7.12 Electron Configurations of Atoms

- Electron configurations can be deduced from the positions of elements in the periodic table
  - valence shell principal quantum numbers equal period numbers (**F → 2nd period, *n*=2**)
  - all elements in a period have the same **noble-gas core** configurations (**[He], [Ne], [Ar], …**)
  - all elements in a group have analogous valence shell electron configurations (**F → [He]$2s^2 2p^5$; Cl → [Ne]$3s^2 3p^5$; all halogens → ns^2np^5**)
- The filling order of the orbitals can be obtained from the periodic table

The *n*s, *n*p, (*n-1*)d and (*n-2*)f orbitals are filled in the *n*th period

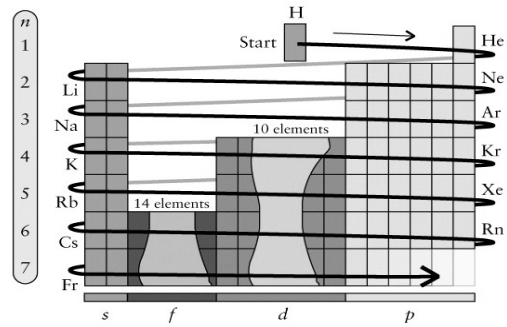


Fig. 7.31

- The ***n*s** orbitals are filled before the **(*n-2*)f** and **(*n-1*)d** orbitals, but are written after them

Sc → noble gas core [Ar] → (K, Ca) $4s^2$ → (Sc) $3d^1$

⇒[Ar]$4s^2 3d^1$ → [Ar]$3d^1 4s^2$

Example: Write the electron configuration of osmium, **Os**.

Os is in the 6th period → valence shell *n*=6

Previous noble gas is Xe → noble-gas core is [Xe]

After Xe → 2 *n*s, 14 (*n-2*)f, and 6 (*n-1*)d elements

⇒ [Xe]$4f^{14}5d^6 6s^2$

- Exceptions to the building-up principle
  - half-filled **d** and **f** subshells have exceptional stability

    Cr → [Ar]$3d^5 4s^1$ instead of **[Ar]$3d^4 4s^2$**
  - completely filled **d** and **f** subshells have exceptional stability

    Cu → [Ar]$3d^{10} 4s^1$ instead of **[Ar]$3d^9 4s^2$**

Example: Write the electron configuration of gold, **Au**.

valence shell *n*=6 noble-gas core [Xe]

After Xe → 2 6s, 14 4f, and 9 5d elements

⇒ [Xe]$4f^{14}5d^9 6s^2$ → [Xe]$4f^{14}5d^{10} 6s^1$

7.13 Electron Configurations of Ions

- Cations
  - for **p-** and **s-elements** electrons are lost first from the ***n*p** subshell followed by the ***n*s** subshell
  - for **d-elements** electrons are lost first from the ***n*s** subshell followed by the **(*n-1*)d** subshell

Example: Write the electron configurations of **Pb^{2+}** and **Pb^{4+}**.

Pb → [Xe]$4f^{14}5d^{10}6s^2 6p^2$

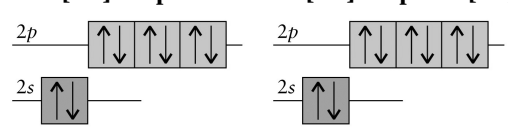
Pb^{2+} → [Xe]$4f^{14}5d^{10}6s^2$ Pb^{4+} → [Xe]$4f^{14}5d^{10}$

- Anions
  - electrons are added until a noble-gas configuration is reached

Example: Write the electron configurations of the **nitride** and **oxide** ions.

N → [He]$2s^2 2p^3$ N^{3-} → [He]$2s^2 2p^6$ → [Ne]

O → [He]$2s^2 2p^4$ O^{2-} → [He]$2s^2 2p^6$ → [Ne]



15 N^{3-} [He]$2s^2 2p^6$ **16** O^{2-} [He]$2s^2 2p^6$

7.14 Electronic Structure and the Periodic Table

- The table is divided into **s**, **p**, **d**, and **f** blocks named by the last occupied subshell being filled
- **Valence electrons** – electrons in the highest occupied principal shell and in partially filled subshells of lower principal shells
- **Main groups** – **s** and **p** elements
  - group 1 ns^1, group 2 ns^2, group 13 ns^2np^1, ..., group 18 ns^2np^6
- **Secondary groups** – **d** elements
  - group 3 $(n-1)d^1ns^2$, …, group 12 $(n-1)d^{10}ns^2$

- The number of valence electrons equals the group # (group # - 10 for p-elements)

Example: Write the valence electron configurations of groups 7 and 15.

Group 7 → 7 valence electrons, d-elements

2 s-electrons and 5 d-electrons

⇒ $(n-1)d^5ns^2$

Group 15 → 15-10=5 valence electrons, p-elements

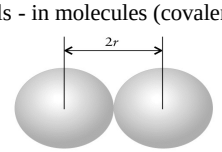
2 s-electrons and 3 p-electrons

⇒ ns^2np^3

The Periodicity of Atomic Properties

7.15 Atomic Radius

- Half of the distance between the centers of two neighboring atoms
  - metals - in the solid phase
  - nonmetals - in molecules (covalent radius)



17 Atomic radius

- Atomic radii **increase down a group** and **decrease** from left to right **across a period**

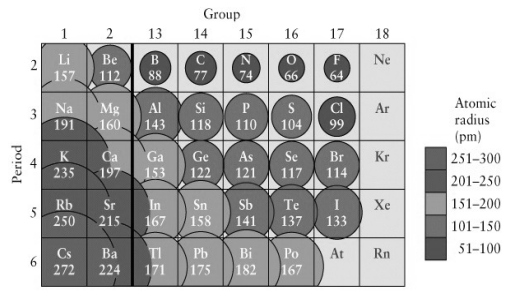


Fig. 7.34

- **Down a group** – the principal quantum number increases ⇒ orbitals and electron clouds become larger
- **Across a period** – the nuclear charge increases and the new electrons enter the same principal shell (do not shield each other effectively) ⇒ the effective nuclear charge increases ⇒ draws the electrons closer to the nucleus

Example: Compare the sizes of **Ge**, **Sn** and **Se**.

Sn is below Ge ⇒ Sn>Ge

Ge is to the left of Se ⇒ Ge>Se