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The amazing colors of fireworks result from electron transfer between energy levels of atoms.

Electromagnetic Radiation:

A form of energy that runs a continuum from radio to X-rays, visible light to microwaves.

Each form of radiation shares common characteristics: they display wavelike properties and travel at the same speed.

Basic Properties of Waves

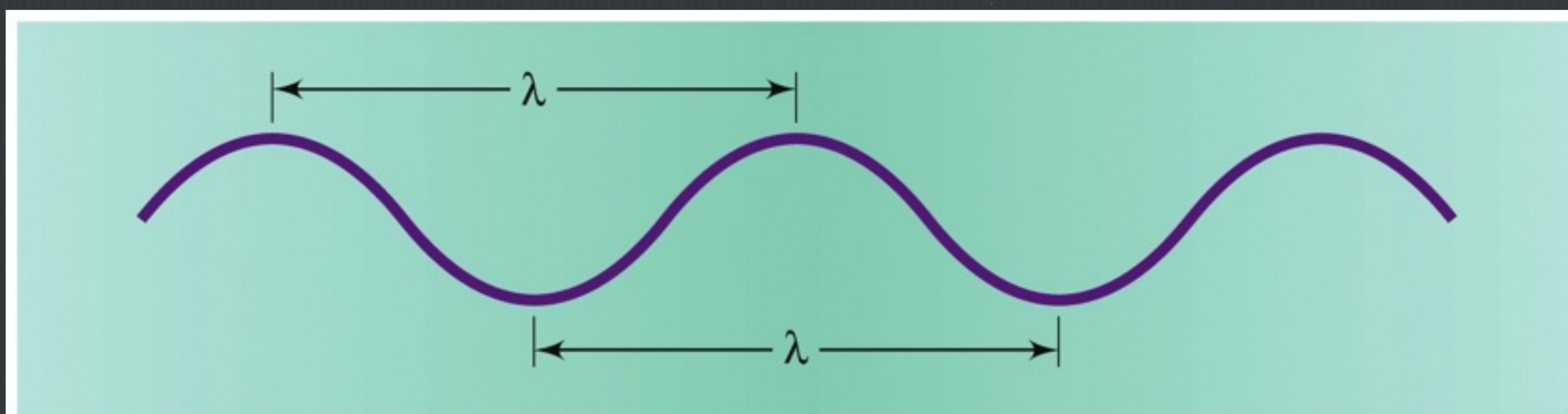
Wavelength (λ): the distance between consecutive peaks (or troughs) of a wave.

Frequency (ν): the number of times a wave passes through a certain point per second.

Speed: how fast a wave moves through space.

Basic Properties of Waves

Wavelength (λ): the distance between consecutive peaks (or troughs) of a wave.



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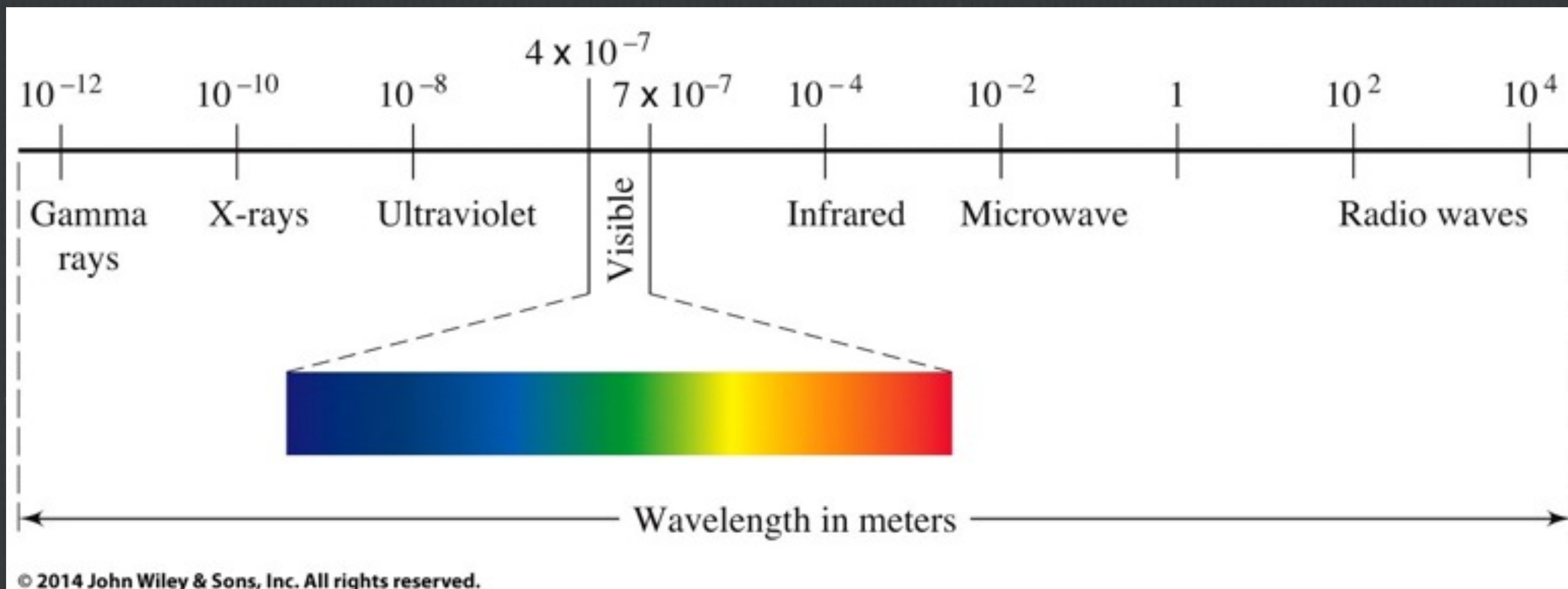
Electromagnetic Spectrum:
The full range of electromagnetic radiation, arranged based on wavelength.



Electromagnetic radiation has both wave-like and particle properties.

Radiation can behave like tiny packets ("particles") of energy called photons.

Making Light…

At high temperatures or when high voltages are applied, elements radiate (emit) colored light.

When this light is passed through a prism, a set of brightly colored lines result.

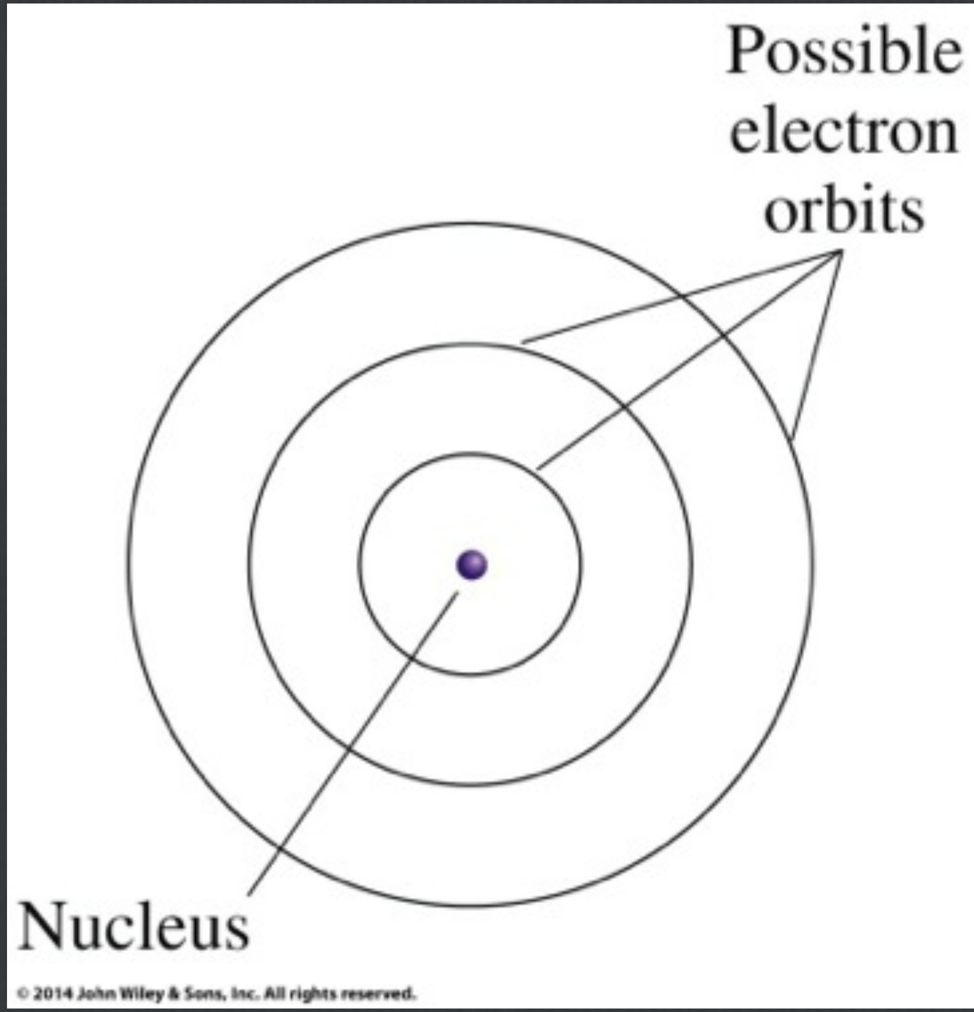
Line spectrum of hydrogen

These line spectra indicate the light emitted has only specific wavelengths/frequencies.

Each element possesses a characteristic and unique line spectrum.

Explaining the spectra

From his study of the line spectrum of hydrogen, Bohr proposed a revised theory of the atom.

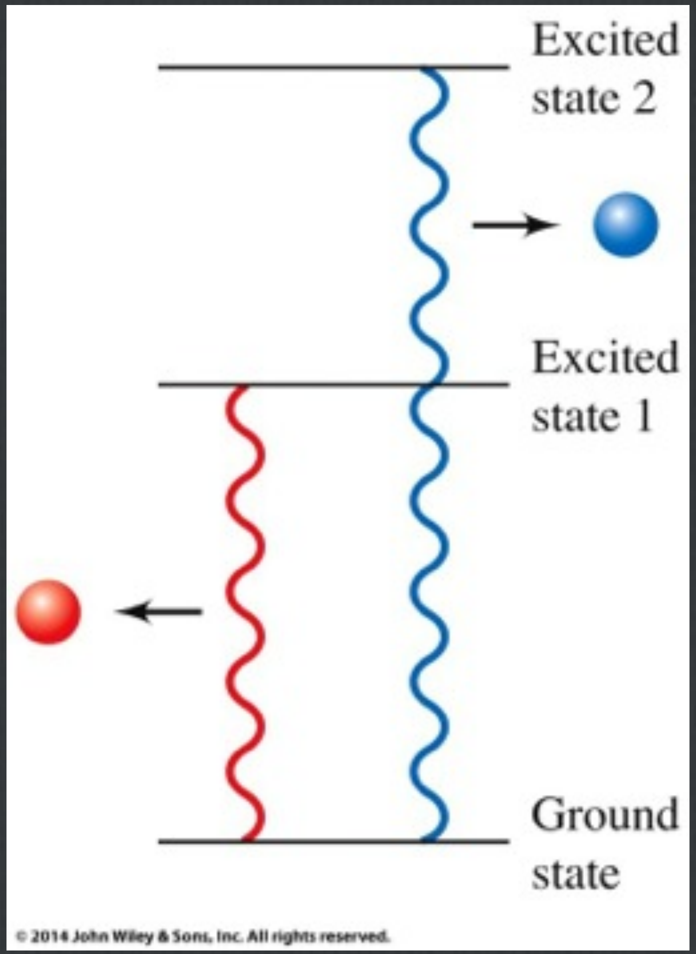


Bohr suggested electrons exist in specific regions at defined distances from the nucleus.

The electrons then move about the nucleus in circular orbits at a fixed distance from the nucleus.

The Bohr Model

Bohr also suggested energy absorbed or emitted by an atom is quantized (has discrete fixed units).



Bohr proposed that electrons can orbit the nucleus at different distances. Each orbit is a distinct, discrete (quantized) energy level.

When an atom absorbs energy, the electrons can be promoted to higher energy levels.

When an atom emits energy, the electrons can decays to a lower energy level.

Ground state: lowest energy level for an atom.

Each line in the spectrum corresponds to emission of energy as an electron relaxes from a higher to lower energy level.

Color of light emitted depends on the gap between the energy levels.

Bohr’s theory worked very well to explain and predict the line spectrum of hydrogen…. BUT….

In 1924, DeBroglie (a French chemist) proposed matter could be treated as either a wave or a particle.

For large objects, the wave properties are negligible due to their very small wavelengths.

For smaller objects with less mass, i.e. an electron, the wavelike properties of matter become very important.

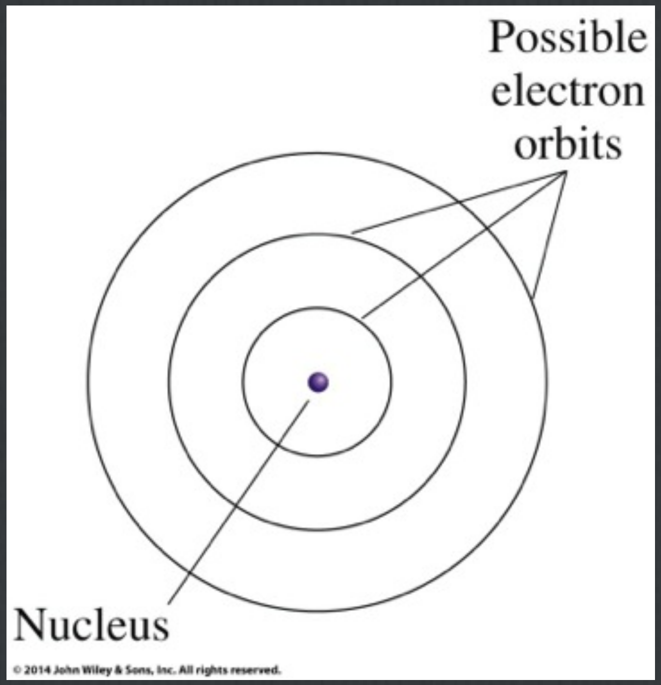
Schrödinger in 1926 expanded the wavelike properties of matter by developing an equation to describe electrons as waves.

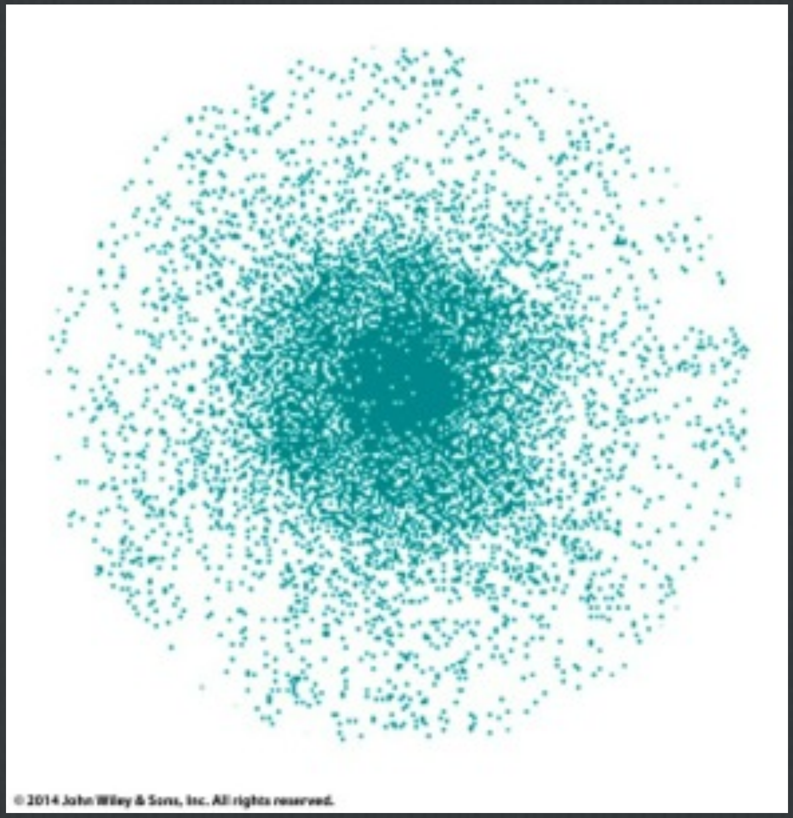
This theory, called quantum mechanics, allows one to calculate the probability of finding an electron in space.

Orbits vs Orbitals

Quantum mechanics deal in electron probabilities; orbits from Bohr theory are replaced by orbitals.

Orbitals: regions of space with a high probability of finding an electron.

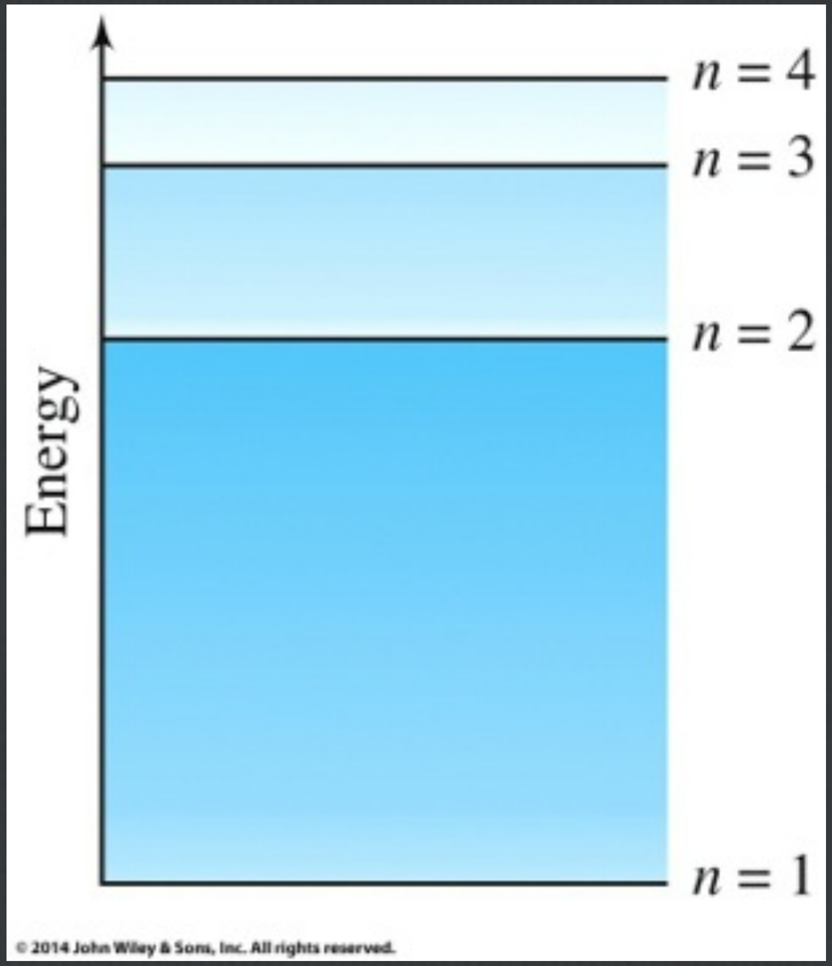




An orbital for a hydrogen atom.

Bohr's idea of quantized energy levels does have parallels in quantum mechanics.

For example, quantum mechanics predicts discrete, quantized principal energy levels for electrons in an atom.



Principal energy level (n):
provides a general idea of the distance of an electron from the nucleus
(as n increases, so does the distance from the nucleus)

n can be a positive integer (1,2,3, etc.)

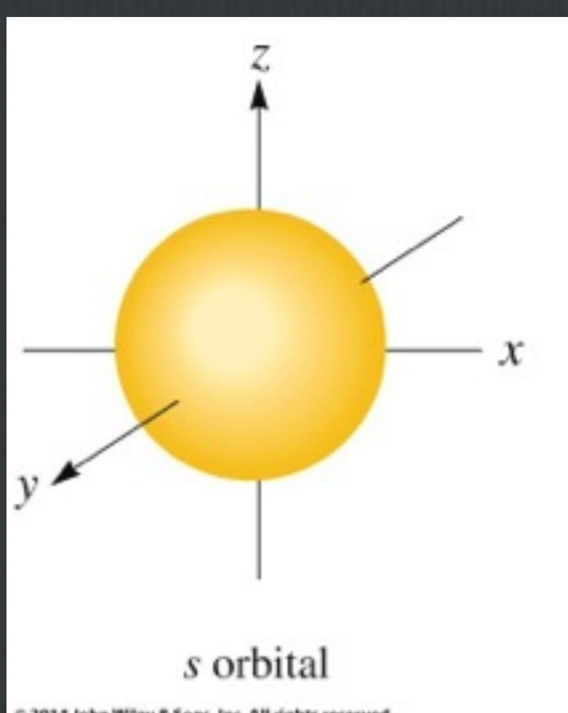
Principle Energy Levels

Each energy level can be divided into sublevels.

Each sublevel corresponds to a different type of orbital that can house electrons.

Principal quantum level n =1

Contains only one sublevel/orbital (1s orbital)



The orbital is spherical in nature, as are all s orbitals.

The orbital is a space where an electron has a 90% probability of being located.

Inside the orbital

Pauli Exclusion Principle: an atomic orbital can hold two electrons, which must have opposite spins.

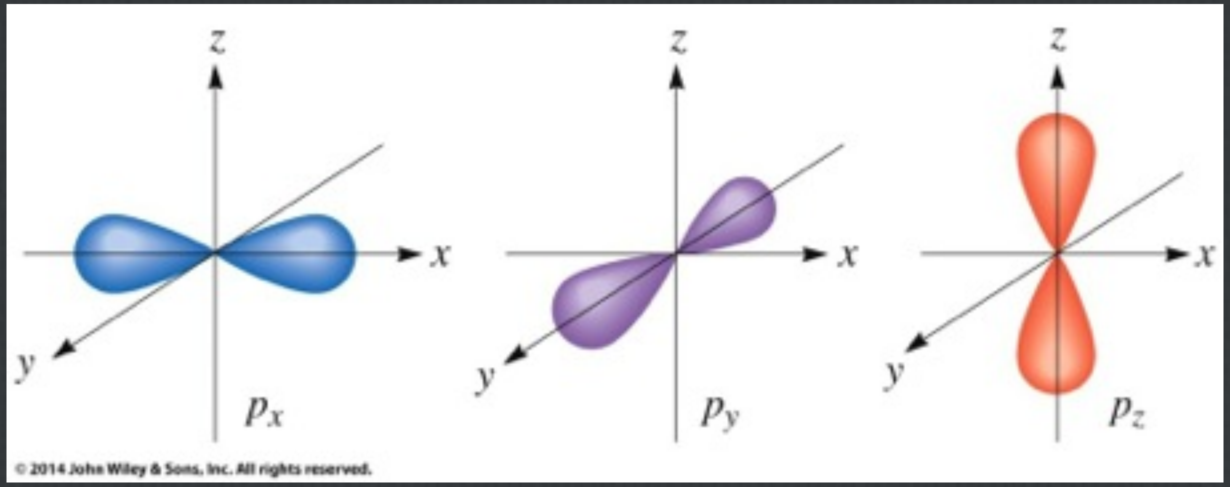
Electron spin is represented by arrows (↑ or ↓)

Principal quantum level n =2

Contains two sublevels/orbitals (2s and 2p orbitals)

The 2s orbital is spherical like a 1s orbital, but larger and higher in energy.

The 2p sublevel consists of three orbitals: px, py and pz. The shape of the 3 orbitals is the same; they differ in their orientation in space.



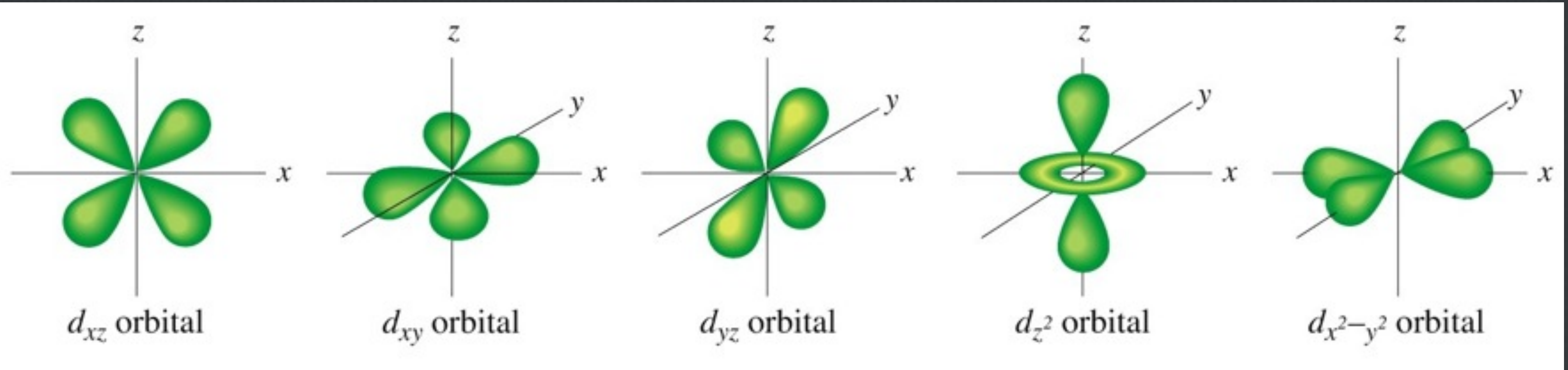
6 total electrons can occupy the 3 p orbitals of a subshell.

Each p orbital has two lobes.

Principal quantum level n =3

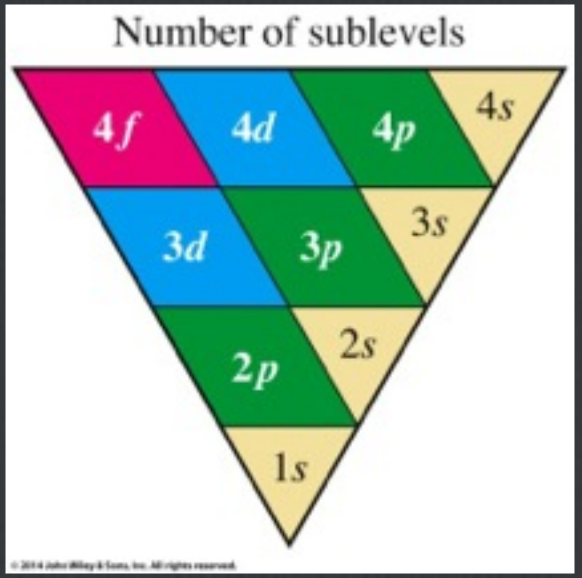
Contains three sublevels/orbitals (3s, 3p and 3d orbitals)

The 5 3d orbitals have unique shapes relative to s and p orbitals.



10 total electrons can occupy the 5 d orbitals of a subshell.

With each new principal quantum level, a new sublevel/orbital type is introduced.



Quantum Level Summary

Principal Quantum Level (n)	Sublevels			
1	1s	_____	_____	_____
2	2s	2p 2p 2p	_____	_____
3	3s	3p 3p 3p	3d 3d 3d 3d 3d	_____
4	4s	4p 4p 4p	4d 4d 4d 4d 4d	4f 4f 4f 4f 4f 4f 4f

What is the maximum number of electrons in a 4d sublevel?

a. 2

b. 4

c. 6

d. 10

Any d sublevel has 5 total orbitals.
Two electrons can be housed in each orbital.

What is the maximum number of electrons that can occupy the n = 3 sublevel?

a. 8

b. 2

c. 18

d. 10

The n = 3 sublevel has 3 types of orbitals:
s (1), p (3), and d (5).

Two electrons can occupy each orbital.
(9 x 2 = 18)

How electrons are arranged

Hydrogen: consists of a nucleus with one proton and one electron in a 1s orbital.

Assuming multielectron atoms have orbitals similar to that of the hydrogen atom, rules for filling electrons can be developed.

Guidelines for filling electrons:

1. Only up to two electrons can occupy an orbital.

2. Electron will occupy lower energy orbitals first. Orbital energies: $s < p < d < f$ for a given n value.

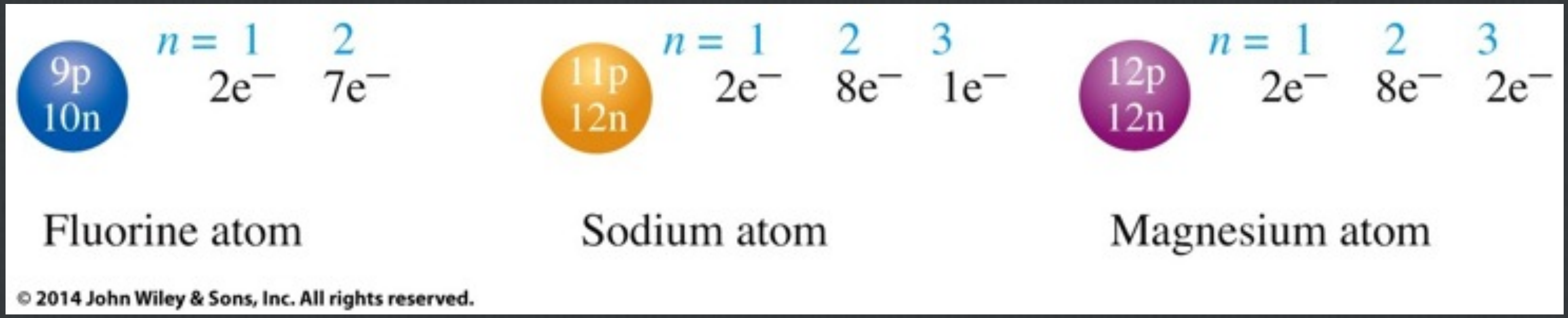
3. Each orbital in a sublevel must contain an electron before another electron can be added to any of the orbitals.

Depicting Electronic Structure

Three ways of depicting atomic and electronic structure:

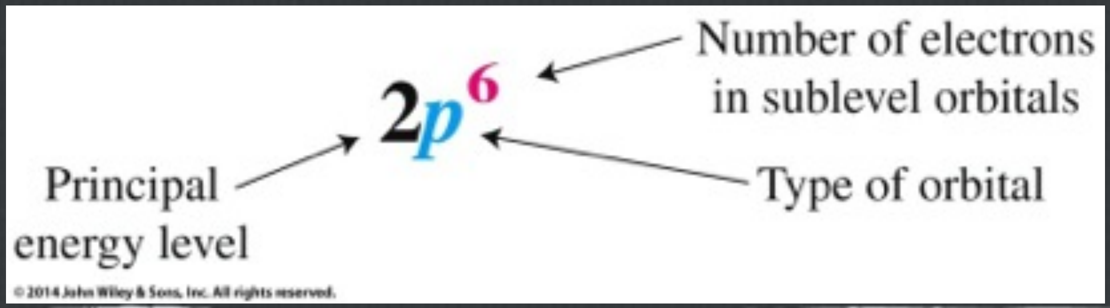
I) Atomic Structure Diagrams

Depict both the nuclear and electronic structure



II) Electron configuration

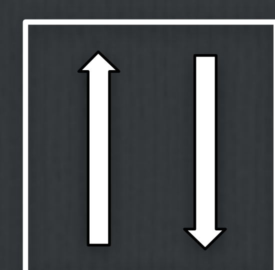
Each orbital is listed, in order of increasing energy, while depicting the number of electrons in each orbital/sublevel.



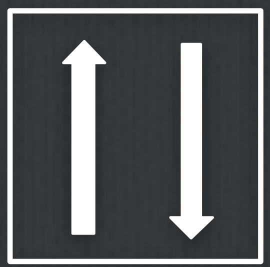
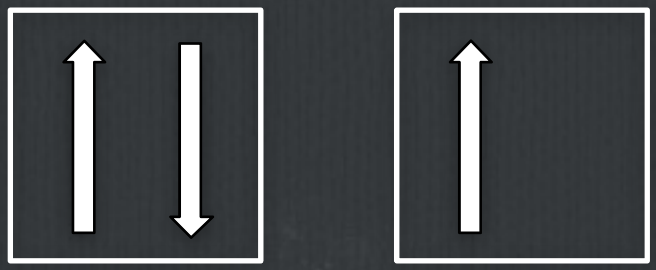
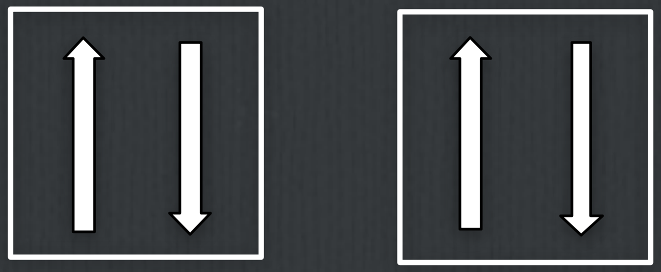
Depicting Electronic Structure

III) Orbital Diagram

Boxes represent orbitals and electron spins are represented as arrows.

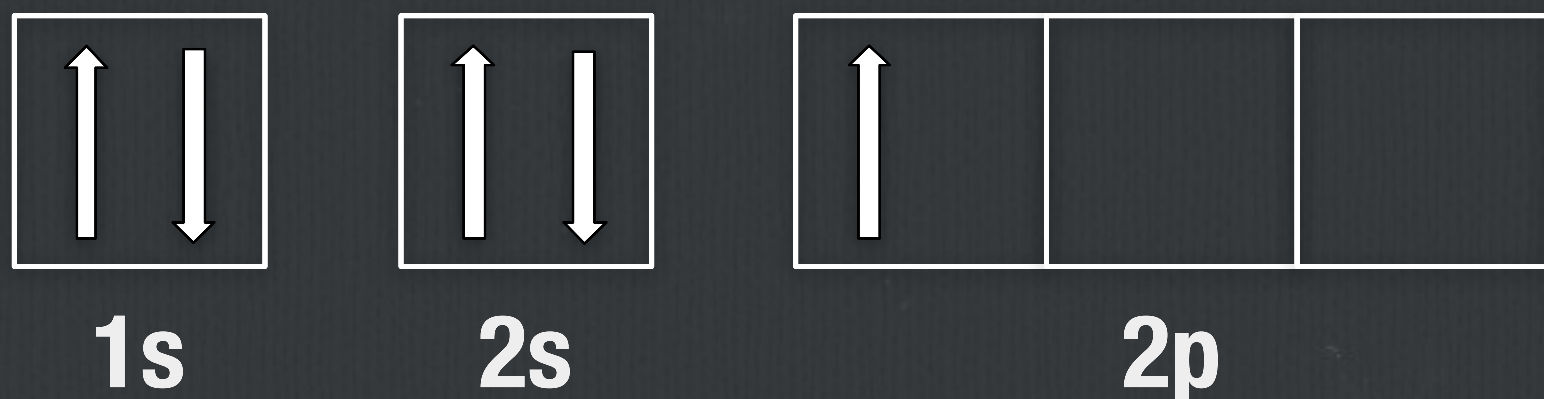


3p

	Orbital Diagram	Electron configuration
Hydrogen	 1s	$1s^1$
Helium	 1s	$1s^2$
Lithium	 1s 2s	$1s^2 2s^1$
Berylium	 1s 2s	$1s^2 2s^2$

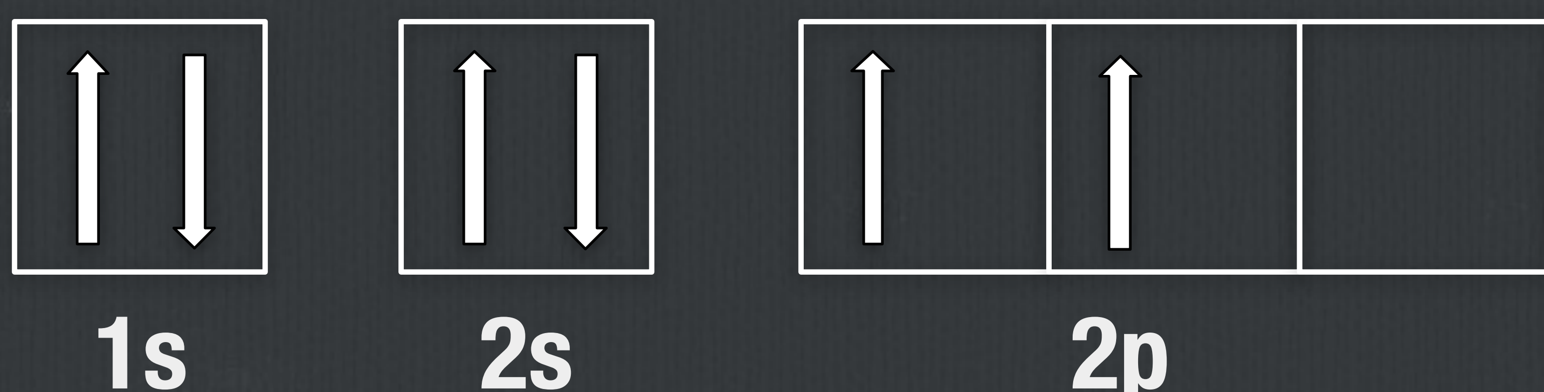
Filling the p subshell

Boron



$1s^2 2s^2 2p^1$

Carbon



$1s^2 2s^2 2p^2$

Note the electrons are in different orbitals per Rule 3.
The similar spin will be discussed shortly.

TABLE 10.1 Orbital Filling for the First Ten Elements*

Atomic number	Element	Orbitals 1s	2s	2p	Electron configuration
1	H	[↑]			$1s^1$
2	He	[↑↓]			$1s^2$
3	Li	[↑↓]	[↑]		$1s^22s^1$
4	Be	[↑↓]	[↑↓]		$1s^22s^2$
5	B	[↑↓]	[↑↓]	[↑] [ ] [ ]	$1s^22s^22p^1$
6	C	[↑↓]	[↑↓]	[↑] [↑] [ ]	$1s^22s^22p^2$
7	N	[↑↓]	[↑↓]	[↑] [↑] [↑]	$1s^22s^22p^3$
8	O	[↑↓]	[↑↓]	[↑↓] [↑] [↑]	$1s^22s^22p^4$
9	F	[↑↓]	[↑↓]	[↑↓] [↑↓] [↑]	$1s^22s^22p^5$
10	Ne	[↑↓]	[↑↓]	[↑↓] [↑↓] [↑↓]	$1s^22s^22p^6$

*Boxes represent the orbitals grouped by sublevel. Electrons are shown by arrows.

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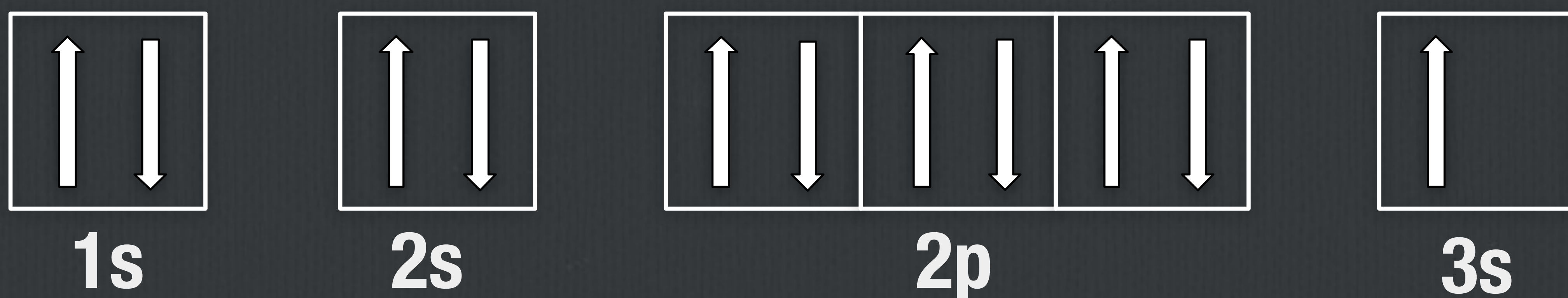
Sodium

$1s^2 2s^2 2p^6 3s^1$



Begin filling the 3n shell.

TABLE 10.2 Orbital Diagrams and Electron Configurations for Elements 11–18

Atomic number	Element	Orbitals 1s	2s	2p	3s	3p	Electron configuration
11	Na	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑		$1s^2 2s^2 2p^6 3s^1$
12	Mg	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓		$1s^2 2s^2 2p^6 3s^2$
13	Al	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ _ _	$1s^2 2s^2 2p^6 3s^2 3p^1$
14	Si	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑ _	$1s^2 2s^2 2p^6 3s^2 3p^2$
15	P	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑ ↑ ↑	$1s^2 2s^2 2p^6 3s^2 3p^3$
16	S	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑ ↑	$1s^2 2s^2 2p^6 3s^2 3p^4$
17	Cl	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑	$1s^2 2s^2 2p^6 3s^2 3p^5$
18	Ar	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	$1s^2 2s^2 2p^6 3s^2 3p^6$

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Valence Electrons

Valence Electrons: electrons located in the highest energy (outermost) orbitals of an atom.

Example: Oxygen

Electron configuration

$1s^2 2s^2 2p^4$

Outermost valence electrons are in the n =2 subshell; oxygen has 6 total valence electrons.

The column number (1A-7A) in the periodic table gives the valence electrons of an element.

Valence electrons participate in bonding to form molecular compounds.

How many valence electrons does phosphorus contain?

a. 5

b. 8

c. 15

d. 18

P is in Group 5A.

It will contain 5 valence electrons.

Electron Configuration

$1s^2 2s^2 2p^6 3s^2 3p^3$

Which element could have the following electron configuration?

$1s^2 2s^2 2p^6 3s^2 3p^6$

a. Cl

b. Ca

c. Ar

d. S

The element contains 18 electrons.

It must also contain 18 protons.

Atomic number 18 = argon.

How does the electronic structure of atoms relate to their positioning on the periodic table?

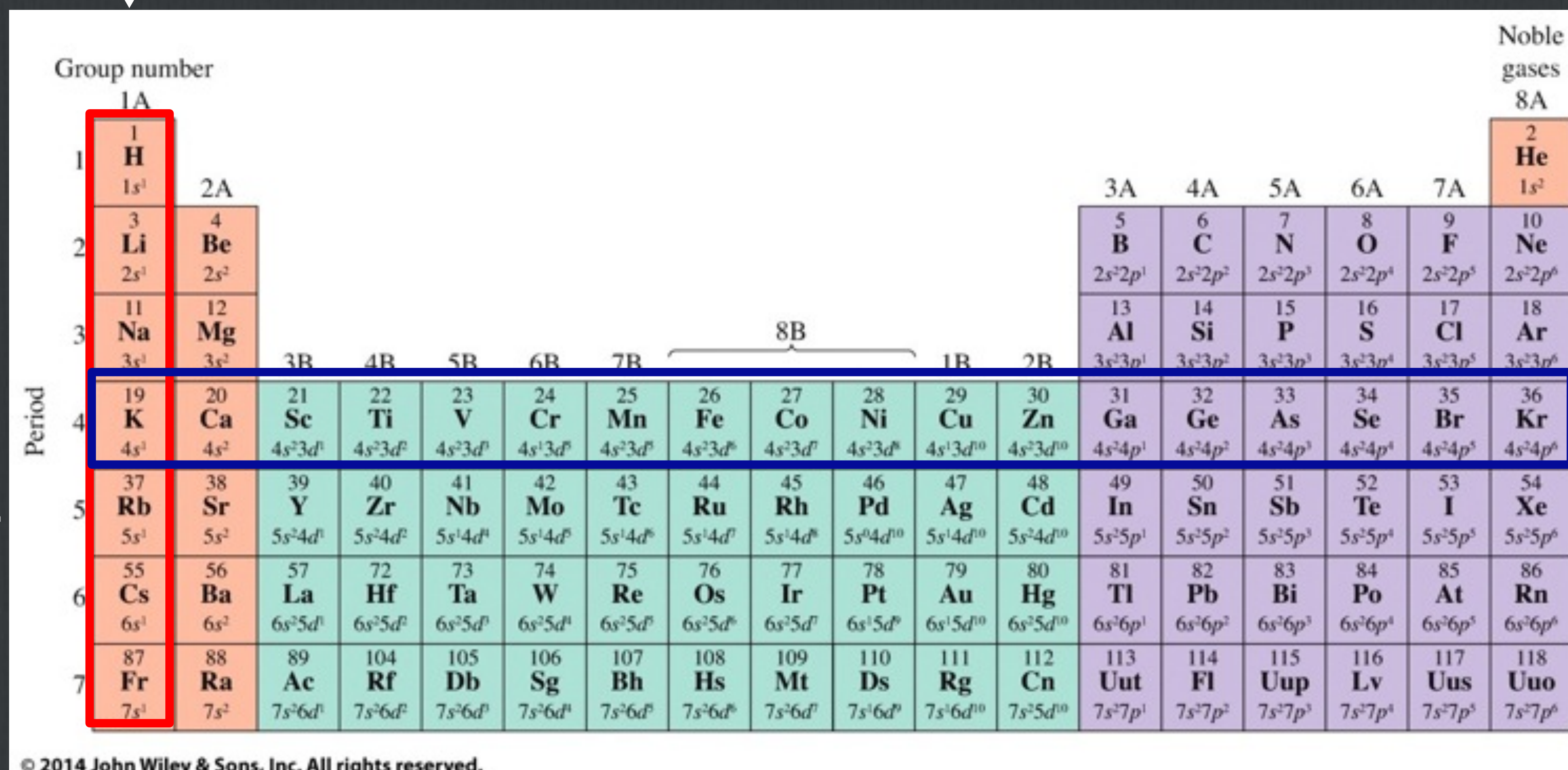
Origins of the Periodic Table

Mendeleev and Meyer proposed organizing elements based on increasing atomic masses.
Their ideas formed the origins of the modern periodic table.

How does the electronic structure of atoms relate to their positioning on the periodic table?

columns (groups/families): elements with similar properties.

rows (periods): increasing Z (atomic number)

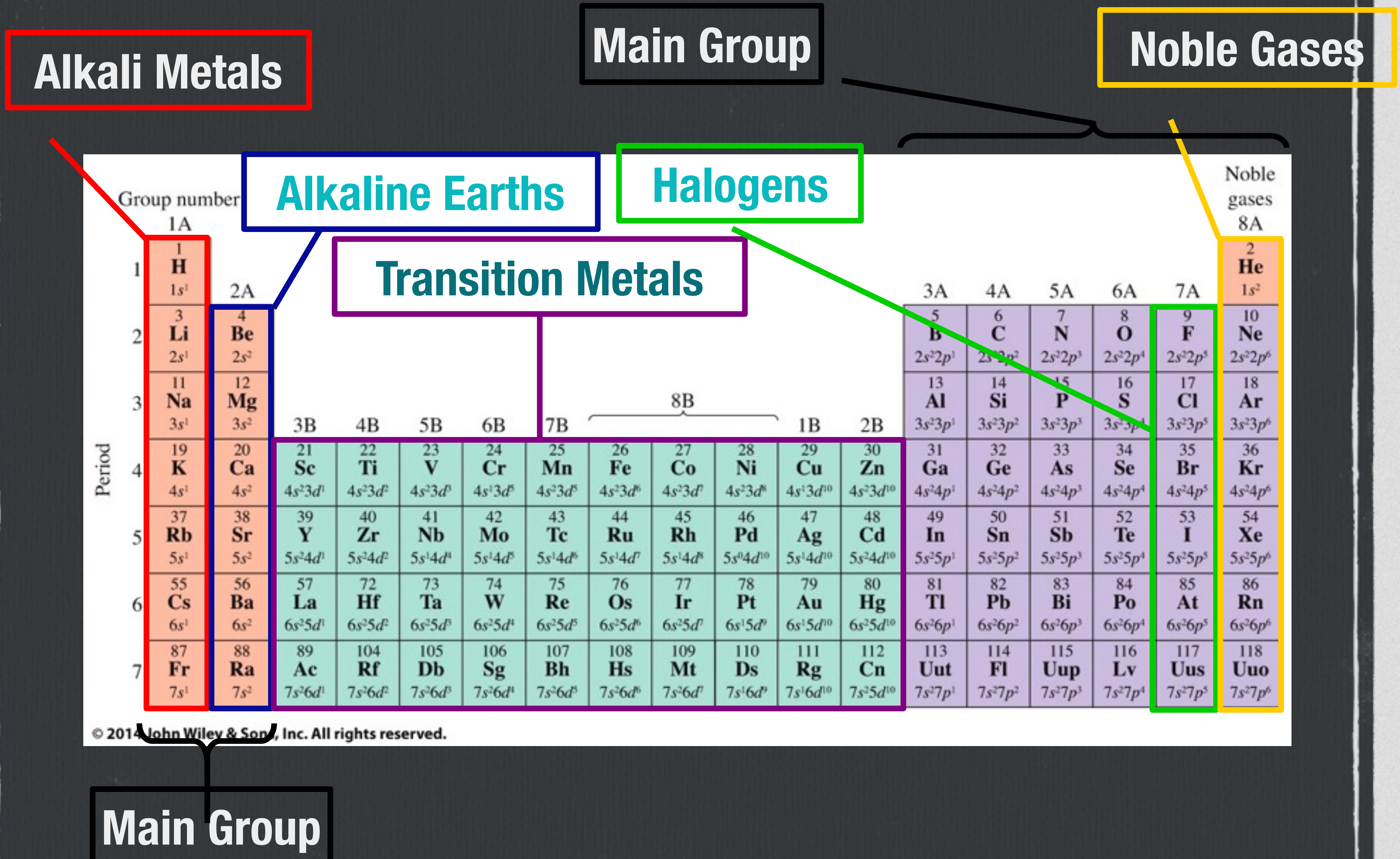


Period number corresponds to the highest principal quantum number n of the elements.

Group numbering: by numbers 1–18 or with numbers/letters (i.e., 7A).

A groups are the representative elements.

B groups are the transition elements.



Alkali Metals
Main Group
Noble Gases
Alkaline Earths
Halogens
Transition Metals
Main Group
Group number
Period
1A
2A
3B 4B 5B 6B 7B 8B 1B 2B
3A 4A 5A 6A 7A
Noble gases 8A
1: 1 H $1s^1$ | 2 He $1s^2$
2: 3 Li $2s^1$ | 4 Be $2s^2$ | 5 B $2s^22p^1$ | 6 C $2s^22p^2$ | 7 N $2s^22p^3$ | 8 O $2s^22p^4$ | 9 F $2s^22p^5$ | 10 Ne $2s^22p^6$
3: 11 Na $3s^1$ | 12 Mg $3s^2$ | 13 Al $3s^23p^1$ | 14 Si $3s^23p^2$ | 15 P $3s^23p^3$ | 16 S $3s^23p^4$ | 17 Cl $3s^23p^5$ | 18 Ar $3s^23p^6$
4: 19 K $4s^1$ | 20 Ca $4s^2$ | 21 Sc $4s^23d^1$ | 22 Ti $4s^23d^2$ | 23 V $4s^23d^3$ | 24 Cr $4s^13d^5$ | 25 Mn $4s^23d^5$ | 26 Fe $4s^23d^6$ | 27 Co $4s^23d^7$ | 28 Ni $4s^23d^8$ | 29 Cu $4s^13d^{10}$ | 30 Zn $4s^23d^{10}$ | 31 Ga $4s^24p^1$ | 32 Ge $4s^24p^2$ | 33 As $4s^24p^3$ | 34 Se $4s^24p^4$ | 35 Br $4s^24p^5$ | 36 Kr $4s^24p^6$
5: 37 Rb $5s^1$ | 38 Sr $5s^2$ | 39 Y $5s^24d^1$ | 40 Zr $5s^24d^2$ | 41 Nb $5s^14d^4$ | 42 Mo $5s^14d^5$ | 43 Tc $5s^14d^6$ | 44 Ru $5s^14d^7$ | 45 Rh $5s^14d^8$ | 46 Pd $5s^04d^{10}$ | 47 Ag $5s^14d^{10}$ | 48 Cd $5s^24d^{10}$ | 49 In $5s^25p^1$ | 50 Sn $5s^25p^2$ | 51 Sb $5s^25p^3$ | 52 Te $5s^25p^4$ | 53 I $5s^25p^5$ | 54 Xe $5s^25p^6$
6: 55 Cs $6s^1$ | 56 Ba $6s^2$ | 57 La $6s^25d^1$ | 72 Hf $6s^25d^2$ | 73 Ta $6s^25d^3$ | 74 W $6s^25d^4$ | 75 Re $6s^25d^5$ | 76 Os $6s^25d^6$ | 77 Ir $6s^25d^7$ | 78 Pt $6s^15d^9$ | 79 Au $6s^15d^{10}$ | 80 Hg $6s^25d^{10}$ | 81 Tl $6s^26p^1$ | 82 Pb $6s^26p^2$ | 83 Bi $6s^26p^3$ | 84 Po $6s^26p^4$ | 85 At $6s^26p^5$ | 86 Rn $6s^26p^6$
7: 87 Fr $7s^1$ | 88 Ra $7s^2$ | 89 Ac $7s^26d^1$ | 104 Rf $7s^26d^2$ | 105 Db $7s^26d^3$ | 106 Sg $7s^26d^4$ | 107 Bh $7s^26d^5$ | 108 Hs $7s^26d^6$ | 109 Mt $7s^26d^7$ | 110 Ds $7s^16d^9$ | 111 Rg $7s^16d^{10}$ | 112 Cn $7s^25d^{10}$ | 113 Uut $7s^27p^1$ | 114 Fl $7s^27p^2$ | 115 Uup $7s^27p^3$ | 116 Lv $7s^27p^4$ | 117 Uus $7s^27p^5$ | 118 Uuo $7s^27p^6$
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Magnesium is a member of what group of elements?

a. Halogens

b. Noble gases

c. Alkali metals

d. Alkaline earth metals

Group Commonalities: All elements in a group have similar valence electron configurations.

This often leads to similar reactivity of elements in a group.

Because electron configurations of larger elements are very long, a shorthand notation is used that highlights valence electrons.

Example Na [Ne]3s1

This indicates the electron configuration of sodium is that for Ne with an additional 3s1 electron.

The shorthand electron configuration always uses the previous noble gas core with the valence electrons for the element of interest.

mple

Initially, one may expect to fill 3d electrons before 4s electrons.

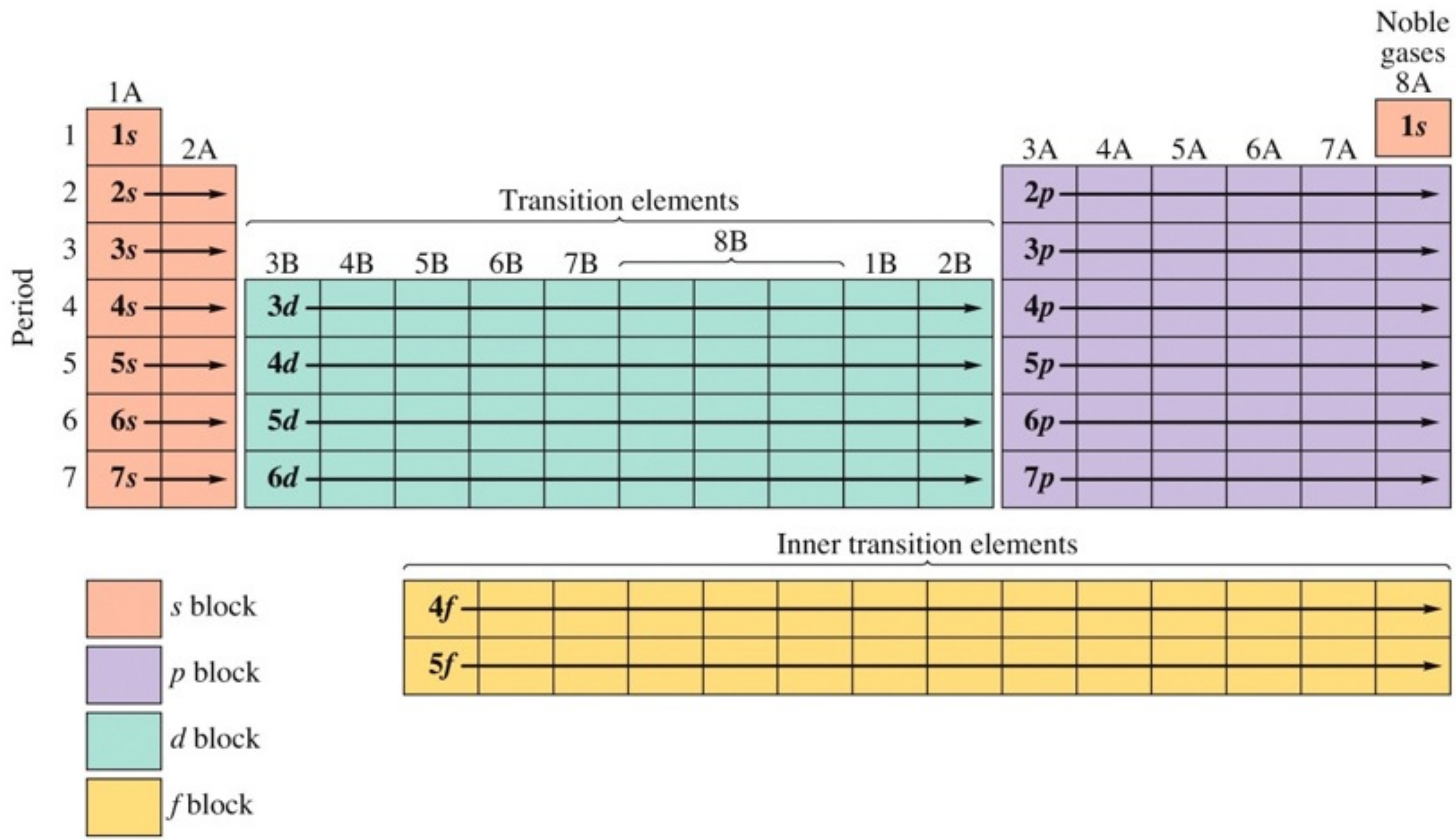
However, based on reactivity of the elements K and Ca, they resemble metals with valence electrons in s orbitals.

K	[Ar]4s1
Ca	[Ar]4s2

Elements 21-30 are transition elements where electrons now fill in the 3d subshell.

Once the 3d subshell is filled, the 4p shell begins to fill.

Sublevel Filling Diagram



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The electron configuration [Ar] 4s1 is the ground state electron configuration of:

- a. K
- b. P
- c. Fluorine
- d. Na

The element contains 1 valence electron in the fourth period (due to n = 4).

The electron configuration [Ne] 3s23p1, is the ground state electron configuration of:

a. Na

b. Al

c. Ar

d. S

The element contains 3 valence electrons in the third period (due to n = 3).