(COMMON IONS AND THEIR CHARGES QUIZZES BEGIN)

Chapter 6

X. Chemical Periodicity

A. Periodic Table

1. Development

a. Dimitri Mendeleev developed first periodic table that related chemical and physical properties of the elements and their atomic weight(1869). He was writing his own text book and wanted to include the new atomic weights of the elements. He began with note cards that had the elements' weight, physical and chemical properties. As he arranged them in order of increasing atomic weight he noticed that their properties started to repeat at definite intervals. He also left blanks for elements that had not yet been discovered. His table was used to predict the properties of elements not yet discovered. Mendeleev used inductive reasoning to build his periodic table (inductive reasoning - predicting a general theory from known details). He then used deductive reasoning to predict the properties of the properties of the elements yet to be discovered (deductive reasoning - predicting details from a known general theory). This didn't always work, some elements were out of order when placed strictly by their weights. Dimitri knew something was wrong and arranged his table according to properties.

b. 44 years later, in 1913, Henry Moseley was working with Rutherford and determined the nuclear charge of many elements. He did this using x-rays, the same way Rutherford had used alpha particles his gold foil experiment, shooting them through various elements noticing the wavelength change as the x-rays came through the element. This allowed Moseley to arrange the elements in order of increasing atomic number and the modern periodic law was complete - *when the elements are arranged in order of increasing atomic number, their chemical and physical properties repeat periodically*.

2. Table design

a. Groups / families - vertical columns on the periodic table in which the elements have similar chemical properties and the same number of valence electrons.

b. Periods / series - horizontal rows on the periodic table in which each successive element increases by one proton and have electrons in the same outer shell.

c. Groups and electron configurations

(a) Group 1	Alkali metals	s^1
(b) Group 2	Alkaline earth metals	s^2
(c) Group 3-12	Transition elements	s^2 (d^1 - d^{10})
(d) Group 13	Aluminum group	s^2 , p^1
(e) Group 14	Carbon group	s^2 , p^2
(f) Group 15	Nitrogen group	s^2 , p^3
(g) Group 16	Oxygen group	s^2 , p^4
(h) Group 17	Halogens	s^2 , p^5

(i) Group 18 Noble gases (octets) s^2 , p^6
(j) Lanthanide rare earth elements s^2 ($4f^1$-$4f^{14}$)
(k) Actinide series s^2 ($5f^1$-$5f^{14}$)

B. Periodic Trends

1. Atomic Radii
   a. Groups - as atomic number increases the atomic radii increases. Due to additional energy level and shielding effect - inner electrons shield outer electrons from the pull of the nucleus.
   b. Periods - as atomic number increases the atomic radii decreases. Due to increasing positive negative attraction and electrons are filling in same energy level.
2. Ionization Energy - the energy needed to remove an electron from an atom
   a. ion - an atom or group of atoms that have a positive or negative charge.
   b. A + energy ➔ $A^+ + e^-$ endothermic always
      Na + 496 kJ / mol ➔ $Na^{+1} + e^-$
   c. Groups - as atomic number increases the ionization decreases. Due to electron being held less tightly because of increased distance and shielding effect.
   d. Periods - as atomic number increases the ionization energy increases. Due to increased positive - negative attraction and getting closer to an octet
   e. Removal of successive electrons always requires more energy. Wherever the largest increase in successive ionization energies is the most stable configuration. (Table 5-3 pg. 145)
   f. Noble gases have highest I.E. Thus the octet is most stable; next is full s orbital; then half filled orbital. Stability also determines the most common ion. If largest inc. is between 1st and 2nd then +1 is most common. If largest increase is between 3rd and 4th then +3 is most common.
3. Electron Affinity - the energy change when an atom gains an electron.
   a. A + energy + e^- ➔ A^- endothermic (positive)
   b. A + e^- ➔ A^- + energy exothermic (negative)
   c. Processes in nature tend toward lower energy and higher entropy (state of disorder). Thus the more negative the more stable.
   d. E. A. tells us how much an atom wants an electron. Halogens have "highest" E. A. More evidence for octet being most stable.
   e. Groups - as atomic number increases the E. A. generally decreases
   f. Periods - as atomic number increases the E. A. generally increases
4. Ionic radii
   a. Positive ions - cations - are always smaller than the corresponding atom since losing electron(s) and electrons give the atom volume and the positive - negative attraction is greater. Less stable in E considerations; more stable in configuration considerations.
   b. Negative ions - anions - are always larger since gaining an electron and electrons give atom volume and the positive - negative attraction is smaller. If E.A. is negative then more stable in E considerations and configuration considerations.

c. Group and Period trends are the same as atomic radii trends for the same reasons.

5. Electronegativity - the tendency for atoms to attract electrons to itself when chemically combined with another atom.
   a. arbitrary scale developed by Linus Pauling called Pauling electronegativity scale (won his 1st Nobel Prize for this)
   b. based numerous factors including electron affinity and ionization energies.
   c. values help us determine the type of bonds involved in chemical compounds and molecules and also the molecular polarity
   d. Group trends - very generally decrease as atomic number increases; reasons- shielding effect and addition of energy levels electrons held more loosely
   e. Period trends - very generally increase as atomic number increases; reasons- increased positive negative attraction and getting closer to the octet