Notes: ATOMS AND THE PERIODIC TABLE

GOOD WORK

Atomic Structure:

- Atom: the smallest particle that has the properties of an element.
- From the early Greek concept of the atom to the modern atomic theory, scientists have built on and modified existing Models of the atom.

Atom Basics:

·Atoms are composed of a positively charged nucleus surrounded by an electron cloud.

- Nucleus (99% of atom's mass): uncharged neutrons and positively charged protons.
- Electron cloud: negatively charged electrons in constant motion creating a "cloud" like a fan.

DEMOCRITUS:

- In 400 B.C., this Greek philosopher suggested that the universe was made of indivisible units.
- "Atom" – Greek word meaning "unable to be divided".

JOHN DALTON:

Dalton's Atomic Theory:

- ~~in 1808 he proposed his atomic~~ All elements are made of tiny atoms
- Atoms cannot be subdivided
- Atoms of the same element are exactly alike
- Atoms of different elements can join to form molecules

THOMPSON AND MILLIKAN:

·As it turns out, the atom can be divided into subatomic particles.

·**Thompson** and **Millikan** are given credit for the first discoveries relating to electrons.

RUTHERFORD:

·**Rutherford** discovered the positively charged nuclei.

NIELS BOHR:

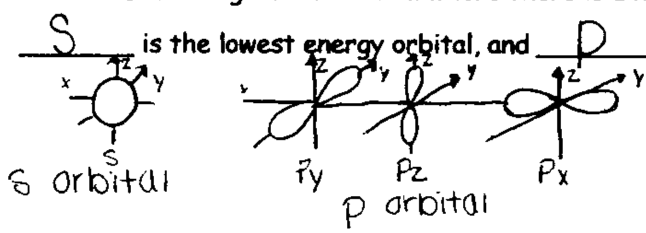
- In 1913, this Danish scientist suggested that electrons "orbit" the nucleus.
- In Bohr's model, electrons are placed in different energy levels based on their distance from the nucleus.

MODERN ATOMIC MODEL:

- By 1925, Bohr's model of the atom no longer explained all observations. Bohr was correct about energy levels, but wrong about electron movement.
- Electrons occupy the lowest energy levels available.
- Energy increase as distance from the nucleus increases.
- Electrons move in patterns of "wave functions" around the nucleus.
- It is impossible to know both an electrons velocity and location at any moment in time.

ORBITALS:

- ORBITAL: the regions in an atom where there is a high probability of finding electrons.
- S is the lowest energy orbital, and P is slightly higher



- d and f are the next two orbitals. They occupy even higher energy levels and take on more complex shapes than *s* & *p*

VALENCE ELECTRONS:

- Electrons in the outermost energy level are called Valence electrons.
- Valence electrons determine how an atom will act in a chemical reaction.
- Atoms with equal numbers of valence electrons have similar properties.

DMITRI MENDELEEV: 1834-1907

1869: created first periodic table of elements.

Arranged elements in order of increasing atomic mass.

HENRY MOSELY: 1887 - 1915

One of Rutherford's students.

1914: Arranged the elements in order of increasing atomic number (responsible for TODAY'S periodic table).

ORGANIZATION OF THE PERIODIC TABLE:

PERIODICITY: regular variations (or patterns) of properties with increasing atomic number. Both chemical and physical properties vary in a periodic (repeating) pattern.

- Period: horizontal row of elements on P.T.
- Group (Family): vertical column of elements on P.T.

PERIODIC KEY:

6	Atomic # (Z)
C	Element's symbol
Carbon	Element's name
12.011	Atomic mass (A)

protons = Z

electrons = # of protons (in a neutral atom)

neutrons = A - Z

ISOTOPES

- Isotopes are atoms that have the same # of protons, but a different # of neutrons.
- Example: Carbon-12 vs. Carbon-14

^{12}C Mass # = 12; Atomic # = 6 (6 P, 6 E, 6 N)

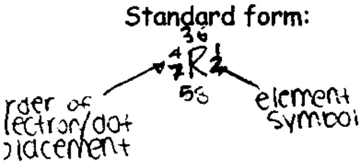
^{14}C Mass # = 14; Atomic # = 6 (6 P, 6 E, 8 N)

IONS

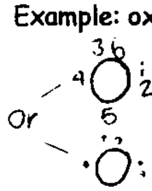
- Ionization: the process of adding or removing electrons from an atom or group of atoms.
- An ion has a net electric charge.
- Cation: ion with a positive charge. Ex: Na^+
- Anion: ion with a negative charge. Ex: O^{2-}

ELECTRON DOT DIAGRAMS: (diagram of valence electrons)

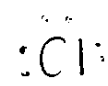
Standard form:



Example: oxygen



Example: chlorine



DETERMINING # OF PROTONS, NEUTRONS, AND ELECTRONS FROM CHEMICAL SYMBOLS:

Example 1:
protons = 6
#electrons= 6
#neutrons= 8

Mass # No net charge

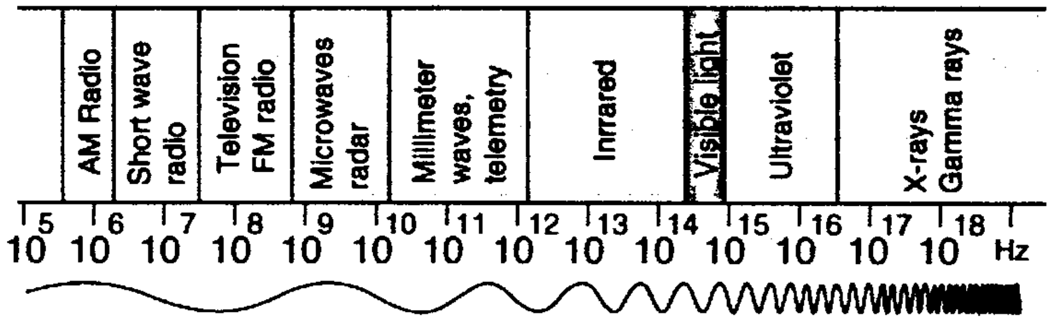
$^{14}_{6}C$

Atomic #

Example 2:
protons = 7
electrons = 10
neutrons = 8

$^{15}_{7}N^{3-}$ ← Net charge of ~~negative~~ -3

Notes: Light, Photon Energies, and Atomic Spectra



Low frequency
Long wavelength
Low quantum energy

High frequency
Short wavelength
High quantum energy

- Electromagnetic radiation (radiant energy) is characterized by its:
  - **-wavelength (color):** λ (Greek letter lambda)
  - **-frequency (energy):** ν (Greek letter nu)

- They are related by the equation:

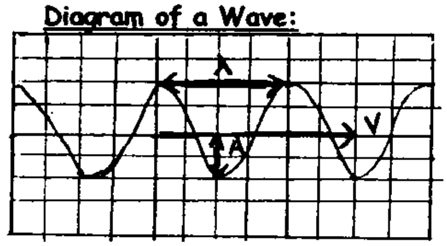
$$\boxed{\nu = \frac{c}{\lambda}}$$

where c = 3.00 x 10^8 m/s (the speed of light in a vacuum)

Wavelength = distance between successive "crests"

Frequency = the # of crests passing a given point per second

Diagram of a Wave:



Example: The frequency of violet light is 7.31 x 10^{14} Hz, and that of red light is 4.57 x 10^{14} Hz. Calculate the wavelength of each color.

Violet: $7.31 \times 10^{14}\ \text{Hz} = \dfrac{3.00 \times 10^8 \frac{m}{s}}{\lambda}$

$\lambda = \dfrac{3.00 \times 10^8 \frac{m}{s}}{7.31 \times 10^{14} \frac{1}{s}}$ $\lambda = 4.10 \times 10^{-7}\ m$

$4.57 \times 10^{14}\ \text{Hz} = \dfrac{3.00 \times 10^8 \frac{m}{s}}{\lambda}$

$\lambda = \dfrac{3.00 \times 10^8 \frac{m}{s}}{4.57 \times 10^{14} \frac{1}{s}}$ $\lambda = 6.56 \times 10^{-7}\ m$

- When sunlight or white light is passed through a prism, it gives the **continuous spectrum** observed in a rainbow.

- We can describe light as composed of particles, or **PHOTONS**.
- Each photon of light has a particular amount of energy (a **quantum**).
- The amt. of energy possessed by a photon depends on the color of the light.

- The energy of a photon is given by this equation:

$$\boxed{E = h\nu}$$

where h = 6.6262 x 10^{-34} J·s
and ν = frequency (Hz)

Example: Calculate the energy, in joules, of an individual photon of violet and red light.

violet: $E = (6.6262 \times 10^{-34}\ J\cdot s)(7.31 \times 10^{14} Hz)$
$E = (6.6262 \times 10^{-34}\ J\cdot s)(7.31 \times 10^{14} \frac{1}{s})$
$E = 4.84 \times 10^{-19}\ J$

red: $E = (6.6262 \times 10^{-34}\ J\cdot s)(4.57 \times 10^{14} Hz)$
$E = \cancel{(6.6262 \times 10^{-34}\ J}\cdot s)(4.57 \times 10^{14} \frac{1}{s})$
$E = 3.03 \times 10^{-19}\ J$

What does this have to do with electron arrangement in atoms?

• When all electrons are in the lowest possible energy levels, an atom is said to be in its **GROUND STATE**.

• When an atom absorbs energy so that its electrons are "boosted" to higher energy levels, the atom is said to be in an **EXCITED STATE.**

• The light emitted by an element when its electrons return to a lower energy state can be viewed as a **bright line emission spectrum**. (see figure 6.3 on page 147)

• The light absorbed by an element when white light is passed through a sample is illustrated by the **absorption spectrum**.

Note: The wavelengths of light that are absorbed by the gas show up as black lines, and are equal to the wavelengths of light given off in the emission spectrum.

Why?
• Electronic energy is **quantized** (only certain values of electron energy are possible).
• When an electron moves from a lower energy level to a higher energy level in an atom, energy of a characteristic frequency (wavelength) is **absorbed**.

• When an electron falls from a higher energy level back to the lower energy level, then radiation of the same frequency (wavelength) is **emitted**.

• The bright-line emission spectrum is unique to each element, just like a fingerprint is unique to each person. *see figure 6.3, p. 147 – Harcourt text (honors only)

Example: A green line of wavelength 486 nm is observed in the emission spectrum of hydrogen. Calculate the energy of one photon of this green light.

Green: $E = \frac{hc}{\lambda} = \frac{(6.6262 \times 10^{-34}\ J\cdot s)(3.00 \times 10^{8} \frac{m}{s})}{486 \times 10^{-9}\ m}$ $E = 4.09 \times 10^{-19}\ J$

Example: The green light associated with the aurora borealis is emitted by excited (high-energy) oxygen atoms at 557.7 nm. What is the frequency of this light?

Green: $V = \frac{3.00 \times 10^{8} \frac{m}{s}}{557.7 \times 10^{-9}\ m}$

$V = 5.38 \times 10^{14}\ Hz$

Notes: Electron Configurations

• The quantum mechanical model of the atom predicts energy levels for electrons; it is concerned with the probability, or likelihood, of finding an electron in a certain position.

• Regions where electrons are likely to be found are called orbitals.
EACH ORBITAL CAN HOLD UP TO 2 ELECTRONS!

• In quantum theory, each electron is assigned a set of quantum numbers
(*analogy: like the mailing address of an electron)

1) Principal Quantum Number (n):
• describes the energy level that the electron occupies
• n = 1, 2, 3, 4
• the larger the value of n, the farther away from the nucleus and the higher the energy of the electron.

2) Sublevels (l):
• the # of sublevels in each energy level = the quantum #, n, for that energy level.
• sublevels are labeled with a # that is the principal quantum #, and a letter: s, p, d, f
(ex: 2p is the p sublevel in the 2nd energy level)

Principal Energy Level	Sublevels	Orbitals
n=1	1s	one (1s)
n=2	2s 2p	one (2s) three (2p)
n=3	3s 3p 3d	one (3s) three (3p) five (3d)
n=4	4s 4p 4d 4f	one(4s) three(4p) five (4d) seven (4f)

Sublevel	# of orbitals	Max. # of electrons
s	1	2
p	3	6
d	5	10
f	7	14

3) spin quantum number (m_s):
• labels the orientation of the electron;
• electrons in an orbital spin in opposite directions; these directions are designated as $+\frac{1}{2}$ and $-\frac{1}{2}$

Pauli Exclusion Principle: states that no 2 electrons have an identical set of four quantum #'s; ensures that no more than 2 electrons can be found within a particular orbital.

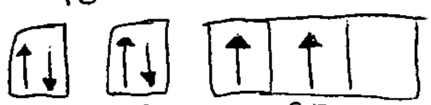
Hund's rule: orbitals of equal energy are each occupied by one electron before any pairing occurs. (repulsion between electrons in a single orbital is minimized)

All electrons in singly occupied orbitals must have the same spin; when 2 electrons occupy an orbital they have opposite spins.

Orbital diagrams:
-each orbital is represented by a box
-each electron is represented by an arrow

hydrogen: 

helium: 

carbon: 

Electron configurations: an abbreviated form of the orbital diagram.

helium: $1s^2$

boron: $1s^2 2s^2 2p^1$

neon: $1s^2 2s^2 2p^6$

aluminum: $1s^2 2s^2 2p^6 3s^2 3p^1$

uranium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6 7s^2 5f^4$

Abbreviated electron configurations: an abbreviated form of the electron configuration.

helium: $1s^2$

boron: $[He] 2s^2 2p^1$

aluminum: $[Ne] 3s^2 3p^1$

cobalt: $[Ar] 4s^2 3d^7$

uranium: $[Rn] 7s^2 5f^4$

N^{3-}: $[He] 2s^2 2p^6 = [Ne]$

Se^{2-}: $[Kr]$

Mg^{2+}: $[Ne]$

Notes: Periodic Groups and Trends

PERIODIC GROUPS:

Alkali Metals

- Group 1 on the periodic table.
- very reactive
- Soft Solids
- Readily combine with halogens
- Tendency to lose one electron

Alkaline Earth Metals

- Group 2 on the periodic table.
- Abundant metals in the earth
- Not as reactive as alkali metals
- Higher density and melting point than alkali metals

Transition Metals

- Groups 3-12 on the periodic table.
- Important for living organisms

Halogens

- Group 17 on the periodic table.
- "Salt former" combines with groups 1 and 2 to form salts (ionic bonds)

Noble Gases

- Group 18 on the periodic table.
- Relatively inert, or nonreactive
- Gases at room temperature

Lanthanides

- Part of the "inner transitiom metals"
- Sofl silvery metals
- Tarnish readily in air
- React slowly with water

Actinides

- Radioactive elements
- Part of the "inner transition metals"

PERIODIC TRENDS

1) Atomic Radii:

- *Trend:* increases down a group
- **Why?**

  The atomic radius gets bigger because electrons are added to energy levels farther away from the nucleus.

  PLUS, the inner electrons shield the outer electrons from the positive charge ("pull") of the nucleus; this is known as the ***SHIELDING EFFECT.***

- *Trend:* decreases across a period
- **Why?**

  As the # of protons in the nucleus increases, the positive charge, and as a result, the "pull" on the electrons, increases.

2) Ionization Energy: energy required to remove an outer electron

- *Trend:* decreases down a group
- **Why?**

  Electrons are in a higher energy levels as you move down a group; they are further away from the nucleus, and thus easier to remove.

- *Trend:* increases across a period
- **Why?**

  The increasing charge in the nucleus as you move across a period exerts greater "pull" on the electrons; it requires more energy to remove an electron.

3) Ionic Radii

- Cations are always smaller than the metal atoms from which they are formed. (fewer electrons)
- Anions are always larger than the nonmetal atoms from which they are formed. (more electrons)

4) Electronegativity: the tendency of an atom to attract electrons to itself when chemically combined with another element

- *Trend:* decreases down a group
- **Why?**

  Although the nuclear charge is increasing, the larger size produced by the added energy levels means the electrons are farther away from the nucleus; decreased attraction, so decreased electronegativity; plus shielding effect.

- *Trend:* increases across a period (noble gases excluded)
- ***Why?***

  Nuclear charge is increasing, atomic radius is decreasing, so the attractive force that the nucleus can exert on another electron increases.