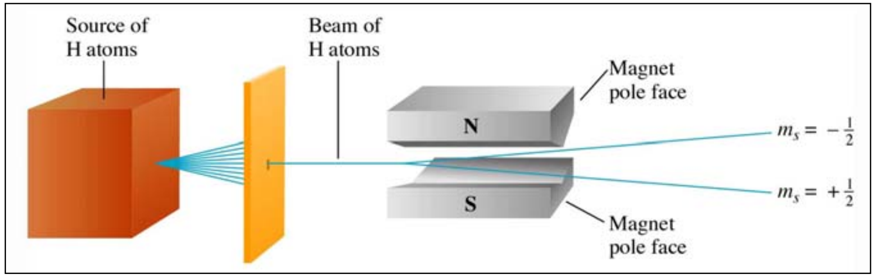
# Figure 8.2: The Stern-Gerlach experiment: showing two values of electron spin

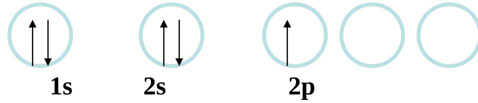



The electrons behave like tiny bar magnets. A beam of hydrogen atoms is split in two by a magnetic field due to these magnetic properties of the electrons.



# Electron Configuration

- An **orbital diagram** is used to show how the orbitals of a sub shell are occupied by electrons.
  - Each orbital is represented by a circle (or a bar, as we did in class).
    - Each group of orbitals is labeled by its sub shell notation.




  - Electrons are represented by arrows: up for $m_s = +1/2$ and down for $m_s = -1/2$

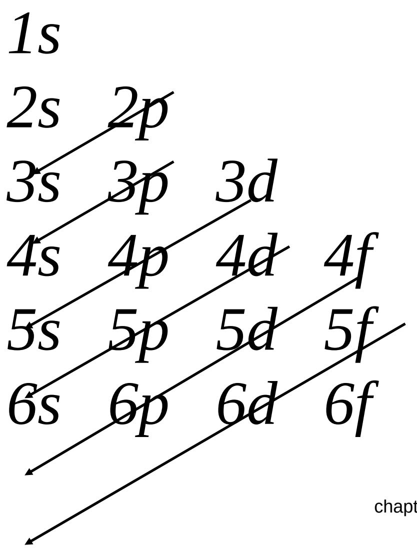
- The **Aufbau principle** is a scheme used to reproduce the ground state (lowest energy) electron configurations of atoms by following the “building up” order.
  - Listed below is the order in which all the possible sub-shells fill with electrons.

**1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f**

You will remember it by looking at Periodic Table



# Order for Filling Atomic Subshells



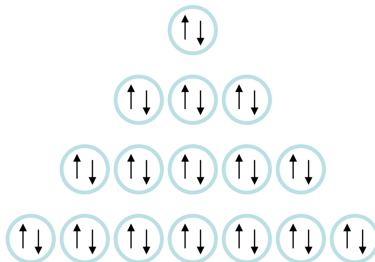



**Electron Configurations Rules review:**

The **Aufbau principle** is a scheme used to reproduce the ground state (lowest energy) electron configurations of atoms by following the “building up” order.

- **Hund’s rule** states that the lowest energy arrangement (the “ground state”) of electrons in a sub-shell is obtained by putting electrons into separate orbitals of the sub shell with the same spin before pairing electrons.
- The **Pauli exclusion principle,** which summarizes experimental observations, states that no two electrons can have the same four quantum numbers.
  - In other words, an orbital can hold at most two electrons, and then only if the electrons have opposite spins.

*(Illustrate each rule by an example!)*



| Sub shell | Number of Orbitals | Maximum Number of Electrons |
|---|---|---|
| s ($l$ = 0) | 1 | 2 |
| p ($l$ = 1) | 3 | 6 |
| d ($l$ =2) | 5 | 10 |
| f ($l$ =3) | 7 | 14 |

Few examples:

- With boron (Z=5), the electrons begin filling the 2p subshell.

| | | | | |
|---|---|---|---|---|
| **Z=5** | **Boron** | $1s^2 2s^2 2p^1$ | **or** | $[He]2s^2 2p^1$ |
| **Z=6** | **Carbon** | $1s^2 2s^2 2p^2$ | **or** | $[He]2s^2 2p^2$ |
| **Z=7** | **Nitrogen** | $1s^2 2s^2 2p^3$ | **or** | $[He]2s^2 2p^3$ |
| **Z=8** | **Oxygen** | $1s^2 2s^2 2p^4$ | **or** | $[He]2s^2 2p^4$ |
| **Z=9** | **Fluorine** | $1s^2 2s^2 2p^5$ | **or** | $[He]2s^2 2p^5$ |
| **Z=10** | **Neon** | $1s^2 2s^2 2p^6$ | **or** | $[He]2s^6 2p^6$ |

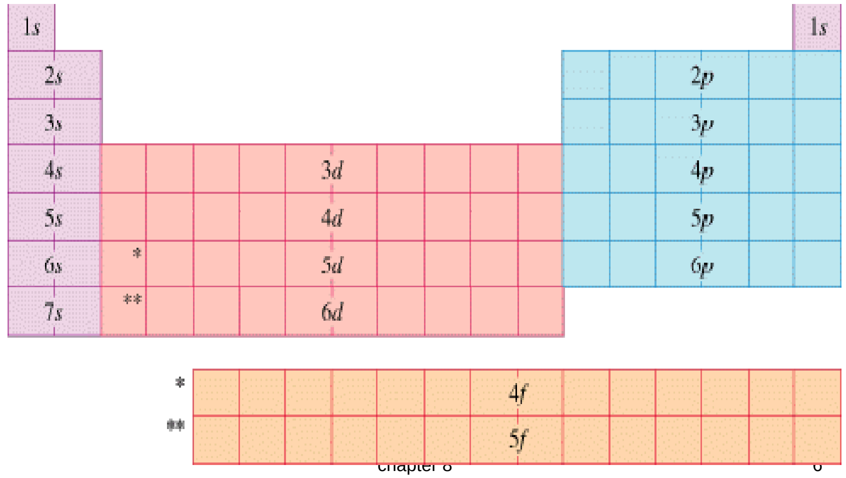
Complete octet (that is 8 electrons maximum) in Neon on the shell with n = 2 can not be filled any more – hence this is a very stable electron configuration for this and all other noble gases in group VIII

| | | | | |
|---|---|---|---|---|
| **Z=13** | **Aluminum** | $1s^2 2s^2 2p^6 3s^2 3p^1$ | **or** | $[Ne]3s^2 3p^1$ |

Electrons outside of the noble gas core are most reactive and are called VALENCE electrons

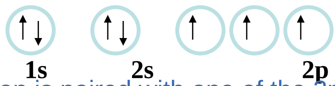


Fill in Configurations and Group names for each group in the Periodic Table: alkali metals, alkaline earth metals, transition metals, group III, IV, V, VI elements, halogens and noble gases, inner transition metals.

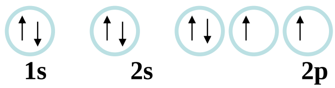





- To apply Hund's rule to oxygen, whose ground state configuration is **$1s^2 2s^2 2p^4$,** we place the first seven electrons as follows.




  – The last electron is paired with one of the 2p electrons to give a doubly occupied orbital.

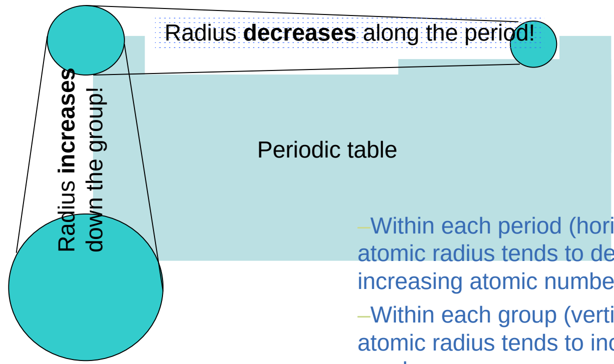



Note that oxygen atom has 2 unpaired electrons. Oxygen is in fact **paramagnetic** – that is attracted by magnetic field due to the existence of unpaired electrons. Another example of paramagnetic substance is **iron, Fe.** (Practice writing Fe electron configuration: ______________________________ )

Substances that are not affected or slightly repelled by magnetic field usually do not have available unpaired electrons and are called diamagnetic. For example all noble gases are diamagnetic.



- The **periodic law** states that ***when the elements are arranged by atomic number, their physical and chemical properties vary periodically***.
  – **Atomic radius ← most critical!**
  – **Ionization energy**
  – **Electron affinity**




– Within each period (horizontal row), the atomic radius tends to decrease with increasing atomic number (nuclear charge).
– Within each group (vertical column), the atomic radius tends to increase with the period number.

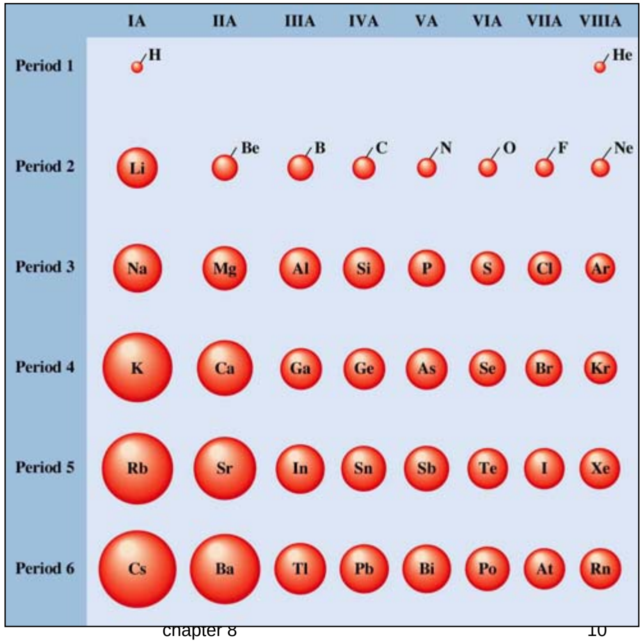
# Periodic Properties

- Two factors determine the size of an atom.
  - One factor is the principal quantum number, **n**. The larger is “n”, the larger the size of the orbital.
  - The other factor is the **effective nuclear charge,** which is the positive charge an electron experiences from the nucleus minus any “shielding effects” from intervening core electrons.

***Coulombs Law:*** **Attraction Force = Q1 Q2/$r^2$**
**Force of attraction** between two charges **Q1** and **Q2** separated by distance **r:**
**is proportional to the charges value and inverse to the square of distance between them.**
*Stronger force will hold two charges closer together and result in smaller radius **r**.*



Figure 8.17: Representation of atomic radii (covalent radii) of the main-group elements.






## Periodic Properties: Ionization energy

- The **first ionization energy** of an atom is the minimal energy needed to remove the highest energy (outermost) electron from the neutral atom.
  - For a lithium atom, the first ionization energy:

$$Li(1s^2 2s^1) \rightarrow Li^+(1s^2) + e^-$$

**Ionization energy = 520 kJ/mol**

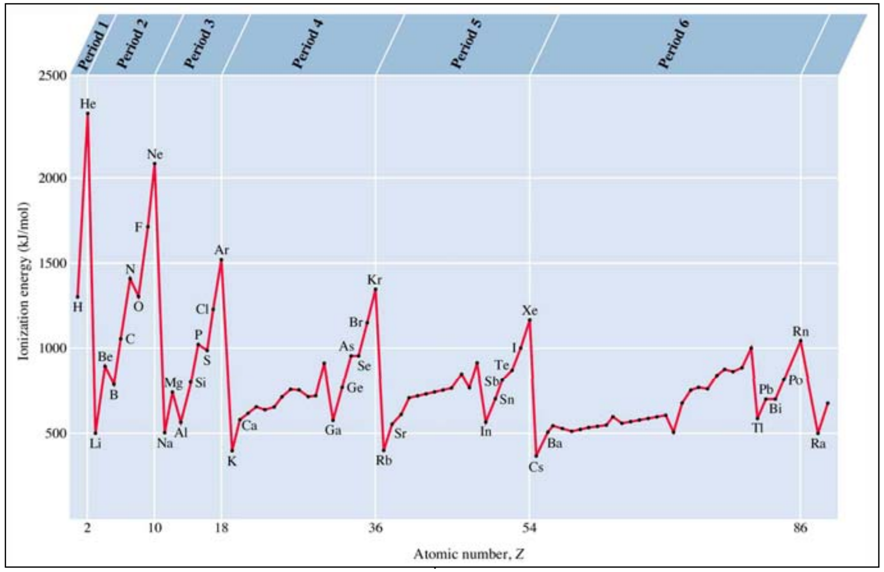
•There is a general trend that ionization energies increase with atomic number within a given period.

–This follows the trend in size, **as it is more difficult to remove an electron that is closer to the nucleus.**

–For the same reason, we find that ionization energies, again following the trend in size, decrease as we descend a column of elements.



Figure 8.18: Ionization energy versus atomic number.



**Ionization energy:** The electrons of an atom can be removed successively, one after another. The energies required at each step are known as the ***first*** ionization energy, the ***second*** ionization energy, etc.

**Table 8.3**

**Successive Ionization Energies of the First Ten Elements (kJ/mol)***

| Element | First | Second | Third | Fourth | Fifth | Sixth | Seventh |
|---|---|---|---|---|---|---|---|
| H | 1312 | | | | | | |
| He | 2372 | 5250 | | | | | |
| Li | 520 | 7298 | 11,815 | | | | |
| Be | 899 | 1757 | 14,848 | 21,006 | | | |
| B | 801 | 2427 | 3660 | 25,025 | 32,826 | | |
| C | 1086 | 2353 | 4620 | 6222 | 37,829 | 47,276 | |
| N | 1402 | 2857 | 4578 | 7475 | 9445 | 53,265 | 64,358 |
| O | 1314 | 3388 | 5300 | 7469 | 10,989 | 13,326 | 71,333 |
| F | 1681 | 3374 | 6020 | 8407 | 11,022 | 15,164 | 17,867 |
| Ne | 2081 | 3952 | 6122 | 9370 | 12,177 | 15,238 | 19,998 |

*Ionization energies to the right of a vertical line correspond to removal of electrons from the core of the atom.

# Periodic Properties: **Electron Affinity**

- The **electron affinity** is ***the energy change for the process of adding an electron to a neutral atom in the gaseous state to form a negative ion***.
  - For a chlorine atom, the first electron affinity is:

$$Cl([Ne]3s^2 3p^5) + e^- \rightarrow Cl^-([Ne]3s^2 3p^6)$$

**Electron Affinity = -349 kJ/mol**

–The more negative the electron affinity, the more stable the negative ion that is formed.

–Broadly speaking, the general trend goes from lower left to upper right as **electron affinities become more negative** (i.e. it is more favorable to add electrons in upper right corner of periodic table – to Chlorine or fluorine atoms especially).



**Table 8.4**

**Electron Affinities of the Main-Group Elements (kJ/mol)***

| Period | IA | IIIA | IVA | VA | VIA | VIIA |
|---|---|---|---|---|---|---|
| 1 | **H** −73 | | | | | |
| 2 | **Li** −60 | **B** −27 | **C** −122 | **N** 0 | **O** −141 | **F** −328 |
| 3 | **Na** −53 | **Al** −44 | **Si** −134 | **P** −72 | **S** −200 | **Cl** −349 |
| 4 | **K** −48 | **Ga** −30 | **Ge** −120 | **As** −77 | **Se** −195 | **Br** −325 |
| 5 | **Rb** −47 | **In** −30 | **Sn** −121 | **Sb** −101 | **Te** −190 | **I** −295 |
| 6 | **Cs** −45 | **Tl** −30 | **Pb** −110 | **Bi** −110 | **Po** −180 | **At** −270 |

*Atoms of the alkaline earth metals (Group IIA) and the noble gases (Group VIIIA) do not form stable negative ions.



# Chemical Periodicity in Mendeleev's Periodic Table (1872)

| Reihen | Gruppe I. — $R^2O$ | Gruppe II. — RO | Gruppe III. — $R^2O^3$ | Gruppe IV. $RH^4$ $RO^2$ | Gruppe V. $RH^3$ $R^2O^5$ | Gruppe VI. $RH^2$ $RO^3$ | Gruppe VII. RH $R^2O^7$ | Gruppe VIII. — $RO^4$ |
|---|---|---|---|---|---|---|---|---|
| 1 | H = 1 | | | | | | | |
| 2 | Li = 7 | Be = 9,4 | B = 11 | C = 12 | N = 14 | O = 16 | F = 19 | |
| 3 | Na = 23 | Mg = 24 | Al = 27,3 | Si = 28 | P = 31 | S = 32 | Cl = 35,5 | |
| 4 | K = 39 | Ca = 40 | – = 44 | Ti = 48 | V = 51 | Cr = 52 | Mn = 55 | Fe = 56, Co = 59, Ni = 59, Cu = 63. |
| 5 | (Cu = 63) | Zn = 65 | – = 68 | – = 72 | As = 75 | Se = 78 | Br = 80 | |
| 6 | Rb = 85 | Sr = 87 | ?Yt = 88 | Zr = 90 | Nb = 94 | Mo = 96 | – = 100 | Ru = 104, Rh = 104, Pd = 106, Ag = 108. |
| 7 | (Ag = 108) | Cd = 112 | In = 113 | Su = 118 | Sb = 122 | Te = 125 | J = 127 | |
| 8 | Cs = 133 | Ba = 137 | ?Di = 138 | ?Ce = 140 | — | — | — | — — — — |
| 9 | (—) | — | — | | — | — | — | |
| 10 | — | — | ?Er = 178 | ?La = 180 | Ta = 182 | W = 184 | — | Os = 195, Ir = 197, Pt = 198, Au = 199. |
| 11 | (Au = 199) | Hg = 200 | Tl = 204 | Pb = 207 | Bi = 208 | — | — | |
| 12 | — | — | — | Th = 231 | — | U = 240 | — | — — — — |

Predicted gallium (Ga, #31) properties!

## The Main-Group Elements:

- The physical and chemical properties of the main-group elements clearly display periodic behavior.
  - Variations of metallic-nonmetallic character.
  - Basic-acidic behavior of the oxides.

**Group IA, Alkali Metals;**
- Largest atomic radii
- React violently with water to form H2
- Readily ionized to 1+
- Metallic character, oxidized in air
- $R_2O$ in most cases

**Group IIA, Alkali Earth Metals :**
- Readily ionized to 2+
- React with water to form H2
- Closed s shell configuration
- Metallic

**Transition Metals:**
*May have several oxidation states *Metallic *Reactive with acids



**Group III A:**
Metals (except for boron)
Several oxidation states (commonly 3+)

**Group IV A:**
Form the most covalent compounds
Oxidation numbers vary between 4+ and 4-

**Group V A:**
Form anions generally(1-, 2-, 3-), though positive oxidation states are possible
Form metals, metalloids, and nonmetals

**Group VI A:**
Form 2- anions generally, though positive oxidation states are possible
React vigorously with alkali and alkali earth metals
Nonmetals

**Halogens (group VII):**
Form monoanions
High electronegativity (electron affinity)
Diatomic gases
Most reactive nonmetals (F)

**Noble Gases (group VIII):**
Minimal reactivity
Monatomic gases
Closed shell



# Operational Skills

- Applying the Pauli exclusion principle.
- Determining the configuration of an atom using the Aufbau principle.
- Determining the configuration of an atom using the period and group numbers.
- Applying Hund’s rule.
- Applying periodic trends.