CHEM 101: CHAPTER 10: MODERN ATOMIC THEORY AND THE PERIODIC TABLE pgs. 194 - 212

ELECTROMAGNETIC RADIATION

This is energy in the form of radio waves, microwaves, infrared radiation, visible light, ultraviolet light (UV), x-rays, and gamma rays.
Our eyes are capable of seeing only in the visible light region (400 -700 nm) of the electromagnetic spectrum. SEE pg. 196 in text.

Wavelength and Energy

As the wavelength (symbol λ) decreases, the energy for that wave increases. This is an indirect relationship. So, as you go from gamma rays (the shortest λ) to radio waves (the longest λ) the electromagnetic radiation has less and less energy.

For example, being exposed to x-rays is much more damaging than being exposed to visible light. X-rays are known to trigger cancerous tumor growth and cause deep tissue damage.

Electromagnetic energy has dual properties of both a particle (termed a photon) and that of a wave. Different situations use either property to explain physical phenomenon such as some metals giving off electrons when exposed to light (termed the photoelectric effect). This is explained by seeing light as a particle.

THE BOHR ATOM

Recall that Niels Bohr was responsible for the Bohr or Planetary Model of the Atom. He put forth a model that electrons are different distances away from the nucleus like planets revolving around the sun.

Bohr developed this model with the aid of past work done by German physicist *Max Planck*. Planck proposed that energy is released in small bundles he termed quanta (for quantity). Bohr used this information along with knowledge of the bright-line spectrum (a display of light in which bright lines represent light of a particular wavelength beingemitted or given off) for hydrogen to explain that this spectrum was a result of electrons jumping up to higher energy levels and then falling back down and emitting their energy they absorbed to get up there as light of a certain wavelength.

THE *LOWEST ENERGY LEVEL* AN ELECTRON CAN FALL TO IS TERMED THE *GROUND STATE*.

As each element has its own unique bright-line spectrum, it serves as a helpful identifier of an unknown element. Astronomers in particular use this to identify elements in stars. This was how HELIUM (from the Greek "helios" for sun) was first discovered in a star (our sun), not on the earth.

Regretfully, since Bohr's model only worked for light but not heavier atoms, more work on atomic structure was done by French physicist *Louis de Broglie*.

De Broglie proposed that all matter has wave properties. For large objects such as humans, these waves are insignificant. However, for small particles like an electron, they become important. With this hypothesis proven by others and the work of Austrian physicist *Erwin Schrodinger*, a mathematical model was developed that not only described electrons as waves but also allowed for the determination of the probability of finding an electron, although not precisely, in a region around the atom.

THESE AREAS OF PROBABILITY ARE TERMED ORBITALS. *These orbitals have differing geometric shapes.*

ENERGY LEVELS OF ELECTRONS AND THE PAULI EXCLUSION PRINCIPLE

The Bohr model aided in developing the wave-mechanical model in which electrons have a principal energy level *n*. *n* is a whole number integer ranging from 1, 2, 3, and on. *N*= 1 is termed the ground state, the lowest possible energy level an electron can occupy.

The principal energy level is further divided into sublevels. These sublevels (represented by the letter *l*) range from 1, 2, 3, and on as the principal levels did and contain spaces for the electrons to occupy termed ORBITALS.

EACH ORBITAL CAN HOLD A MAXIMUM OF 2 ELECTRONS EACH, **AND THESE ELECTRONS MUST HAVE OPPOSITE SPINS. ELECTRONS SPIN ON THEIR AXES AND CAN EITHER SPIN UP (+1/2) OR SPIN DOWN (-1/2). POSITIVE SPIN IS ALSO INDICATED WITH AN ↑ ARROW, WHILE NEGATIVE SPIN A ↓. THIS RULE IS TERMED THE *PAULI EXCLUSION PRINCIPLE.***

Principal Energy Levels and the Number and Name of Possible Orbitals

For n = 1, there is 1 orbital. It is termed an "s" orbital and is written 1s. It can hold a maximum of 2 e-.

For n=2, there are 4 orbitals, one 2s orbital and three 2p orbitals. Each p orbital can also hold only 2 e−, so the 2 from the 2s plus the 6 from the three 2p orbitals gives a maximum of 8 e− in the 2nd principal energy level.

For n = 3, there are 9 orbitals. One 3s, three 3p, and five 3d. This gives a total of 2 e- for 3s, 6 e- for 3p, and 10 e- (5 x 2) for 3d. Hence, a maximum of 2 + 6 + 10 = 18 e- are possible for the 3rd principal energy level.

For n = 4, there are 16 orbitals. One 4s, three 4p, five 4d, and seven 4f orbitals. This gives a total of 2 + 6 + 10 + 14 = 32 e- max possible.

For n = 5, we only are concerned with the 16 orbitals as for n = 4.

ELECTRON CONFIGURATIONS

An electron configuration represents a short-hand way of displaying the energy levels, orbitals, and positions of the electrons in an atom.

In writing e- configurations, the following guidelines are used.

a. **No more than 2 e- can be placed in each orbital, and they must have their spins be opposite.**

b. **Electrons fill in the lowest principal energy levels first. If an electron goes into a p sublevel with 3 orbitals, the electrons will go unpaired into each of the 3 first before filling up the remaining vacant space to give a total of 6 e-. All five d orbitals must contain 1 electron each before the 6^{th} electron can be added. This is to obey the PAULI EXCLUSION PRINCIPAL.**

c. Coefficients are used to indicate the energy level followed by the symbol for the orbital and a superscript with it to indicate the number of electrons total it contains. (A p orbital could have amaximum of a superscripted ^{6}, while a d orbital a max of ^{10}.)

Starting with H, the simplest atom, the e- configuration would be written $1s^1$.

He: $1s^2$ (It is atomic number 2 so it has 2 e- with it and each of these go into a 1s orbital as it can hold a max of 2.)

Li: $1s^22s^1$; Be: $1s^22s^2$; B: $1s^22s^22p^1$; C: $1s^22s^22p^2$; N: $1s^22s^22p^3$

O: $1s^22s^22p^4$; F: $1s^22s^22p^5$; Ne: $1s^22s^22p^6$

SEE PG. 208 OF TEXT AND HANDOUT FOR ELECTRON CONFIGURATIONS OF MOST ELEMENTS

NOTE*: *ELECTRON CONFIGURATIONS CAN BE ABBREVIATED USING THE NEAREST NOBLE OR INERT GAS BEFORE THAT PARTICULAR ELEMENT AND WRITING THE NOBLE GAS'S SYMBOL IN BRACKETS.

EXAMPLE: Carbon's configuration could be written $1s^2 2s^2 2p^6 3s^2 3p^4$
OR the abbreviated form would be [Ne]$3s^2 3p^4$.

Valence electrons

Valence refers to outermost and signifies those electrons that are furthest away from the nucleus of the atom. THESE ARE THE ELECTRONS RESPONSIBLE FOR CHEMICAL REACTIONS! *These are found by looking at the outermost s and p electrons.*

Example: How many valence electrons does Mg have?

The electron configuration for Mg is $1s^2 2s^2 2p^6 3s^2$. The highest energy level *n* present is 3 and since 2 e- are in the s orbital and none in the p, Mg has 2 valence e-. NOTE: All alkaline-earth metals have similar endings for their e- configuration and so have the same number of valence e- (2).

How many valence e- does F have? $1s^2 2s^2 2p^5$:
(abbreviated form = [He]$2s^2 2p^5$) *n* = 2 is the highest energy level; hence F has 2 + 5 = 7 valence electrons.

Note: All halogens have similar ending e- configurations and thus the same number (7) of valence e-.

How many valence electrons does Ti have? $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^2$

So, the highest principal energy level is n = 4, so Ti has 2 valence e- from an s orbital.

EVEN THOUGH IT APPEARS d ELECTRONS ARE FURTHEST AWAY, THEY ARE NOT VALENCE ELECTRONS AS THEY ARE NOT IN s OR p orbitals. ***The same applies to f electrons.***

Families of Elements

The Periodic Table has rows (periods) and groups (columns). A group of elements can have a special name as they share similar physical and chemical properties. Periods can also be used to classify elements.

Group (Column) 1: Alkali metals;

Group 2: Alkaline-earth metals;

Groups 3 – 12: Transition Metals;

Group 13: Boron family;

Group 14: **Carbon family;**

Group 15: **Nitrogen family;**

Group 16: **Oxygen family or Chalcogens;**

Group 17: **Halogens;**

Group 18: **Noble or Inert Gases**

Periods 4f and 5f are termed the *innertransition metals* or *Lanthanides and Actinides*, respectively.

The Periodic Table we have today was developed by Henry Mosely who organized elements by increasing atomic number (number of protons) as opposed to increasing atomic mass (by Mendeleev and Meyer).