Notes: ATOMS AND THE PERIODIC TABLE

Atomic Structure:

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

Atom Basics:

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

–_________________ (99% of atom's mass): uncharged neutrons and positively charged protons.

–__________________________________: negatively charged electrons in constant motion creating a "cloud" like a fan.

DEMOCRITUS:

- In ____________, this Greek philosopher suggested that the universe was made of __________ ______________________________.
- "Atom" – Greek word meaning "____"

JOHN DALTON:

Dalton's Atomic Theory:

– ____

– ____

– ____

– ____

THOMPSON AND MILLIKAN:

•As it turns out, the atom can be divided into ____.

•**Thompson** and **Millikan** are given credit for the first discoveries relating to ________________.

RUTHERFORD:

•**Rutherford** discovered the ____.

NIELS BOHR:

- In 1913, this Danish scientist suggested that electrons _________________________________.
- In Bohr's model, electrons are placed in different ___________________________________ based on their ____.

MODERN ATOMIC MODEL:

- By _________, Bohr's model of the atom no longer explained all observations. Bohr was correct about _______________________, but wrong about _____________________________.
- Electrons occupy the ___________________________________ levels available.
- Energy ____________________ as distance from the nucleus ________________________.
- Electrons move in patterns of "________________________________" around the nucleus.
- It is impossible to know both an electrons ___________________ and ____________________ at any moment in time.

ORBITALS:

- ORBITAL: the regions in an atom where there is a high _____________ of finding electrons.
- _______ is the lowest energy orbital, and __________ is slightly higher
- ____________________ 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 determine how an atom will ____.
- Atoms with equal numbers of valence electrons have __________________________________.

DMITRI MENDELEEV: 1834-1907

_________: created first periodic table of elements.

Arranged elements in order of increasing _________________________________.

HENRY MOSELY: 1887 – 1915

One of ___________________________ students.

____________: Arranged the elements in order of increasing _____________________________ (responsible for TODAY'S ______________________________________).

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.

- ______________________________: horizontal row of elements on P.T.
- ______________________________: vertical column of elements on P.T.

PERIODIC KEY:

protons =

electrons =

neutrons =

ISOTOPES

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

^{12}C Mass # = ____; Atomic # = ____ (___P, ___E, ___N)

^{14}C Mass # = ____; Atomic # = ____ (___P, ___E, ___N)

IONS

- ______________: the process of adding or removing electrons from an atom or group of atoms.
- An ________ has a net __________________________.
- Cation: ion with a ______________________ charge. Ex: _______
- Anion: ion with a ______________________ charge. Ex: _______

ELECTRON DOT DIAGRAMS: (diagram of valence electrons)

Standard form: Example: oxygen Example: chlorine

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

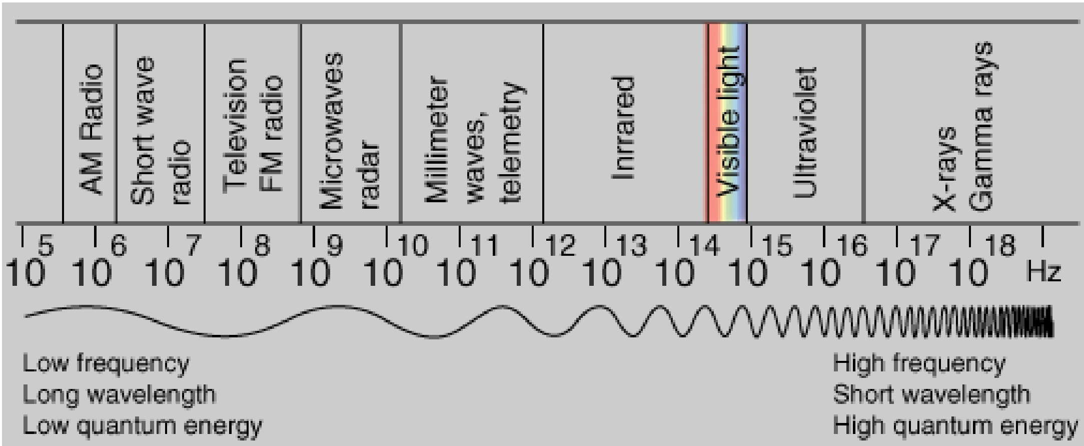
Example 1:
protons = ____
#electrons=____
#neutrons=____

$^{14}_{6}C$

Example 2:
protons = ____
electrons = ____
neutrons = ____

$^{15}_{7}N^{3-}$

Notes: Light, Photon Energies, and Atomic Spectra



• Electromagnetic radiation (radiant energy) is characterized by its:

-wavelength (color):

-frequency (energy):

• They are related by the equation:

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

Wavelength =

Diagram of a Wave:

Frequency =

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.

• 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:

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

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

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.

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?

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 ( ):
• 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 ( ):
• 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 2^{nd} energy level)

Principal Energy Level	Sublevels	Orbitals

Sublevel	# of orbitals	Max. # of electrons

3) spin quantum number ( ):
• 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:

boron:

neon:

aluminum:

uranium:

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

helium: N^{3-} :

boron: Se^{2-} :

aluminum: Mg^{2+} :

cobalt:

uranium:

Notes: Periodic Groups and Trends

PERIODIC GROUPS:

Alkali Metals

- Group ______ on the periodic table.
- ____
- ____
- Readily combine with ___________________________
- Tendency to ____

Alkaline Earth Metals

- Group ______ on the periodic table.
- Abundant metals ____
- Not as reactive as ___________________________________
- Higher ____________________ and ______________________________ than alkali metals

Transition Metals

- Groups ___________ on the periodic table.
- Important for living organisms

Halogens

- Group ______ on the periodic table.
- "________________________" combines with groups ____ and ____ to form salts (ionic bonds)

Noble Gases

- Group ______ on the periodic table.
- ____
- ____

Lanthanides

- Part of the "____"
- ____
- _________________________ readily in air
- React slowly with ______________________

Actinides

- ____
- Part of the "____"

PERIODIC TRENDS

1) Atomic Radii:

- ***Trend:***____
- **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:*** ____
- **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:*** ____
- **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:*** ____
- **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

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

- ***Trend:*** ____
- **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:*** ____
- ***Why?***
  Nuclear charge is increasing, atomic radius is decreasing, so the attractive force that the nucleus can exert on another electron increases.