Chapter 5
Outline

Main Ideas

I. History of the Periodic Table
- A. First International Congress of Chemists
  - 1. Met in Karlsruhe, Germany in 1860
  - 2. Adopted Stanislao Cannizzaro's method for measuring relative atomic masses accurately
- B. Dmitri Mendeleev (1860's)
  - 1. Arranged elements by atomic mass to group similar chemical properties of elements
    - a. Some discrepancies were noted and spaces were left for not-yet-found elements
  - 2. First published his periodic table in 1869
- C. Henry Moseley (1911)
  - 1. Atoms fit into patterns better when arranged in order of nuclear charge
    - a. Led to modern definition of atomic number
  - 2. Arranged elements in current periodic table
- D. Noble gases
  - 1. Helium discovered in 1868 as component of the sun based on emission spectrum of sunlight
    - a. Sir Ramsey showed helium also existed on Earth
  - 2. Argon discovered by Lord Rayleigh & Sir Ramsey in 1894 as component of air
  - 3. Sir Ramsey proposed addition of new group between halogens and alkali metals
    - a. Ramsey discovered Kr & Xe in 1898
    - b. Friedrich Ernst Dorn discovered Rn in 1900
- E. The Lanthanides
  - 1. Have very similar chemical & physical properties; took many scientists to differentiate them all
  - 2. Belong to period 6 and constitute part of the *f* block
- F. The Actinides
  - 1. Similar to Lanthanides
  - 2. Belong to period 7 and constitute part of the *f* block

II. Electron Configuration and the Periodic Table
- A. Periods & Blocks
  - 1. Periods = rows (horizontal)
    - a. Signify the highest energy level attained for those elements
    - b. Length = # of electrons that can be attained by filling all sublevels
      - Period 1 = 2 e^- (H, He)
      - Period 2 = 8 e^- (Li, Be, B, C, N, O, F, Ne)
      - Period 3 = 8 e^- for Na, Mg, Al, Si, P, kS, Cl, Ar but 3rd energy level holds 18 e^- later
      - Period 4 = 18 e^- with 10 e^- in the 3*d* suborbital (Sc thru Zn); 4th energy level holds 32 max.
      - Period 5 = 18 e^- with 10 e^- in the 4*d* suborbital (Y thru Cd); 5th energy level holds 32 max
      - Period 6 = 32 e^- with 10 e^- in the 5*d* suborbital (La thru Hg) & 14 e^- in 4*f* (Lanthanides)
      - Period 7 = 32 e^- with 10 e^- in the 6*d* suborbital (La thru Hg) & 14 e^- in 5*f* (Actinides)
  - 2. Blocks
    - a. Based on electron configurations of elements; corresponds to suborbitals *s, p, d,* & *f*
      - *s* block = Groups IA & IIA **(your text's Groups 1 & 2)** alkali & alkaline earth metals
      - *p* block = Groups IIIB – VIIIB **(your text's Groups 13-18)**
      - *d* block = Groups IIIA – IIB **(your text's Groups 3-12)** = the transition metals
      - *f* block = Lanthanide & Actinide series (rare earth metals)
    - b. *s* block = most reactive metals + the nonmetal Hydrogen
      - are too reactive to be found free/ uncombined in nature & so are stored in kerosene
      - alkali metals react strongly with H_2O to yield $H_{2(g)}$ & alkalis
      - He has a filled *s* suborbital but is located in Group VIIIB with the other noble gases
    - c. *d* block = transition elements; begins in Period 4 but is part of the 3rd energy level, $n = 3$

- has slightly higher energy level than 4s, so the order of filling is 1*s* 2*s* 2*p* 3*s* 3*p* 4*s* 3*d* 4*p*
- each *d* sublevel has 5 orbitals (d_{x2-y2}, d_{xy}, d_{yz}, d_{xz}, d_{z2}) with a max of 2 e^- in each
- these elements do **not** necessarily have identical outer electron configurations
- Pd, Pt, & Au are the least reactive of all metals

d. *p* block = "main-group elements", together with the *s* block elements

- filled only after corresponding *s* block is filled
- has 3 orientations: p_x, p_y, & p_z, with a max of 2 e^- in each
- has great variety within this block: metals, metalloids, nonmetals $\pm$ halogens, noble gases
  Halogens are the most reactive nonmetals!!! They form salts when reacted with metals.
  Metalloids are great semiconductors!!! They share properties of metals & nonmetals.

e. *f* block = rare earth metals = Lanthanide + Actinide series

- found between Groups 3 & 4 in Periods 6 & 7
- Lanthanides are shiny metals similar in properties to alkaline-earth metals
- Actinides are all radioactive with the 1st 4 (Th - Np) found naturally on Earth

III. Electron Configuration & Periodic Properties (FOR MAIN-GROUP ELEMENTS)

A. Atomic Radii
1. Gradual decrease as you progress *to the right* across the periodic table due to rise in nuclear charge
2. Generally increase as you progress *down* the periodic table due to rise in # of energy levels

B. Ionization energy
1. Energy required to remove 1st electron (IE_1), 2nd electron (IE_2), 3rd electron (IE_3), etc...
   a. Amount of energy required increases significantly as you strip away the electrons.
2. Gradual increase as you progress *to the right* across the periodic table (e^- are harder to remove)
3. Generally decrease as you progress *down* the periodic table (e^- easier to remove with > shielding)

C. Electron affinity
1. The energy change that occurs when an electron is acquired by a neutral atom $A + e^- \rightarrow A^- + kJ/mol$
   a. Values are typically negative, increasingly so as atoms electron affinities increase.
   b. Most atoms *release* energy when they acquire an electron, although some are forced to gain e^-.
2. Generally, electronegativity increases as you progress *right* across the periodic table.
   a. An exception occurs between Carbon & Nitrogen. See page 157 in text.
3. Generally decreases as you progress *down* the periodic table.
   a. A slight increase in effective nuclear charge down a group increases e^- affinities
   b. A bigger increase in atomic radii down a group decreases electron affinities. (except for Cl)
   c. For isolated ions in gas phase, it's always harder to add a 2nd electron to the anion.

D. Ionic radii
1. Formation of cation (removal of valence e^-) always leads to a decrease in radii => (smaller e^- cloud)
   a. Cationic radii decrease across a period because nuclear charge increases, thus pulling in e^-.
   b. Cationic radii increase down a group as the # of energy levels increase.
2. Formation of anion (addition of valence e^-) always leads to an increase in radii => (larger e^- cloud)
   a. Anionic radii decrease across a period because nuclear charge increases, thus pulling in e^-.
   b. Anionic radii increase down a group as the # of energy levels increase.

E. Electronegativity
1. Linus Pauling devised a scale of numerical values reflecting tendencies of atoms to attract electrons.
   a. Fluorine, the most electronegative, was assigned a value of 4. Others are relative to this.
2. Electronegativities tend to increase across a period.
3. Electronegativities tend to decrease down a group OR STAY THE SAME.

IV. Periodic Properties of the *d*- & *f*-Block elements pp 163-164

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A. *d*-block elements have 0-2 e^- in their *s* orbital & 1-10 e^- in their *d* orbital, all able to interact with others.
B. Atomic radii
    1. Radii tend to <u>decrease</u> slightly across period for *d*-block atoms. [See Figure 14, page 152.]
        a. As the # of e^- in *d* orbitals increase, the radii increase due to repulsion between electrons.
C. Ionization energy
    1. Generally <u>increase</u> across the period, as was seen in main-group elements.
    2. IE_1 generally also <u>increase</u> down the group as the *s* e^- are less shielded from nuclear charge by *d* e^-.
    3. Atoms lose the *s*-orbital electrons first. Most *d*-block cations therefore have a 2^+ charge.
        a. Fe & Cr also form 3^+ cations.
        b. Ag only forms 1^+ cations.
        c. Cu forms 2+ & 1^+ cations.
D. Electronegativity
    1. *d*-block elements all have values between 1.1 & 2.54, with only Group IA & IIA having lower #s.
    2. Electronegativity values tend to increase as radii decrease, and vice versa.
    3. *f*-block elements are similar with values from 1.1-1.5.