4. To understand the general trends in properties in the periodic table

Section 11.4

Electron Configurations and Atomic Properties

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

- H atom
  - Electron configuration – electron arrangement – $1s^1$
  - Orbital diagram – orbital is a box grouped by sublevel containing arrow(s) to represent electrons

H: $1s^1$ (Configuration) $\quad$ $1s$ ↑ (Orbital diagram)

Number of electrons in the orbital → $1s^1$ ← Value of n (principal energy level); Type (shape) of orbital

Section 11.4

Electron Configurations and Atomic Properties

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

- He atom
  - Electron configuration– $1s^2$
  - Orbital diagram

Two electrons in 1s orbital

He: $1s^2$ $\quad$ $1s$ ↑↓

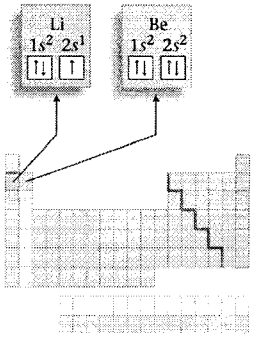
Section 11.4

Electron Configurations and Atomic Properties

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

- Li atom
  - Electron configuration– $1s^2\,2s^1$
  - Orbital diagram

Li: $1s^2 2s^1$ $\quad$ $1s$ ↑↓ $\quad$ $2s$ ↑



Section 11.4

Electron Configurations and Atomic Properties

A. Electron Arrangements in the First 18 Atoms on the Periodic Table

H $1s^1$							He $1s^2$
Li $2s^1$	Be $2s^2$		B $2p^1$	C $2p^2$	N $2p^3$	O $2p^4$	F $2p^5$	Ne $2p^6$
Na $3s^1$	Mg $3s^2$		Al $3p^1$	Si $3p^2$	P $3p^3$	S $3p^4$	Cl $3p^5$	Ar $3p^6$

Section 11.4

Electron Configurations and Atomic Properties

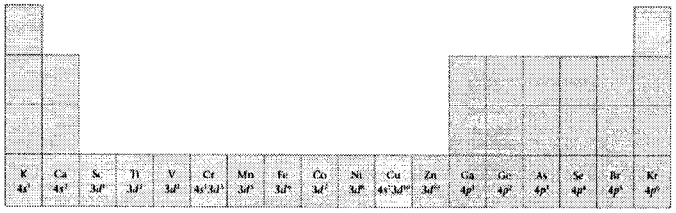
A. Electron Arrangements in the First 18 Atoms on the Periodic Table

Classifying Electrons

- Valence electrons – electrons in the outermost (highest) principal energy level of an atom
- Core electrons – inner electrons
- Elements with the same valence electron arrangement show very similar chemical behavior.

B. Electron Configurations and the Periodic Table

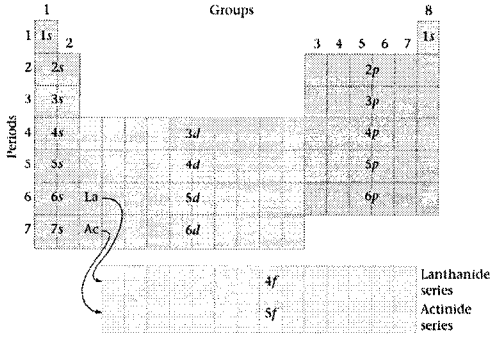
- Look at electron configurations for K through Kr



Section 11.4
Electron Configurations and Atomic Properties

B. Electron Configurations and the Periodic Table

- Orbital filling and the periodic table



Section 11.4
Electron Configurations and Atomic Properties

B. Electron Configurations and the Periodic Table

Periodic Table and Electron Configurations

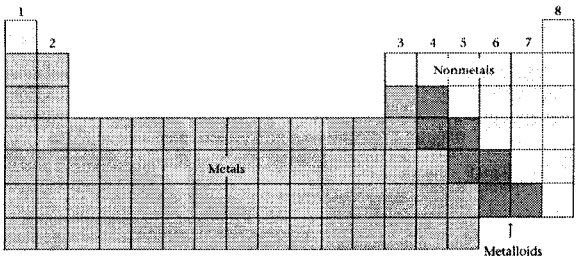
1. The group labels for Groups 1, 2, 3, 4, 5, 6, 7, and 8 indicate the *total number* of valence electrons for the atoms in these groups. For example, all the elements in Group 5 have the configuration ns^2np^3. (Any *d* electrons present are always in the next lower principal energy level than the valence electrons and so are not counted as valence electrons.)
2. The elements in Groups 1, 2, 3, 4, 5, 6, 7, and 8 are often called the **main-group elements**, or **representative elements**. Remember that every member of a given group (except for helium) has the same valence-electron configuration, except that the electrons are in different principal energy levels.

Section 11.4
Electron Configurations and Atomic Properties

C. Atomic Properties and the Periodic Table

Metals and Nonmetals

- Metals tend to lose electrons to form positive ions.
- Nonmetals tend to gain electrons to form negative ions.

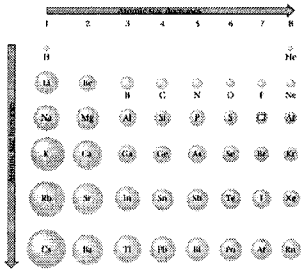


Section 11.4
Electron Configurations and Atomic Properties

C. Atomic Properties and the Periodic Table

Atomic Size

- Size tends to increase down a column.
- Size tends to decrease across a row.



Section 11.4
Electron Configurations and Atomic Properties

C. Atomic Properties and the Periodic Table

Ionization Energies

- Ionization Energy – energy required to remove an electron from an individual atom (gas)

$$M(g) \Rightarrow M^+(g) + e^-$$

ionization energy

- Tends to decrease down a column
- Tends to increase across a row