13: Electron Configuration

Key Chemistry Terms

- **Electron Cloud:** Area outside nucleus where electrons are located.
- **Energy Levels:** Electron cloud is divided into energy levels for electrons.
- **Subshells:** Energy levels of electrons are divided into subshells of equal energy orbitals.
- **Orbitals:** Subdivision of subshell. Each orbital can hold 2 electrons.
- **Valence Electrons:** On the outermost shell.
- **Isoelectric:** Atoms of different elements with the same electron configuration.

Atomic Structure

- **Protons:** Positive, in nucleus, 1 amu, determines the identity of the atom.
- **Neutrons:** Neutral, in nucleus, 1 amu, atoms of the same element with a different number of neutrons are isotopes.
- **Electrons:** Negative, outside nucleus, 0 amu; electrons can be lost or gained to form an atom with a charge (ion).

Determining the number of electrons:
Atomic number = # of protons
Charge = #protons - #electrons

Energy Levels, Subshells and Orbitals

4 types of subshells:

	Subshell	Begins in level	# of orbitals	# of electrons
↓ higher energy	s	1	1	2
	p	2	3	6
	d	3	5	10
	f	4	7	14
Subshell Mnemonic: *spdf = Smart People Don't Fail.*				

Rules for Electron Configurations:

- **Aufbau Principle:** Electrons fill subshells in an order that produces the lowest energy for the atom.
- **Hund's Rule:** When filling orbitals, electrons are placed in each equal-energy orbital before doubling up to produce the lowest energy atom.
- **Pauli Exclusion Principle:** Two electrons occupying the same orbital must be opposite spins (angular momentum).

Mnemonic for Three Electron Configuration Rules:
Aufbau (stays low); Hund (does not double up); Pauli (spin up and down) = "*Alligator stays low; Hippo does not pair up and Penguin jumps up and down.*"

Electron Configurations and the Periodic Table

Every element in a group of the periodic table has the same number of electrons in the highest energy subshell (valence).

Order of Filling for Subshells

Use the periodic table as a guide (read left to right):

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

Or use the diagonal-down method:

1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d
7s 7p
8s

To read the chart, read down one diagonal as far as possible then jump to the top of the next diagonal and repeat

Boxes and Arrow Configurations

Each orbital is shown with a box and each electron with an arrow.
- Determine the number of electrons needed & follow the 3 rules governing electron configurations.

Example:
O (8 electrons): 1s ↑↓ 2s ↑↓ 2p ↑↓ ↑ ↑

Spectroscopic Notation

Spectroscopic is a shorthand notation for electron configurations.
- The number of electrons in each subshell is written as a superscript after the subshell designation.
- The sum of the superscripts is equal to the total number of electrons.

Example:
Br (35 electrons): $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$

Noble Gas Notation

Noble gases have full electron shells.
- The noble gas closest to the number of electrons needed without going over is used to represent the core electrons.
- The spectroscopic notation is written for the valence electrons.

Steps to write a noble gas notation:
1. To write noble gas configuration, determine the number of electrons you need to place.
2. Choose the noble gas closest to that number without going over.
3. Start where that noble gas left off on the periodic table and begin filling with spectroscopic notation.

Example:
Br (35 electrons): [Ar] $4s^2\ 3d^{10}\ 4p^5$

Electron Configuration of Ions

Most ions are formed from losing or gaining electrons to result in a full valence shell.

Example:
Br^- (36 electrons): [Ar] $4s^2\ 3d^{10}\ 4p^6$

Exceptions to the Rules

A half-full "*s*" orbital and a "*d*" subshell with 5 or 10 is more stable than following the Aufbau Principle.
Cr, Mo, W: $s^1\ d^5$
Cu, Ag, Au: $s^1\ d^{10}$

Quantum Numbers

Set of 4 numbers describing the location of an electron in an atom.

Name	Symbol	Describes	Found	Possibilities
Principal energy level	n	Main energy level	Shell #2	Whole # > 0
Azimuthal number	l	Subshell shape	s = 0, p = 1, d = 2, f = 3	Whole # <n
Magnetic	m_l	Which orbital	Number line system (middle is 0)	-l → +l
Spin	m_s	Up or down arrow	Up = + ½ Down = - ½	+ or – ½

Example:
1s ↑↓ 2s ↑↓ 2p ↑↓ ↑ ↑
Quantum #'s for the red arrow: 2, 1, -1, - ½

How to Use This Cheat Sheet: These are the keys related to this topic. Try to read through it carefully twice then write it out from memory on a blank sheet of paper. Review it again before the exams.

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