CHAPTER 5: MODELS OF THE ATOM

Problems: 1-2, 5-9, 11-22, 25-26, 37-48, 61-62, 67, 69-73, 77-80 (a,b only), 81-87, 89-90, 97

5.1 DALTON'S ATOMIC THEORY

* 1. An element is composed of tiny, indivisible, indestructible particles called atoms.
* 2. All atoms of an element are identical and have the same properties.
3. Atoms of different elements combine to form compounds.
4. Compounds contain atoms in small whole number ratios.
   – e.g. each water molecule (H_2O) has one O and 2 H atoms
5. Atoms can combine to form different compounds.
   – e.g. carbon and oxygen can combine to form CO_2 or CO

* Later proven wrong

5.2 THOMSON'S MODEL OF THE ATOM

William Crookes (late 1870s)
– cathode rays are composed of tiny, *negatively* charged subatomic

→ **electrons (e^-)**

J.J. Thomson was given credit for discovery of electron although evidence was accumulated for 20 years before his group's discoveries.

Eugen Goldstein (late 1880s)
– canal rays are composed of *positively* charged subatomic particle

→ **protons (p^+)**

Plum-pudding Model of the Atom
– J.J. Thomson proposed a model where electrons are uniformly dispersed in homogeneous positively, charged spheres
– similar to raisins in English plum pudding (See Fig. 5.1 on p. 108)

And decades later:
– James Chadwick (1935), Nobel Prize winner for his discovery

→ **neutron (n) =** *neutral* subatomic particle

5.3 RUTHERFORD'S MODEL OF THE ATOM

Rutherford's Alpha-Scattering Experiment (Fig. 5.2 on p. 109)

- Alpha (α) particles shot at a thin gold foil
- A detector set up around the foil to determine what happens to the α particles
- Most of α particles went straight through, **but** some were deflected and a few bounced back
  - This was surprising since scientists thought the subatomic particles in the atom were evenly spread throughout the atom and each is too small to stop the much larger alpha particle.
    - It was just like a piece of *tissue paper stopping a bullet from a gun*!

Rutherford's interpretation of the experimental results (Fig. 5.3)

- Most alpha (α) particles passed through foil

  ⇒ Atom is mostly empty space with electrons moving around the space
- Some α particles were deflected or bounced back

  ⇒ Atom must also contain a *dense region*, and α particles colliding with this region are deflected or bounce back towards source

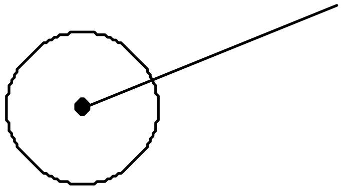
  ⇒ dense region = **atomic nucleus** (contains atom's protons and neutrons)

Rutherford's Planetary Model of the Atom

- *Negatively charged* **e**$^-$ move around the *positively charged* nucleus

Rutherford also estimated the size of the atom and its nucleus:

nucleus (~10^{-13} cm in diameter)



electron orbits within atom (~10^{-8} cm in diameter)

If nucleus = size of a small marble, then atom ≅ Bank One Ballpark (BOB)!

Particle	Symbol	Location	Charge	Relative Mass
electron	e^-	outside nucleus	–1	1/1836 ≈ 0
proton	p^+	inside nucleus	+1	1
neutron	n	inside nucleus	0	1

5.4 ATOMIC NOTATION

Isotopes: Elements always have the same number of protons, but the number of neutrons may vary. Atoms with differing numbers of neutrons are called **isotopes**.

- The convention for distinguishing elements with various isotopes is to give the element name followed by the mass number
  - e.g. carbon-12 (C-12), carbon-13 (C-13) and carbon-14 (C-14) are all isotopes of carbon

Atomic Notation (also called "Nuclear Symbol"):

- shorthand for keeping track of protons and neutrons in the nucleus

atomic number: whole number of p^+ = number of e^- in neutral atom

mass number: whole number sum of p^+ and n in an atom's nucleus

$${}^{\text{mass number}=\mathbf{A}}_{\text{atomic number}=\mathbf{Z}}\mathrm{E} = \text{element symbol}$$

Ex. 1: a. Write the atomic notation for sodium-22 below.

b. How many neutrons are in each neutral sodium-22 atom? ______

Ex. 2: a. Write the atomic notation for chlorine-37 below.

b. How many neutrons are in each neutral chlorine-37 atom? ______

Ex. 3: a. Write the atomic notation for nitrogen-15 below.

b. How many neutrons are in each neutral nitrogen-15 atom? ______

Ex. 4: Indicate the mass number, number of protons, neutrons, and electrons in the following isotopes of carbon:

Isotope of carbon	mass #	# of protons	# of neutrons	# of electrons
carbon-12				
carbon-13				
carbon-14				

5.5 ATOMIC MASS

Atoms are too small to weigh directly

– eg. one carbon atom has a mass of 1.99 x 10^{-23} g—too inconvenient!

$\Rightarrow$ need more convenient unit for mass

$\Rightarrow$ **atomic mass unit (amu)**

Carbon-12 was chosen as the **reference** and given a mass value of **12 amu**

$\Rightarrow$ 1 amu = 1/12 the mass of carbon-12

$\Rightarrow$ Mass of all other atoms measured relative to mass of carbon-12

Atomic Mass of an Element is ***weighted average*** of all naturally occurring isotopes for that element.

- There are two naturally occurring isotopes of carbon: C-12 and C-13
- More carbon exists as carbon-12 (98.89%) compared with carbon-13 (1.11%), so the atomic mass reported for carbon (12.01 amu) is closer to carbon-12.

Example: Use the atomic weight reported on the Periodic Table for the different elements to determine the most abundant naturally occurring isotope for each of the following:

a. The two naturally occurring isotopes for lithium are: (Circle one)

lithium-6 lithium-7

b. The two naturally occurring isotopes for chlorine are: (Circle one)

chlorine-35 chlorine-37

Some elements are ***radioactive*** and ***unstable***

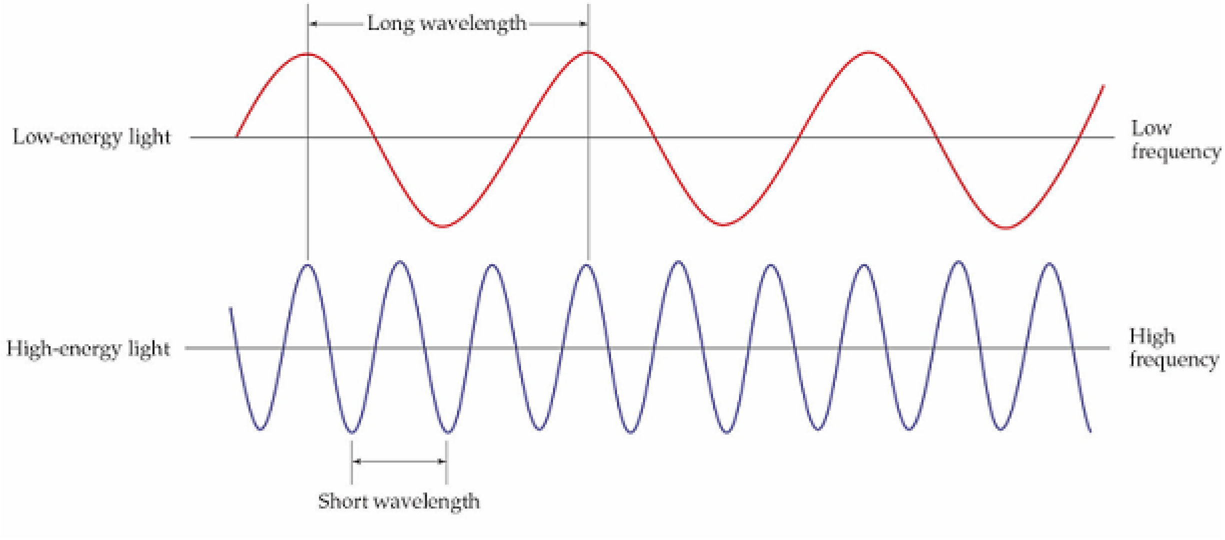
⇒ distinguished on the Periodic Table with parentheses around their ***mass numbers*** (instead of their ***atomic weights*** reported)

- e.g. Element #85, astatine (At), has a mass number of 210, and element #96, curium (Cm), has a mass number of 247.

5.6 THE WAVE NATURE OF LIGHT

Light has two components: **wavelength** and **frequency**. (See Fig. 5.10)

- **Wavelength** is the distance between peaks on adjacent waves
  - usually represented with the Greek letter lambda (λ)
- **Frequency** refers to the number of wave cycles completed in one second
  - usually represented with the Greek letter nu (ν)
- As wavelength ↑, the frequency ↓**,** and the energy ↓
- As wavelength ↓, the frequency ↑**,** and the energy ↑



Light—A Continuous Spectrum

Radiant Energy Spectrum (Fig. 5.12)

- Continuous spectrum from gamma rays to radio waves
- Note: The longer the wavelength ⇒ the lower the energy

Light we observe with the naked eye falls within the ***visible spectrum***.

- Continuous from violet (400 nm) to red (700nm).
- Blue flames have a shorter wavelength than yellow flames
  - → Blue flames are higher in energy and hotter!

5.7 THE QUANTUM CONCEPT

In 1900, Max Planck proposed the controversial idea that **energy was emitted in small bundles (quanta), not continuously.**
→ proposal that **light is also emitted in small bundles (photons)**

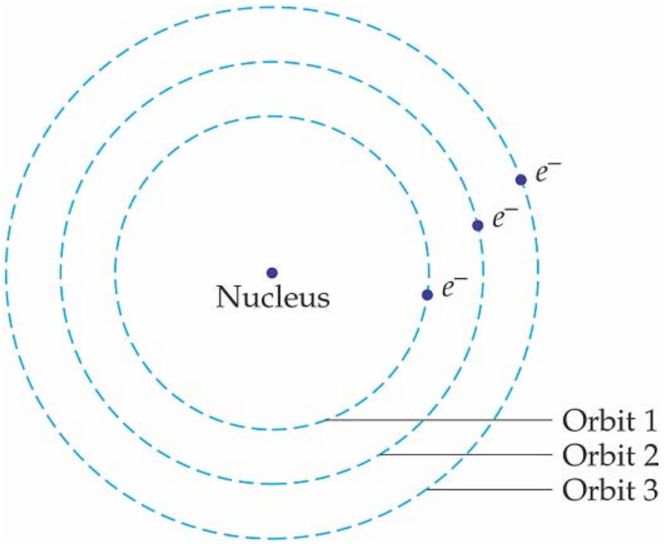
5.8 BOHR MODEL OF THE ATOM

In 1913, Neils Bohr speculated that in the atom, electrons revolve around the nucleus, occupying circular orbits with distinct **energy levels**.

5.9 ENERGY LEVELS AND SUBLEVELS

Bohr's model of the atom

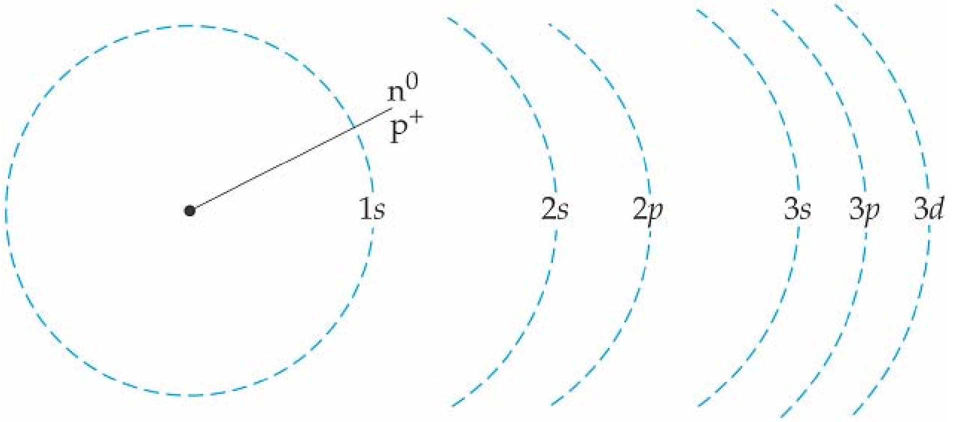
- The electrons orbit around the nucleus like planets orbit around the sun.
- Each orbit has a specific energy. The orbits closest to the nucleus are the lowest in energy, and energy increases with distance from the nucleus.
- A hydrogen atom contains only one proton and one electron, so these energy levels are simply numbered (e.g. 1, 2, 3,…)



For all other elements (w/ more than 1 proton and more than 1 electron), **principal energy levels** (numbered 1, 2, 3,…) are further divided into **energy sublevels**.

principal energy level (n): n=1,2,3,...

energy sublevels: s, p, d, and f



Evidence for Energy Levels

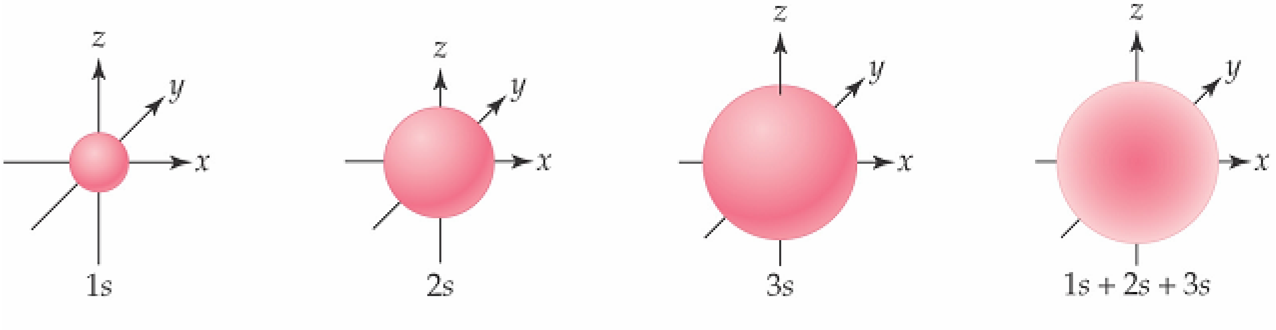
A gas is sealed in a gas discharge tube and energized by electricity

- → electrons gain energy and jump to higher energy levels, but ultimately fall to lowest energy level (or ground state), see Fig. 5.13 on p.121
  - → light separated into narrow bands called an **emission line spectrum**
    - – also called an **"atomic fingerprint"** since each element emits its own unique spectrum, which can be used identify the element

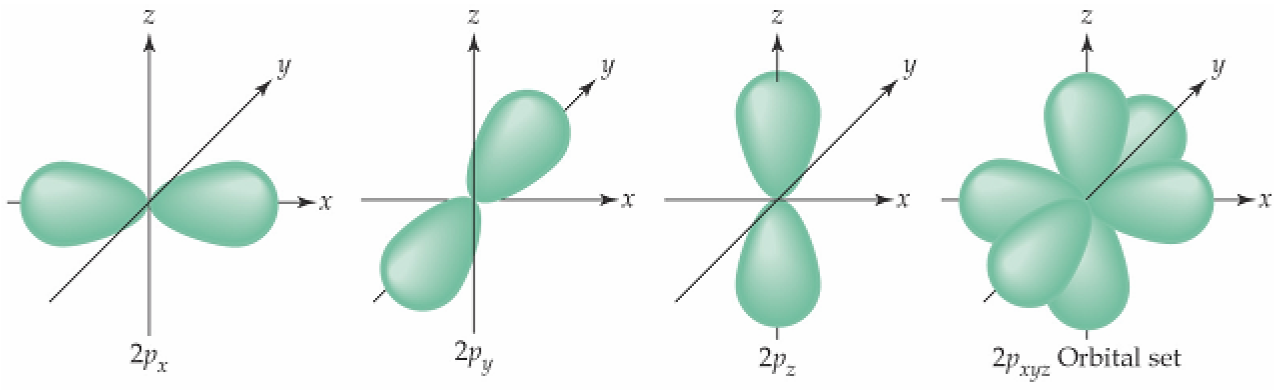
5.11 QUANTUM MECHANICAL MODEL OF THE ATOM

In reality, the electron does not move in a fixed orbit. Instead, we talk about the electron having a high probability of being found within a given **volume** corresponding to the orbital and its energy.

- → **Sizes and Shapes of Orbitals**
  - – **S orbitals are spherical** (see Fig. 5.17)
    - – As principal number ↑, the size and energy of the orbitals ↑



- – **P orbitals resemble dumbbells**, lying along the x,y, and z axes (see Fig. 5.17)
  - – Putting all three p orbitals together results in an orbital set that is close to spherical in shape



Note: You do not need to know about the d orbitals.

5.10 ELECTRON CONFIGURATION

1. Electrons are distributed in orbitals of increasing energy levels, where the lowest energy orbitals are filled first.

2. Once an orbital has the maximum number of electrons it can hold, it is considered “filled.” Remaining electrons must then be placed into the next highest energy orbital, and so on.

   Orbitals in order of increasing energy: (See Corwin p. 125, Fig. 5.16)

   1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p < 6s < 5d < 6p

ELECTRON CONFIGURATION:

– Shorthand description of the arrangement of electrons by sublevel according to increasing energy

REMEMBER!

— ***s orbitals*** can hold **2 electrons**
— a set of ***p orbitals*** can hold **6 electrons**
— a set of ***d orbitals*** can hold **10 electrons**
— a set of ***f orbitals*** can hold **14 electrons**

Ex. 1 He → atomic #=2 ⇒ _____ e-

electron configuration for He: ______________________________

Ex. 2 C → _____ e-

electron configuration for C: ______________________________

Ex. 3 S → _____ e-

electron configuration for S: ______________________________

Ex. 4 K → _____ e-

electron configuration for K: ______________________________

6.6 BLOCKS OF ELEMENTS

The Periodic Table actually corresponds to the order of energy sublevels.

– See Fig. 6.6 on p. 148 to see how electrons for each element are distributed into energy sublevels.

Electron configurations of atoms with many electrons can become cumbersome.

→ an abbreviation called **core notation** using Noble gas electron configurations

- Elements in the last column of the Periodic Table are called "noble gases."
- Since noble gases are at the end of each row in the Periodic Table, all of their electrons are in filled orbitals.

[He] = $1s^2$
[Ne] = $1s^2\ 2s^2\ 2p^6$
[Ar] = $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

– Such electrons are called "**core electrons**" since they are more stable (less reactive) when they belong to completely filled orbitals.

→ Noble gas electron configurations can be used to abbreviate the "core electrons" of all elements

→ **Electron configurations using Noble gas abbreviations** are often called "**core notation**"

Electron Configurations using Core Notation:

a. Electron configuration for K using full notation: $\mathbf{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}\ 4s^1$

Electron configuration for K using core notation: **[Ar]** $4s^1$

b. Electron configuration for Cl using full notation: ________________________

Electron configuration for Cl using core notation: ________________________

c. Electron configuration for Mg using full notation: _______________________

Electron configuration for Mg using core notation: _______________________

d. Electron configuration for Si using full notation: _______________________

Electron configuration for Si using core notation: _______________________

Note: Be able to write electron configurations for elements #1-20.