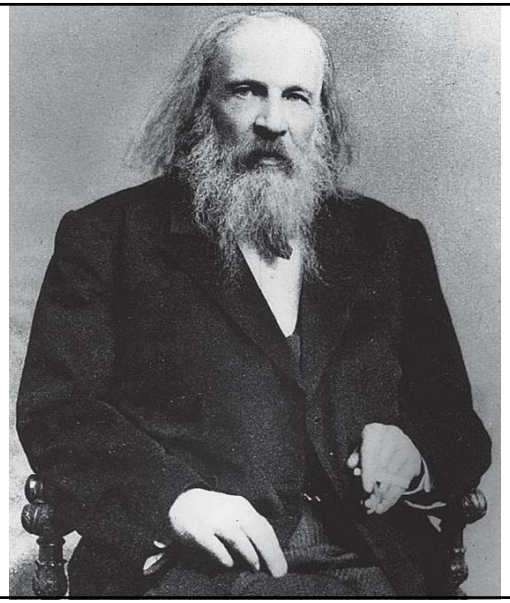


Electron Configuration and Chemical Periodicity

The Periodic Table

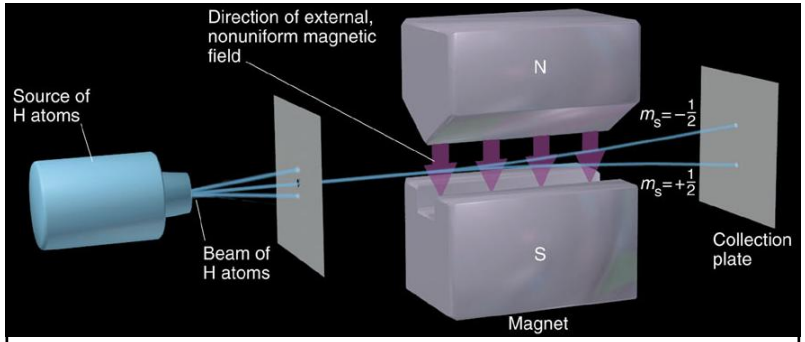
- **Periodic law** (**Mendeleev**, **Meyer**, 1870) – periodic reoccurrence of similar physical and chemical properties of the elements arranged by increasing **atomic mass**
  - Periodic table included the 65 known elements
  - Mendeleev left blank spaces for the undiscovered elements and was able to predict their properties
  - The true basis of periodicity is the **atomic number** not the atomic mass (Mosley, 1913)

8.1 Many-Electron Atoms

- Only approximate solutions of the Schrödinger equation are available
- Electron-electron interactions are important
- The same tree quantum numbers (***n, l*** and m_l) are used to describe the solutions (the orbitals are hydrogen-like)

The Electron Spin

- The electron can be viewed as a ball of spinning charge – has a magnetic moment
- The magnetic moment is quantized – only two orientations of the spin are allowed in a magnetic field



⇒**Spin quantum number (m_s) –** two possible values of m_s **(+1/2 and -1/2)**

- Each electron in an atom is described by four quantum numbers – ***n***, ***l***, m_l, m_s

- **The Pauli exclusion principle** – no two electrons in an atom can have the same set of four quantum numbers

  ⇒**Each orbital can hold no more than two electrons and they must have opposite spins** (paired spins, ↑↓)

Orbital Energies

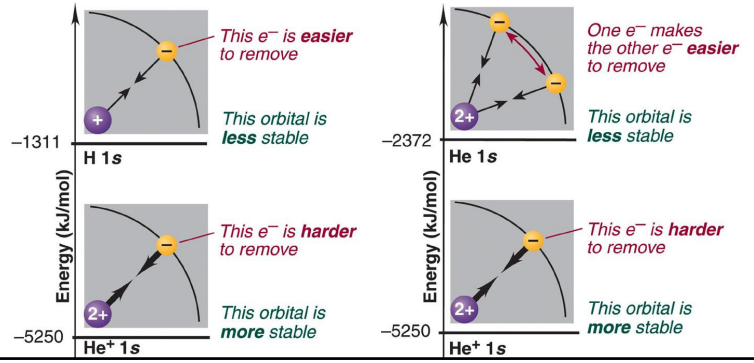
- Orbital energies depend on both ***n*** and ***l***

  $n\uparrow \rightarrow E\uparrow$ $\qquad$ $l\uparrow \rightarrow E\uparrow$

  ⇒ **Orbitals in different subshells of a given principal shell have different energies**

  – Evidence – many-electron atoms have more complex atomic spectra (splitting of *E*-levels)

- Electrons are attracted by the nucleus and repelled by each other
  - **The effect of nuclear charge** (***Z***) – higher ***Z*** lowers the orbital energy
  - **The effect of electron repulsion** – an additional **e**$^-$ in the same orbital raises the orbital energy



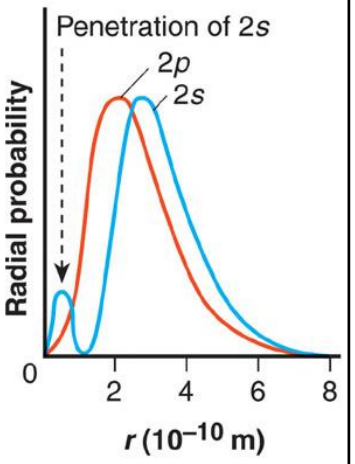
- **Electron shielding** – electrons shield each other from the nuclear charge
  - **Inner electrons** shield **outer electrons** more effectively than electrons in the same orbital or subshell
- **Effective nuclear charge** (Z_{eff}) – smaller than the actual nuclear charge (***Z***) due to electron shielding
- **Penetration** – electrons on orbitals in different subshells of a given shell are shielded to a different extent depending on their penetration (closeness) to the nucleus

  More penetration → less shielding → higher Z_{eff}

- The ***s*-orbitals** have the greatest penetration to the nucleus

  Least shielded → highest Z_{eff} → lowest energy
- The ***p*-orbitals** have a node at the nucleus and lower penetration than the ***s***-orbitals

  More effectively shielded → lower Z_{eff} → higher energy



- Penetration decreases with increasing ***l***
  - More effective shielding → lower Z_{eff} → higher energy
  - Energy order of the subshells in a given shell:

    s < p < d < f < g ...
  - This leads two much greater number of energy levels compared to one-electron species like H