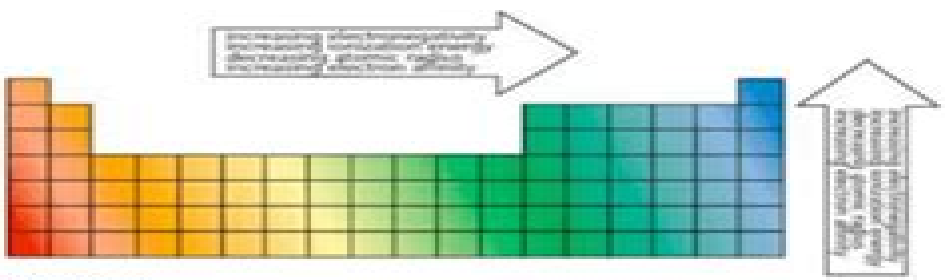
Chemistry:

CHAPTER 1

- Charge of electron ≈ -1.6×10^{-19}; Avogadro's # ≈ 6×10^{23}
- Mass number (A) on top of element, atomic number (Z) below
- $E = \frac{hc}{\lambda} = -R\left[\frac{1}{n_2^2} - \frac{1}{n_1^2}\right]$; E = qV (energy = charge on e- x electron volts); c = fλ ($c = 3 \times 10^8$)
- **Heisenberg uncertainty** – impossible to determine momentum and position simultaneously
- **Pauli exclusion** – no electrons have same quantum #'s
- Max e- in a shell ≈ $2n^2$, max e- in subshell ≈ 4l+2
  - Possible values for l ≈ 0 to n-1, possible values for ml ≈ -l to +l
- S can hold 2 e-, p ≈ 6, d ≈ 10, f ≈ 14
- **Aufbau principle** (building up) – lowest n+l will fill first
- For anions add extra e-, for cations remove e- from subshells with highest n first
- **Hund's** – half filled orbitals with parallel spins (bus seats)
- **Paramagnetic** – unpaired parallel spin e-, attracted to magnetic field
  - **Diamagnetic** – paired e-, repelled by magnetic field
  - **Ferromagnetic** – unpaired e-, randomly oriented dipoles, become magnetized in field
- A faraday is equal to one mole of electric charge

CHAPTER 2

- **Z eff** increases from left to right, constant up and down
- **Ionic radius** is smallest near metalloid line and increases out
- F is smallest, most EN, highest IE and EA; Cs is the largest and least EN
- Complex ions from transition metals associate with water to form **hydration complexes**
- **Atomic radius** has the opposite trend of all others



Periodic Trends

Left → Right
- Atomic radius ↓
- Ionization energy ↑
- Electron affinity ↑
- Electronegativity ↑

Top → Bottom
- Atomic radius ↑
- Ionization energy ↓
- Electron affinity ↓
- Electronegativity ↓

Note: Atomic radius is always opposite the other trends, ionic radius is variable.

CHAPTER 3

- C, N, O, F, Na, Mg always abide by octet rule; H, He, Li, Be, B can have incomplete; period 3 or higher can have expanded
- **Coordinate covalent bond** – both e- are contributed by one atom (lone pair attack, acid-base)
- Ionic ≈ EN difference greater than 1.7, polar covalent ≈ 0.5 to 1.7, nonpolar ≈ less than 0.5
  - Ionic are stronger than covalent (which are both stronger than IM forces)
    - Van der Waals occurs more with close proximity and large molecules
- Dipole moment ≈ p = qd
- Triple bonds are shorter and have more energy
- Major resonance contributor is more stable (less formal charge, negative on more EN atom)
  - Formal charge = valence – nonbonding – ½ bonding
- Electron geometry is when bonding and lone are equal, molecular geo is when lone repel more

Section 2nuclear Chemistry Study Guide Key

David Kirk

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