Rules for Energy-Level Diagrams

Aufbau Principle

Each electron occupies 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

When there is more than one orbital at the same energy level,
each orbital is occupied by one electron before any one of the
orbitals is occupied by a second electron.

Negative Ions

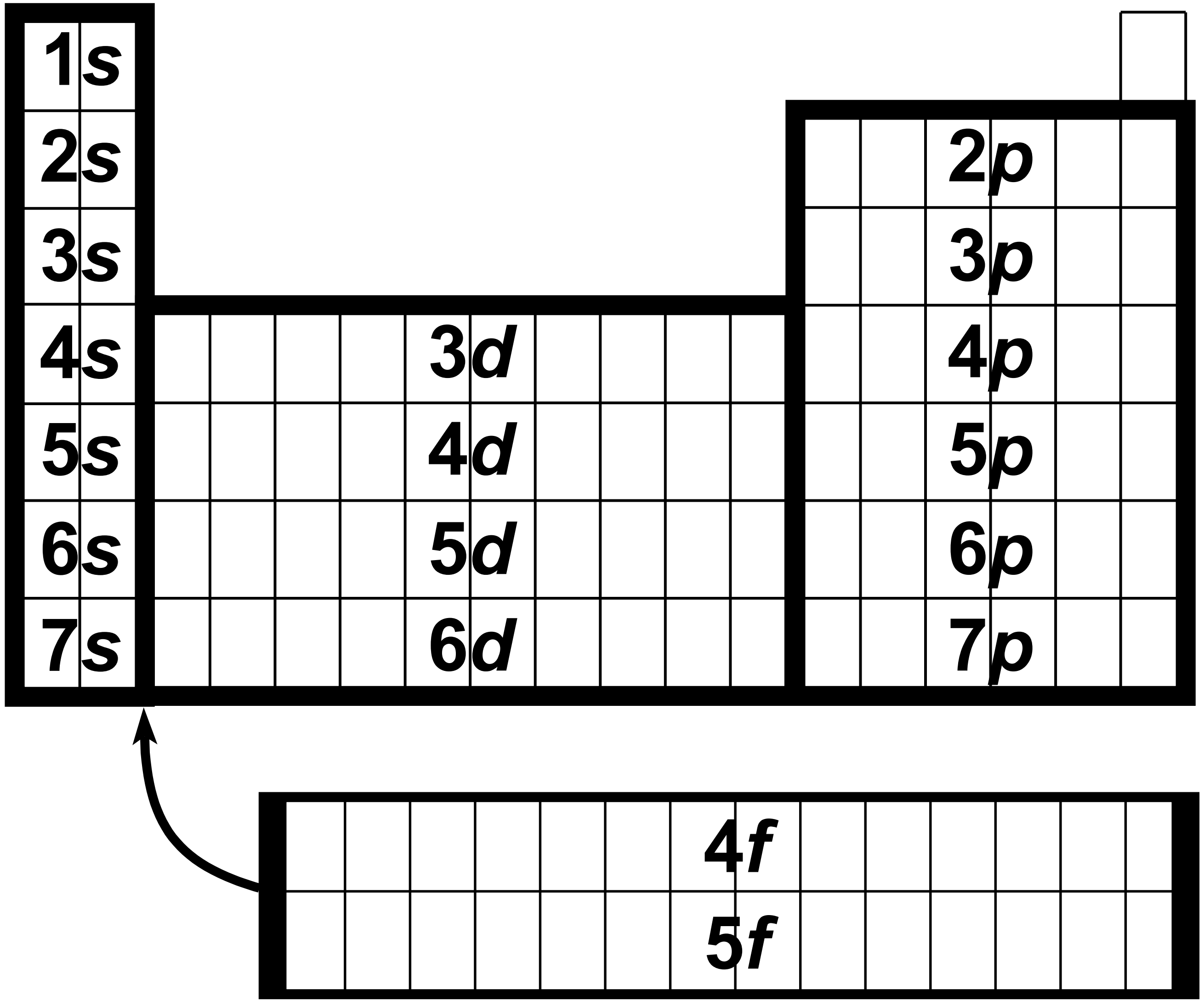
The extra electrons in a negative ion occupy orbitals following the three rules.

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

titanium atom


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

silicon atom

sulfide ion

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**

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

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

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

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