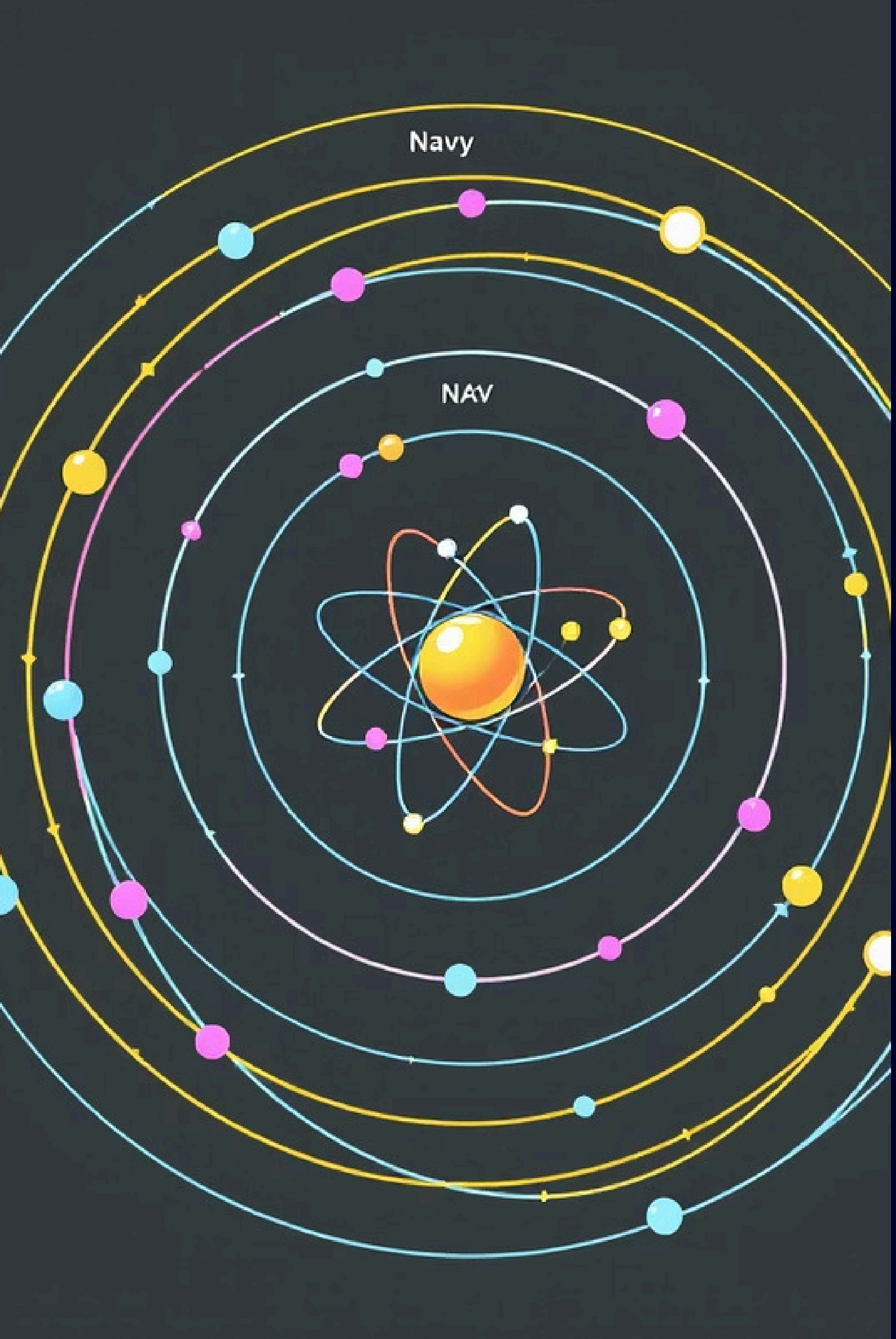


Chemistry Fundamentals

Lecture 10: Electron Configuration Basics

Mohamed Kamal

Why Electron Configuration Matters

Chemical Properties

Electron arrangement determines chemical properties and explains why elements in the same group have similar properties

Bonding & Reactivity

Electrons in outermost shell participate in bonding, allowing us to predict chemical behavior

Applications

Essential for understanding molecular structure, spectroscopy, catalysis, and periodic trends

Quantum mechanics revolutionized our understanding of atomic structure, providing the foundation for modern chemistry.

Learning Goal: Master electron configuration to predict chemical behavior

Energy Levels and Electron Shells

Principal Energy Levels

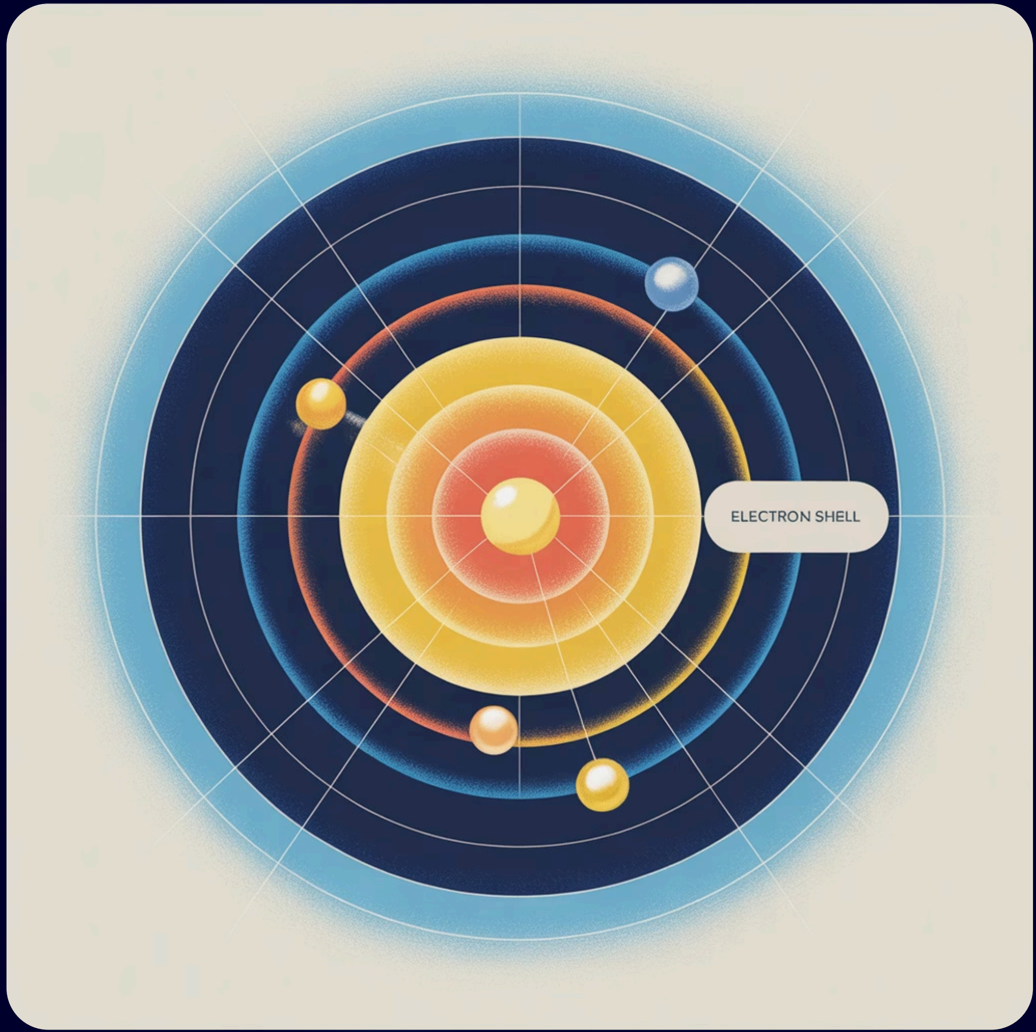
n = 1, 2, 3, 4, 5, 6, 7

Higher n = higher energy = farther from nucleus

Maximum Electrons per Level

$2n^2$ rule determines capacity:

- n = 1: 2 electrons maximum
- n = 2: 8 electrons maximum
- n = 3: 18 electrons maximum
- n = 4: 32 electrons maximum



While Bohr's planetary model helps visualize shells, quantum reality shows electrons occupy probability clouds, not fixed orbits.

Lower energy levels fill before higher ones.

Subshells and Orbital Types

1	**s subshell** 1 orbital, 2 electrons maximum Spherical shape
2	**p subshell** 3 orbitals, 6 electrons maximum Dumbbell shape
3	**d subshell** 5 orbitals, 10 electrons maximum Complex shape
4	**f subshell** 7 orbitals, 14 electrons maximum Very complex shape

Subshell names derive from spectroscopic terms: sharp, principal, diffuse, fundamental

Energy order within same principal level: s < p < d < f

Subshell notation: 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, 4d, 4f, etc.

Electron Configuration Rules

① Aufbau Principle

Electrons fill orbitals in order of increasing energy

② Pauli Exclusion Principle

No two electrons can have identical quantum numbers

Practical meaning: Maximum 2 electrons per orbital with opposite spins

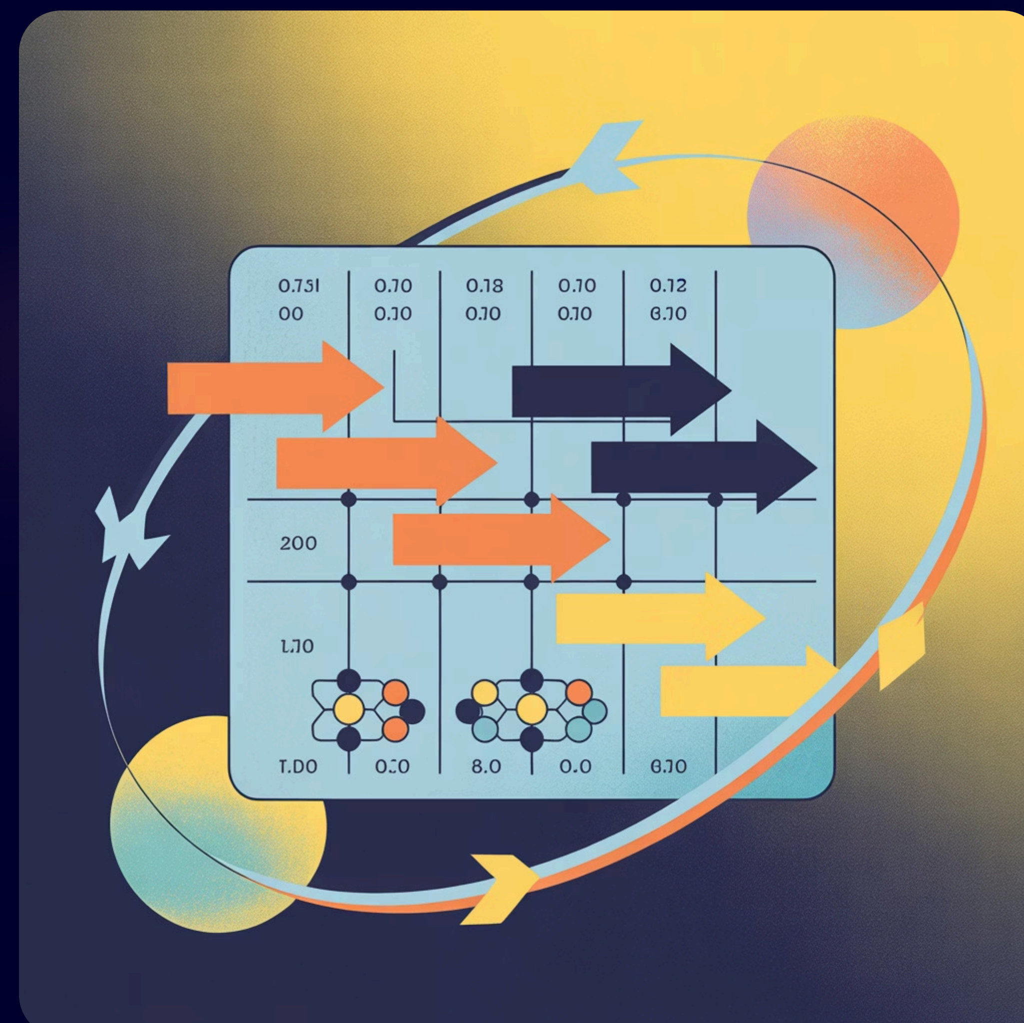
③ Hund's Rule

Electrons occupy orbitals singly before pairing up

Minimizes electron-electron repulsion

Energy Order

1s < 2s < 2p < 3s < 3p < 4s < 3d < 4p < 5s < 4d < 5p...



Memory Device: Use diagonal rule or energy level diagram

Note: Some transition metals have unexpected configurations for stability

Writing Electron Configurations

Standard Notation

List subshells in order of filling with electron count as superscripts

Step 1

Determine number of electrons (= atomic number for neutral atom)

Step 2

Follow energy order to fill subshells

Step 3

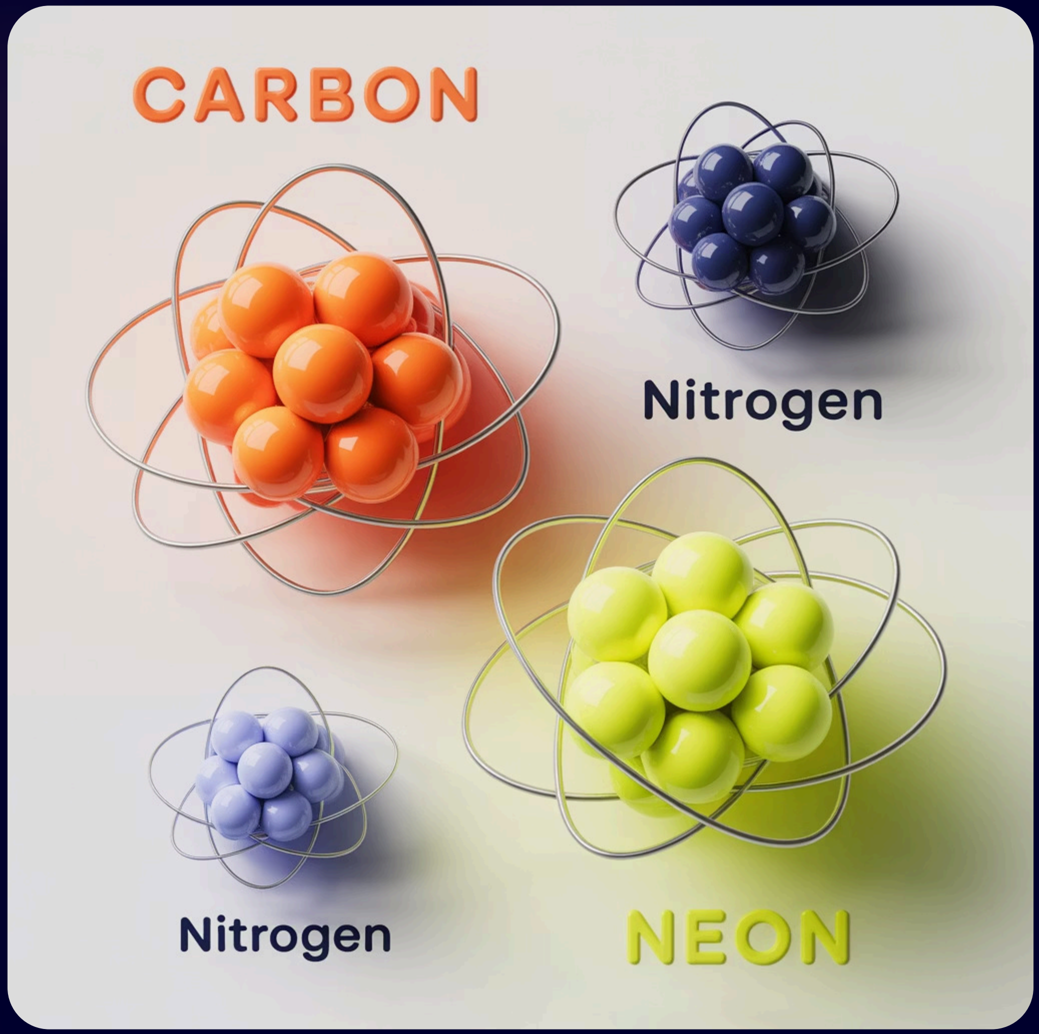
Use superscripts to show electron count

Step 4

Check total electrons = atomic number

Examples

Carbon (Z = 6)	$1s^2\ 2s^2\ 2p^2$
Nitrogen (Z = 7)	$1s^2\ 2s^2\ 2p^3$
Neon (Z = 10)	$1s^2\ 2s^2\ 2p^6$



Verification: Sum of superscripts must equal atomic number

Orbital Diagrams and Electron Spin

Orbital Diagram Basics

Visual representation showing electron placement

Electron spin shown with arrows (↑ or ↓)

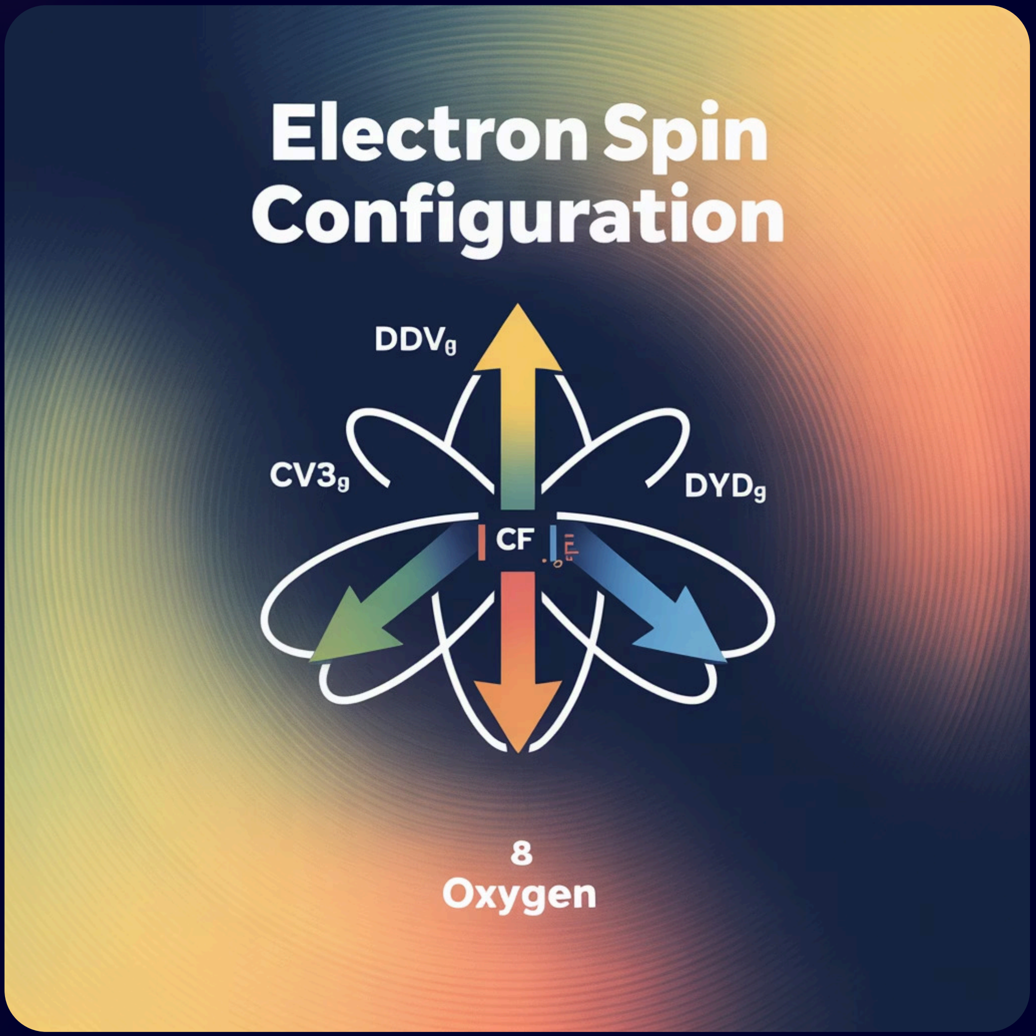
Paired electrons (↑ ↓), unpaired electrons (↑)

Unpaired electrons make atoms paramagnetic

Hund's Rule Application

Fill each orbital singly before pairing

Carbon ($1s^2\ 2s^2\ 2p^2$)	1s: ↑ ↓ 2s: ↑ ↓ 2p: ↑ ↑ _
Oxygen ($1s^2\ 2s^2\ 2p^4$)	1s: ↑ ↓ 2s: ↑ ↓ 2p: ↑ ↓ ↑ ↑



Noble Gas Configuration Shorthand

Abbreviated Notation

Use noble gas core + remaining electrons

Advantages: Shorter notation, emphasizes valence electrons

Noble Gas Cores

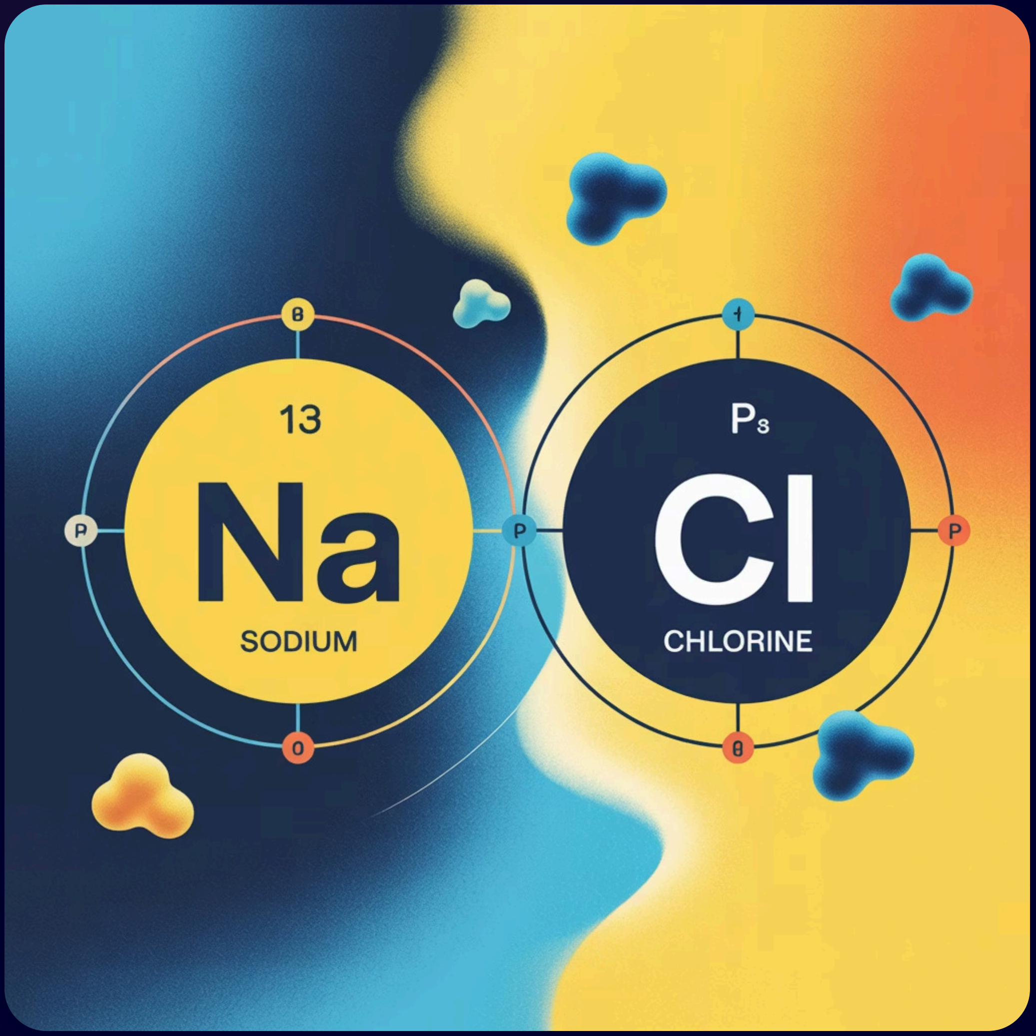
[He] = $1s^2$

[Ne] = $1s^2\ 2s^2\ 2p^6$

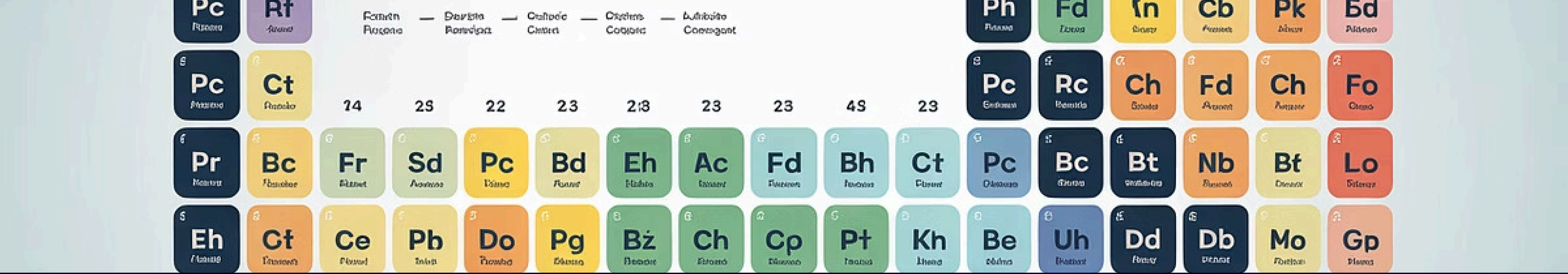
[Ar] = [Ne] $3s^2\ 3p^6$

Examples

Element	Full Configuration	Abbreviated
Sodium (Na, Z=11)	$1s^2\ 2s^2\ 2p^6\ 3s^1$	[Ne] $3s^1$
Chlorine (Cl, Z=17)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	[Ne] $3s^2\ 3p^5$



Valence electrons (those beyond noble gas core) determine chemical properties



Electron Configuration and Periodic Table

Group Patterns

- Group 1 (alkali metals): ns^1 configuration
- Group 2 (alkaline earth metals): ns^2 configuration
- Group 17 (halogens): $ns^2 np^5$ configuration
- Group 18 (noble gases): $ns^2 np^6$ configuration

Block Classification

- s-block: Groups 1-2 (filling s orbitals)
- p-block: Groups 13-18 (filling p orbitals)
- d-block: Groups 3-12 (filling d orbitals)
- f-block: Lanthanides and actinides (filling f orbitals)

Elements in the same group have similar valence electron configurations, explaining their similar chemical properties.

Electron configuration explains periodic trends like atomic size and ionization energy.

Next Lecture:

Chemical Bonding Fundamentals

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