Lesson 2
Electron Configurations

Learning Goals

- ❑ I will be able to draw energy level diagrams and write electron configurations for both atoms and ions.

Rules for Energy-Level Diagrams

Aufbau Principle
Electrons are added to the **lowest** energy orbital available.

Pauli Exclusion Principle
No two electrons in an atom can have the same four quantum numbers.
(i.e., only two electrons can occupy each orbital, one with +½ spin and one with - ½ spin).

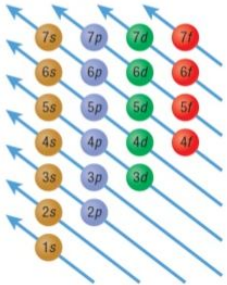
Hund's Rule
Electrons in the same sublevel will **not** pair up (occupy the same orbital) until **all** orbitals in the sublevel are **half**-filled (have 1 electron).

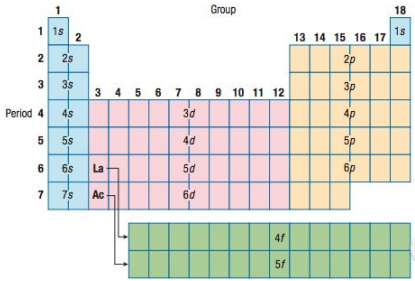
Negative Ions
The extra electrons in a negative ion occupy orbitals following the three rules.

Positive Ions
Electrons are first removed from orbitals with the highest principal quantum number (outer shell) even if these orbitals are not at the highest energy.

Examples:

1. Titanium atom





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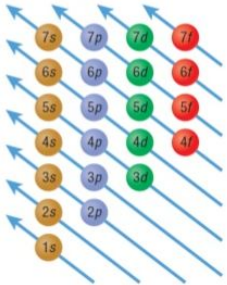
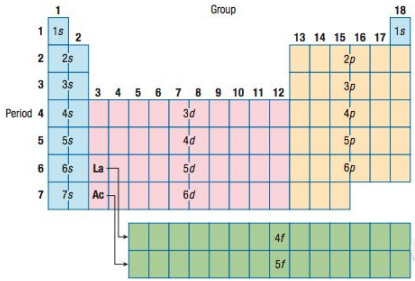
Examples:
2. Silicon atom





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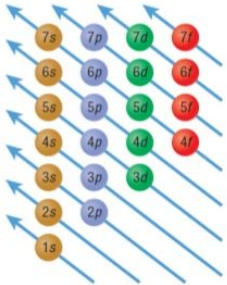
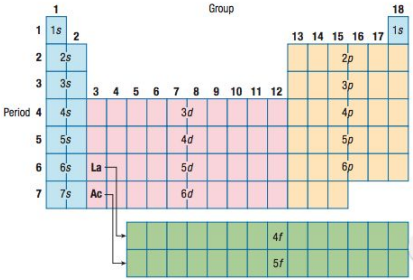
Examples:
3. Sulfide ion





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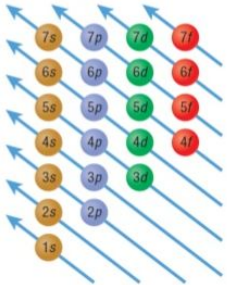
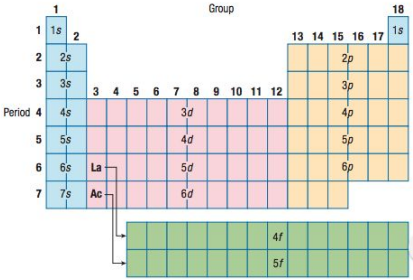
Examples:
4. Manganese (II) ion





Some Tendencies

1. Electrons are most easily lost from the outer shell.
2. Electron configurations tend to be more stable when there is…
   - an **octet** of electrons in the outer shell (ns^2np^6)
   - a **filled *d*-subshell**
   - a **half-filled *d*-subshell**

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Example 1: Explain why carbon is paramagnetic but calcium is not paramagnetic.

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Example 2: Explain the 2+ and 3+ ionic charges for iron.

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Example 3: Explain the anomaly in the electron configuration of chromium.

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Example 4: Explain why nitrogen and boron both form three covalent bonds.

Success Criteria

- ❑ I can draw electron-energy level diagrams for atoms and ions.
- ❑ I can write full and shorthand electron for atoms and ions.
- ❑ I can use these configurations to explain
  - ❑ ionic charges,
  - ❑ anomalies,
  - ❑ paramagnetism, and
  - ❑ bonding.