Chemistry 120 Fall 2016

Instructor: Dr. Upali Siriwardane

e-mail: upali@latech.edu
Office: CTH 311
Phone 257-4941

Office Hours: M,W,F 9:30-11:30 am T,R 8:00-10:00 am or by appointment;

Test Dates:

September 23, **2016 (Test 1): Chapter 1,2 &3**
October 13, **2016 (Test 2): Chapter 4 & 5**
October 31, **2016 (Test 3): Chapter 6, 7 & 8**
November 15, **2016 (Test 4): Chapter 9, 10 & 11**
November 17, **2016 (Make-up test) comprehensive: Chapters 1-11**

Chapter 3. Atomic Structure and the Periodic Table

3-1 Internal Structure of an Atom

Arrangement of Subatomic Particles Within an Atom

Charge Neutrality of an Atom

Size Relationships Within an Atom

3-2 Atomic Number and Mass Number

Electrons and Chemical Properties

3-3 Isotopes and Atomic Masses

Isotopes

Atomic Masses

3-4 **The Periodic Law and the Periodic Table**

Groups and Periods of Elements

The Shape of the Periodic Table

3-5 Metals and Nonmetals

Periodic Table Locations for Metals and Nonmetals

3-6 Electron Arrangements Within Atoms

Electron Shells

Electron Subshells

Electron Orbitals

Electron Spin

Chapter 3. Atomic Structure and the Periodic Table

3-7 Electron Configurations and Orbital Diagrams

Subshell Energy Order

Writing Electron Configurations and Orbital Diagrams

3-8 The Electronic Basis for the Periodic Law and the Periodic Table

Electron Configurations and the Periodic Law

Electron Configurations and the Periodic Table

3-9 Classification of the Elements

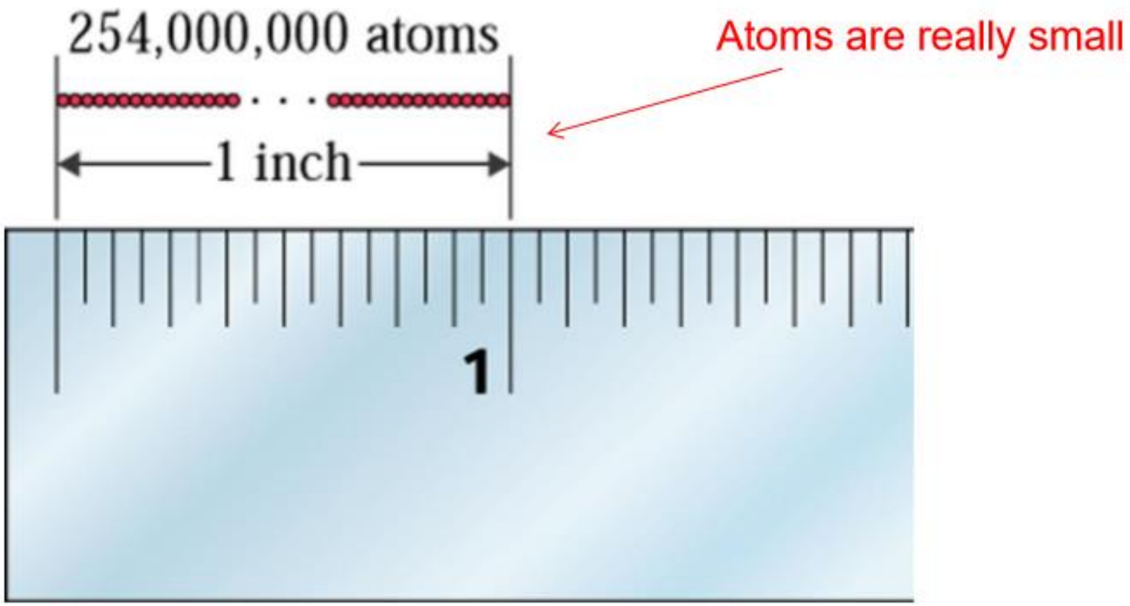
What’s covered in this chapter?

- Structure of an atom
- Atomic number and mass number
- Isotopes and (average) atomic masses
- Periodic table
- How are electrons arranged in atoms?
- How to describe electron arrangements for each element
- Metals, non-metals, and metalloids

Atoms

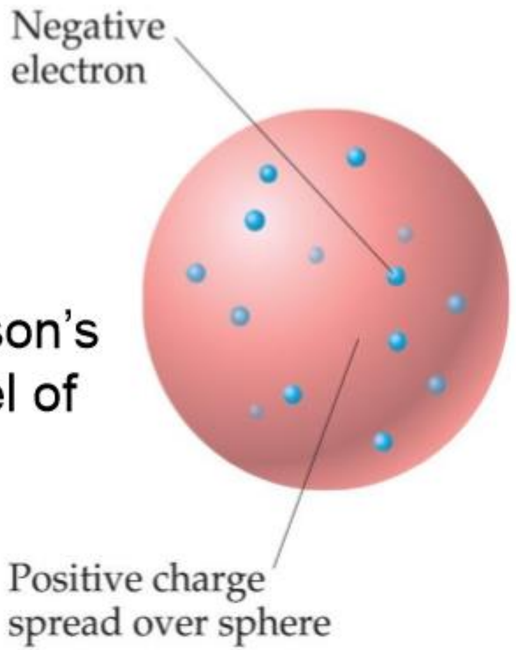
- What are atoms?

– Chapter 1: the smallest particle of an element that can exist and retain all of the properties of the element

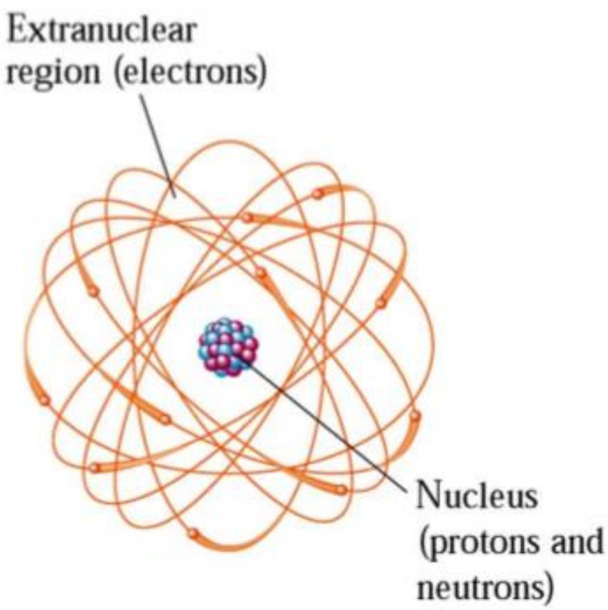


The Nuclear Atom

- The current model of the atom predicts a very small, dense nucleus with the electrons around the outside of the atom.
- Most of the volume of the atom is empty space.

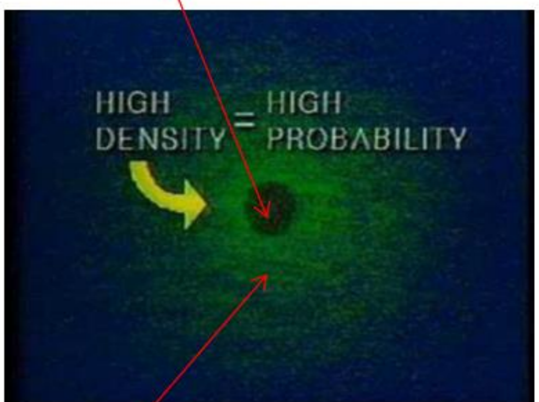


Early 1900's: Thompson's "Plum Pudding" model of the atom.



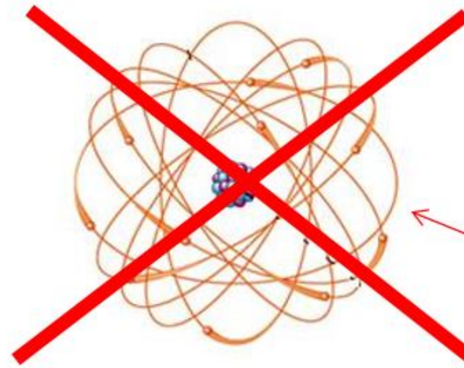
Atomic Structure

dense core (nucleus) in center

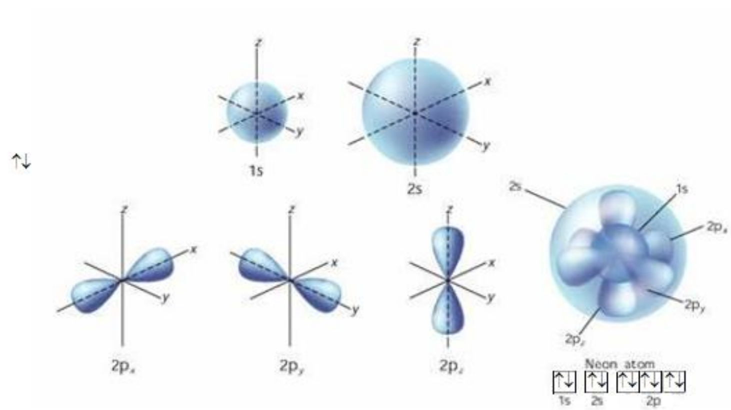


electron "cloud" region around nucleus

Bohr model
of the atom



"orbits"



Subatomic Particles

Atoms are made of three types of subatomic particles: protons, electrons, and neutrons.

- Protons (+) and electrons (-) are the only particles that have a charge.
- Protons and neutrons are found in the nucleus (core); electrons reside outside of the nucleus. Thus protons and neutrons are sometimes called "nucleons"

"atomic mass unit"

Particle	Charge	Mass (amu)
Proton	Positive (1+)	1.0073
Neutron	None (neutral)	1.0087
Electron	Negative (1−)	5.486×10^{-4}

1 amu = 1.66054×10^{-24} g

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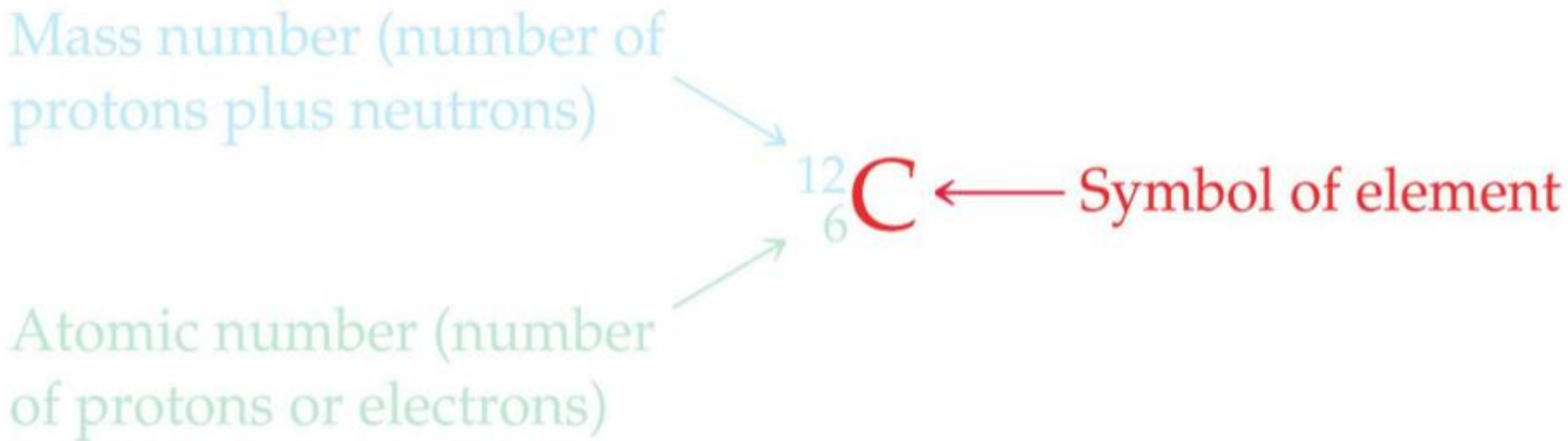
Symbols of Elements

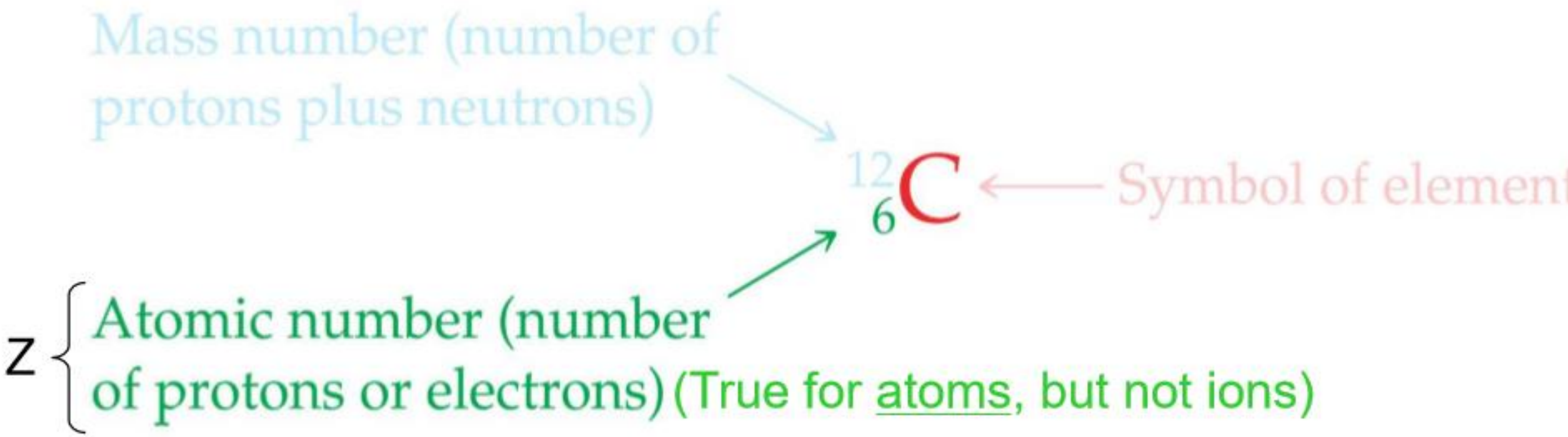
An element's symbol gives information about the number of protons, neutrons, and electrons in an atom or ion

Q: What distinguishes atoms of one element from those of another element?



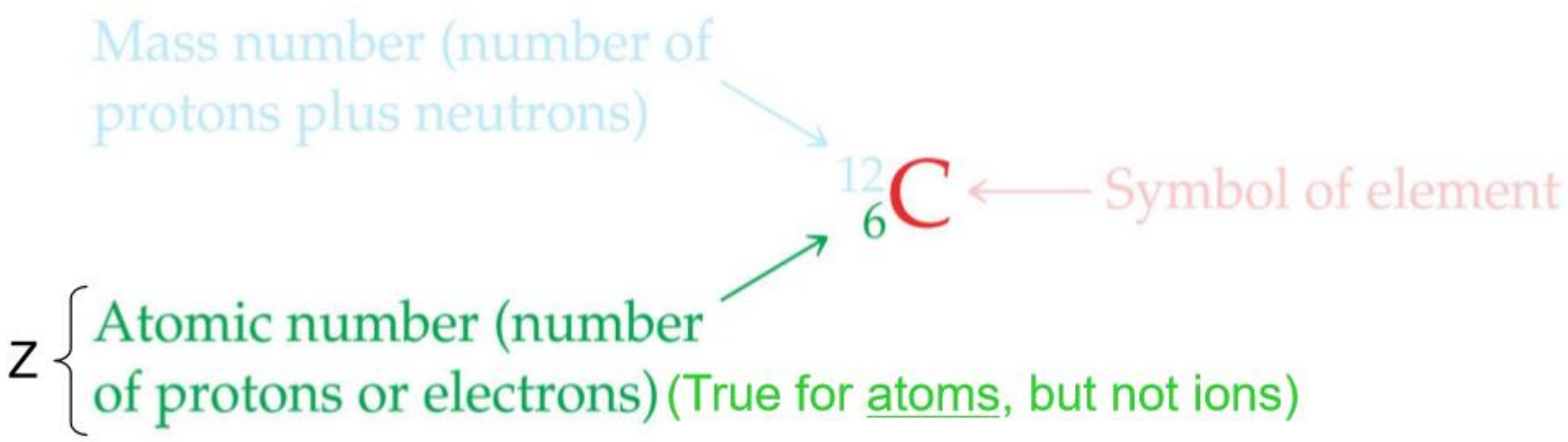
As you know, elements are symbolized by one or two letters (e.g. C, Pb, Na).

Atomic Number



All atoms (and as we'll see later, ions) of the same element have the same number of protons. The atomic number (Z) represents the number of protons in the nucleus of an atom of some element.

Mass Number



The mass of an atom in atomic mass units (amu) is the sum of the number of protons and neutrons (i.e. total number of nucleons) in the atom. The mass number for an element is given the symbol, "A"

An element's atomic number and mass number can be found in the periodic table

Chemical symbols

So the following system is used for quick identification of elements (where "X" is the chem symbol of some element in the periodic table):

$$^{A}_{Z}X$$

Q: If X = O (oxygen, element 8, mass = 16.00 ar

- How many electrons are in an atom of O?
- How many neutrons are in an atom of O?

Q: How many neutrons are in an atom of $^{23}_{11}Na$?

Isotopes

- In nature, it is frequently the case that not all atoms within a sample of an element have the same number of neutrons (this means atoms of the same element with different mass numbers).
- These "different mass" versions are called "isotopes"

Accounts for ~1.07% of all carbon found in nature

mass = 12 amu (exactly)

${}^{11}_{6}C$ ${}^{12}_{6}C$ ${}^{13}_{6}C$ ${}^{14}_{6}C$

mass = 13.00335 amu

Accounts for ~98.93% of all carbon found in nature

1 amu = 1.66054×10^{-24} g

Average Atomic Mass

- Because in the real world we use large amounts of atoms and molecules, we use average masses in calculations.
- Average mass is calculated from the isotopes of an element weighted by their relative abundances.

Example: the average mass of a carbon atom could be found through the following calculation (a weighted average):

(% ^{11}C)(mass of ^{11}C atom)
+ (% ^{12}C)(mass of ^{12}C atom)
+ (% ^{13}C)(mass of ^{13}C atom)
+ (% ^{14}C)(mass of ^{14}C atom)

= (~0)(11.01143 amu)
+ (0.9893)(12 amu)
+ (0.0107)(13.00335 amu)
+ (~0)(14.003241 amu)

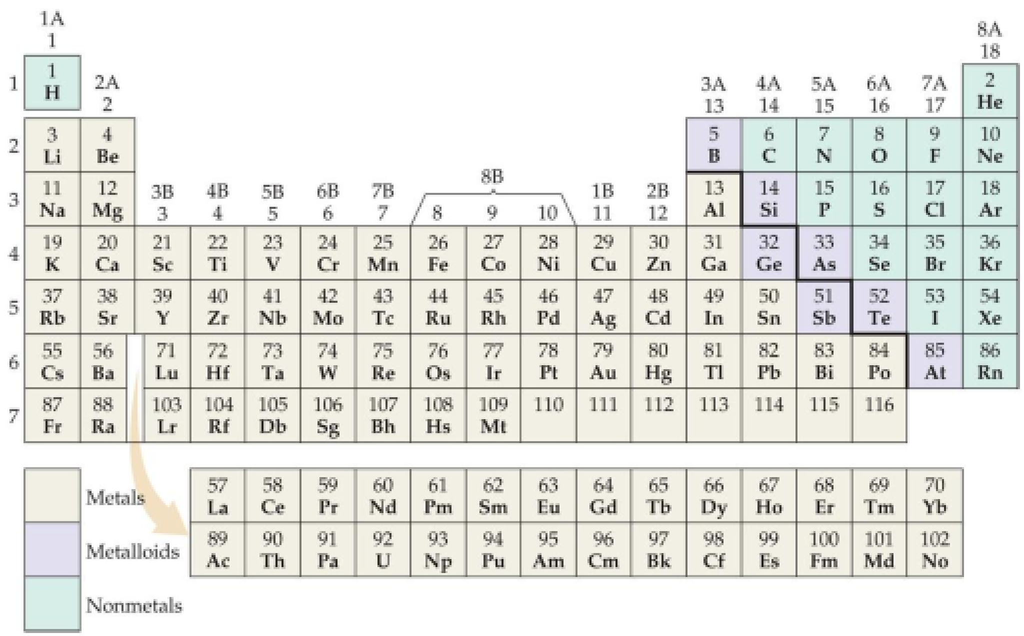
Need to use order of operations here to get sig figs right

= **12.01 amu**

This is why periodic table atomic masses are not whole numbers

IMPORTANT: Percentages expressed as decimals here

Periodic Table



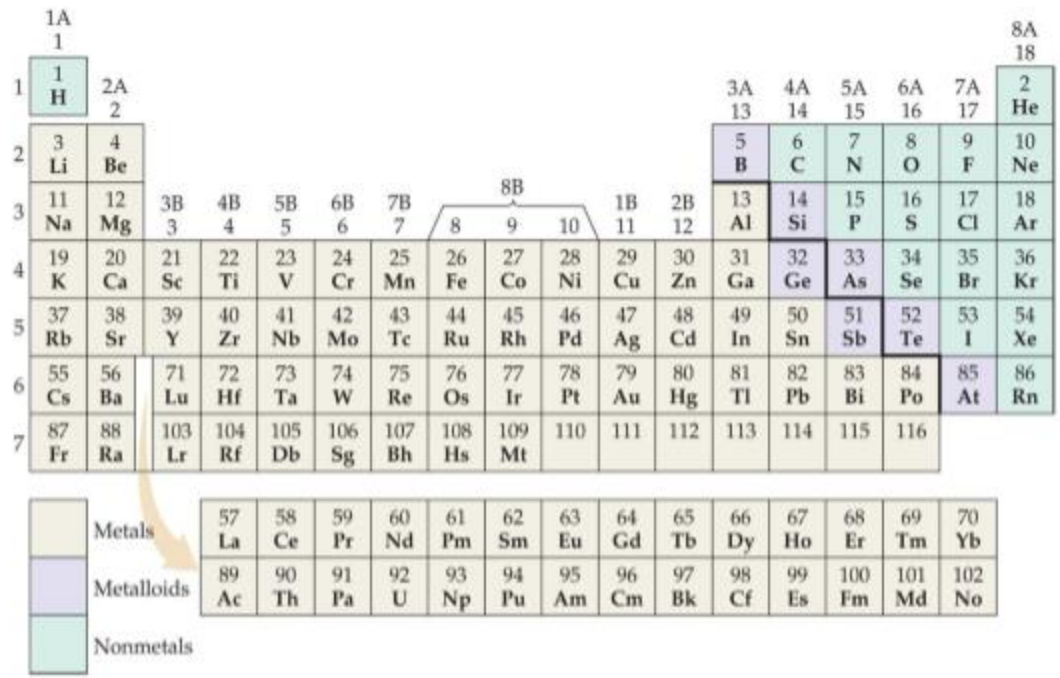
- A systematic catalog of elements.
- Developed by Mendeleev and Meyer, independently.
- Elements are arranged in order of atomic number.

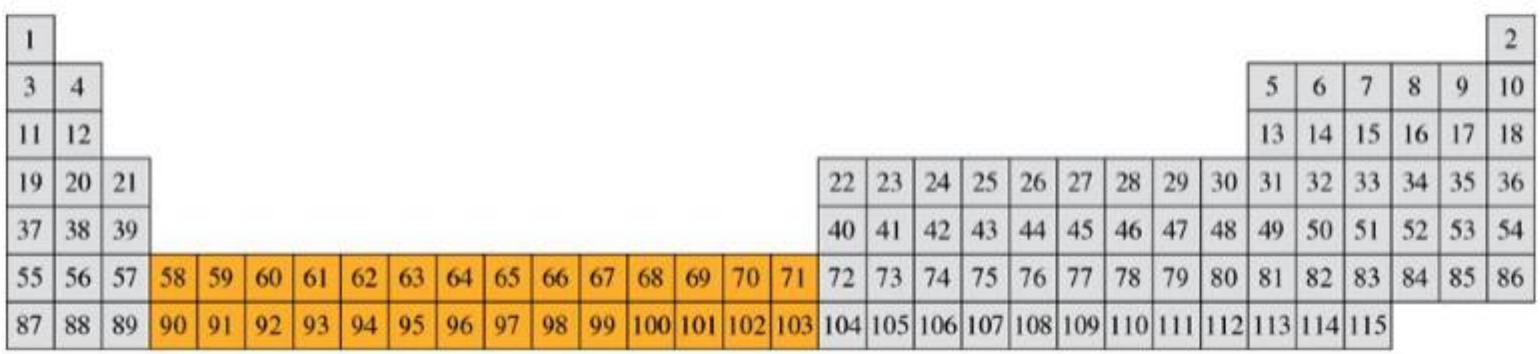
Rows = "period #"
Columns = "group"

Group 1: "alkali metals"
Group 2: "alkaline earth metals"
Group 7: "halogens"
Group 8: "noble gases"

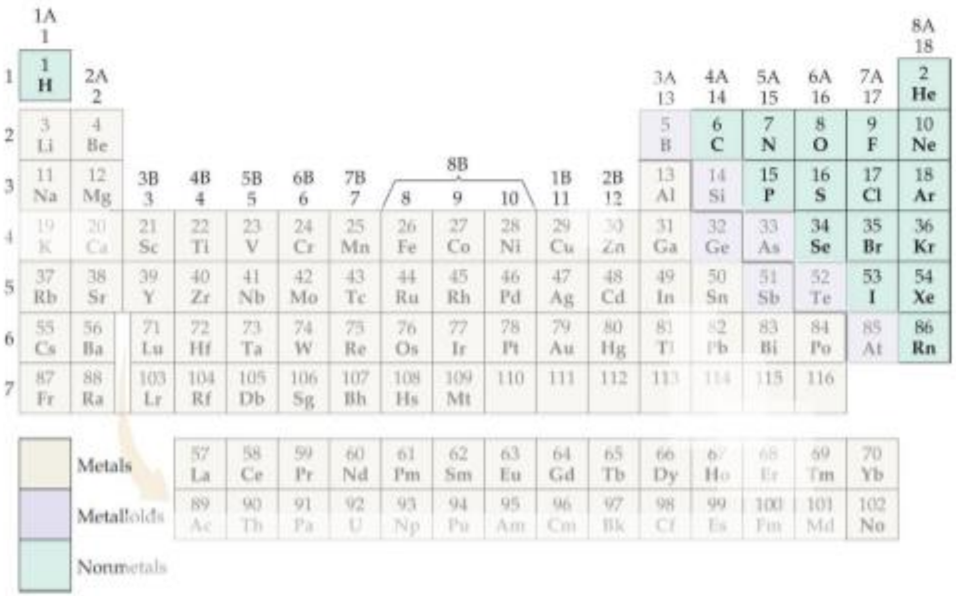
Save as

Periodic Table





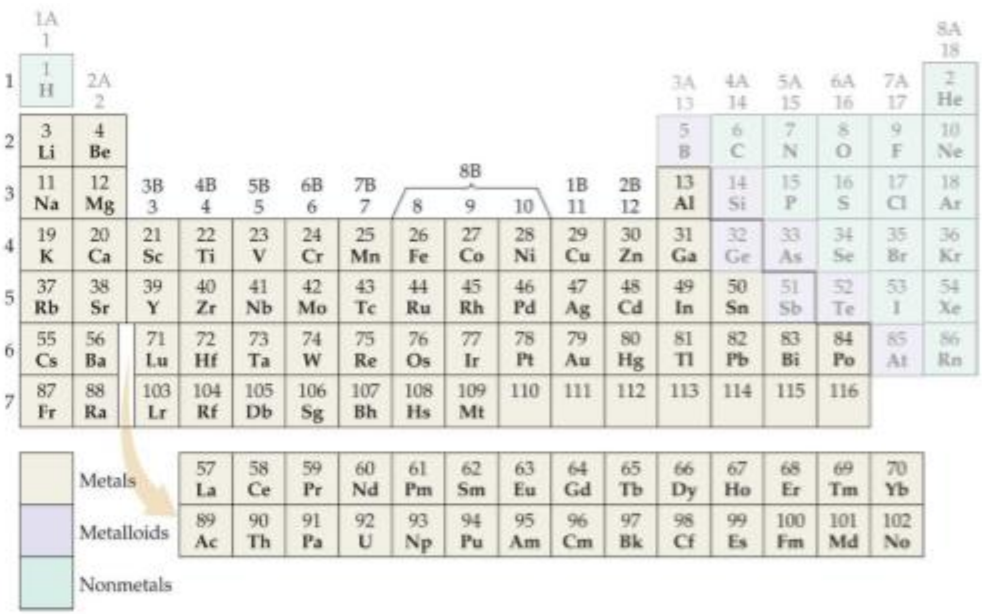
Periodic Table



Nonmetals are on the right side of the periodic table (with the exception of H).

Non-metals: absence of luster, thermal conductivity, electrical conductivity, and malleability/ductility. Tend to be found as solids, liquids, and gases

Periodic Table

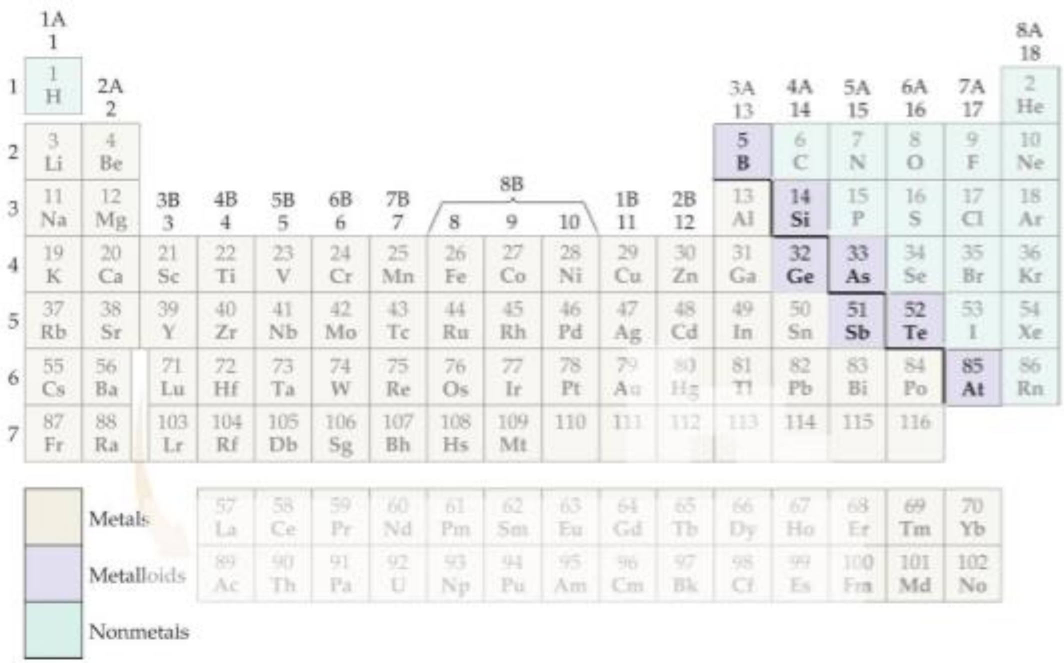


Metals appear on the left side of the periodic table.

Metals: possess characteristic properties of luster, thermal conductivity, electrical conductivity, and malleability/ ductility.

Tend to be found as solids, with few exceptions

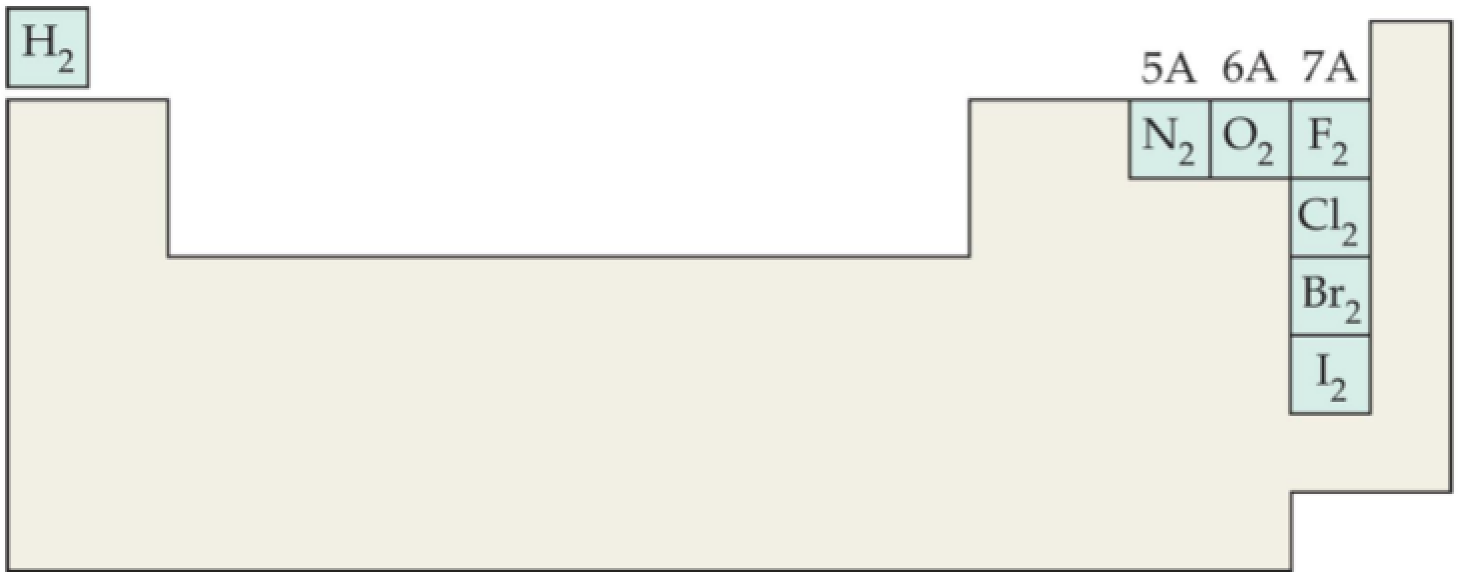
Periodic Table



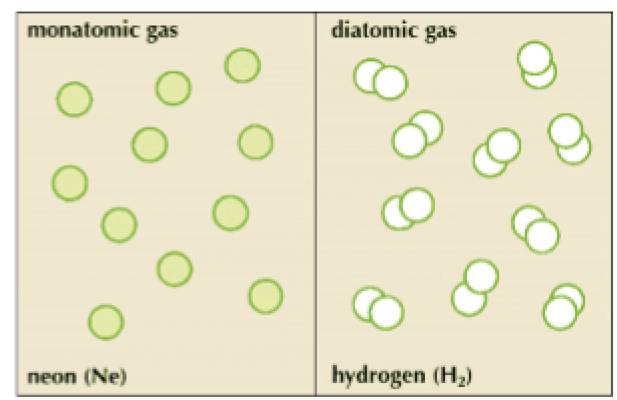
Metalloids border the stair-step line (with the exception of Al and Po).

Silicon

Certain non-metals exist in nature as diatomic molecules



These seven elements occur naturally as molecules containing two atoms.

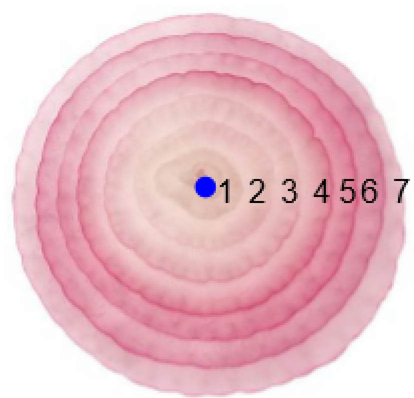


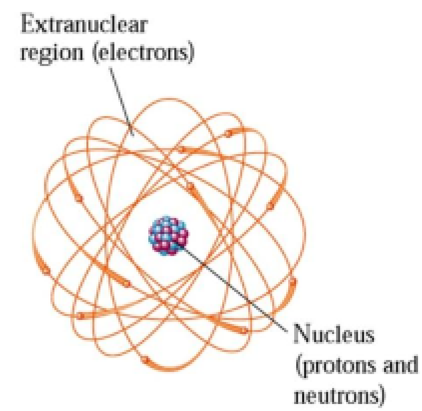
Can remember these using the phrase "Hoffbrinkle" (HOFBrINCl)

How are electrons arranged in the quantum mechanical atom?

Electron arrangements in atoms

- Electrons (abbreviated as "e^-") are contained within "shell" arrangements, around the nucleus. *An electron shell is a region of space that contains electrons that have approximately the same energy and which spend most of their time about the same distance from the nucleus*.
- Shells are numbered 1, 2, 3, … etc., in order of increasing energy. The larger the shell number, the larger the shell and (generally) the more electrons it may hold.

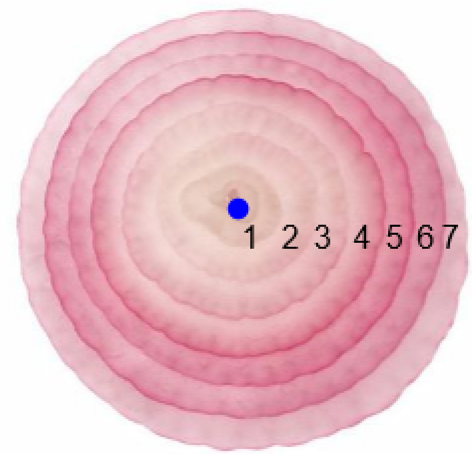




Electron shells co-exist within the same regions of space

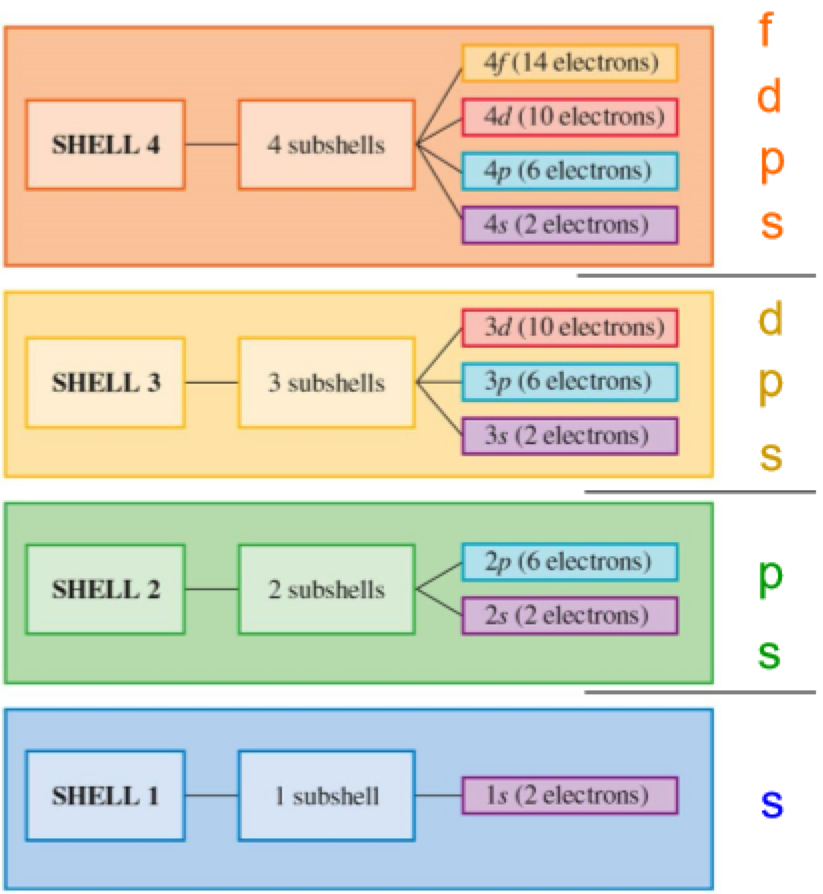
Electron shells and energy

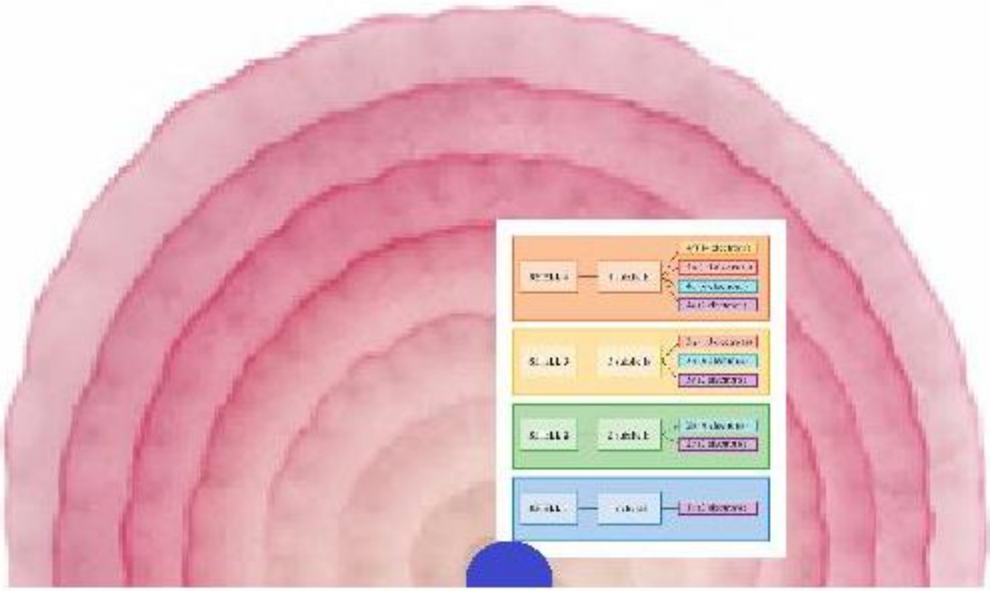
- The closer an electron is to the nucleus, the *lower* its energy becomes.
- Electrons that can exist farther away from the nucleus (in bigger electron shells) have *higher* energy.
- <u>Higher energy means lower stability</u>, so electrons that are farthest from the nucleus are the ones that are involved when atoms participate in chemical reactions. (We'll call these the "valence electrons" later.)

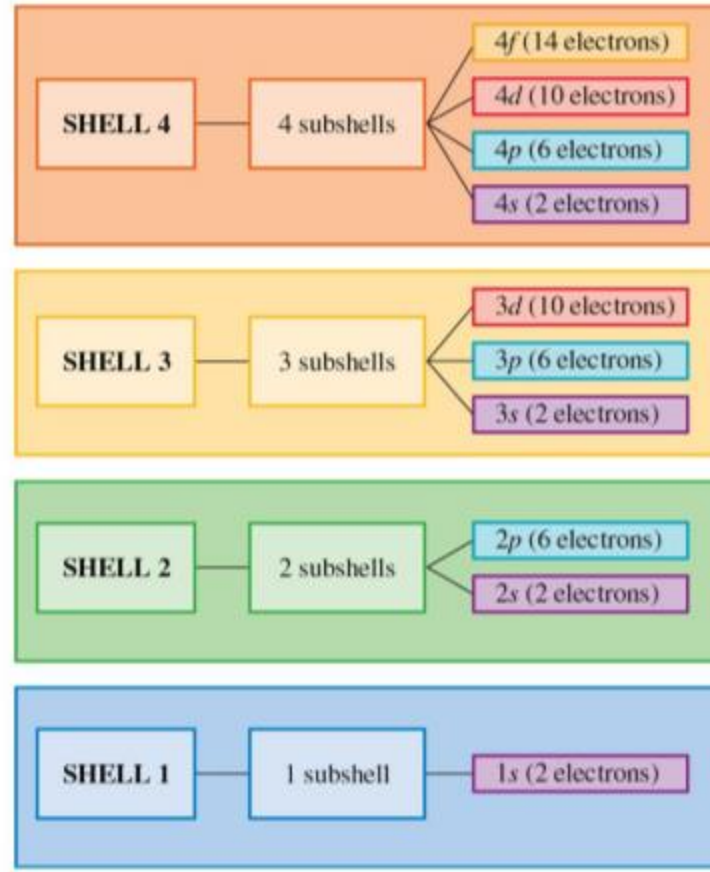


Electron subshells

- Electron *subshells* *are regions of space within electron shells that hold electrons of equivalent energy*.
- The number of subshells within a shell is the same as the shell's number
- Subshell number is designated with a letter-labeling system: s, p, d, f.
- Low → high energy: s<p<d<f
- Both a letter and a number are used to identify a subshell. (e.g. 2p subshell, 3d subshell)
- Subshells can hold different numbers of electrons:
  - s: **2** electrons
  - p: **6** electrons
  - d: **10** electrons
  - f: **14** electrons



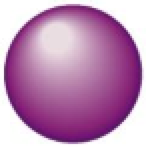



SHELL 4
4 subshells
4f (14 electrons)
4d (10 electrons)
4p (6 electrons)
4s (2 electrons)
SHELL 3
3 subshells
3d (10 electrons)
3p (6 electrons)
3s (2 electrons)
SHELL 2
2 subshells
2p (6 electrons)
2s (2 electrons)
SHELL 1
1 subshell
1s (2 electrons)

Orbitals

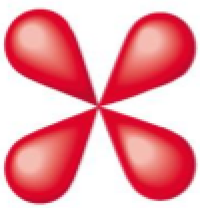
Each of these pictures is one orbital (i.e. the picture for "b" describes *one* orbital, not two)

- Subshells contain orbitals. Orbitals are the containers that hold electrons (*up to* 2 electrons, maximum).
- An *electron orbital is a region of space within an electron subshell where an electron having a specific energy can be found*.
- s-subshells are each comprised of one **s-orbital**
- p-subshells are each comprised of three **p-orbital**s
- d-subshells are each comprised of five **d-orbitals**
- f-subshells are each comprised of seven **f-orbital**s

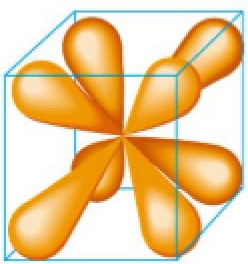


(a) *s* orbital

(b) *p* orbital



(c) *d* orbital



(d) *f* orbital

s: **2** electrons
p: **6** electrons
d: **10** electrons
f: **14** electrons

Orbitals

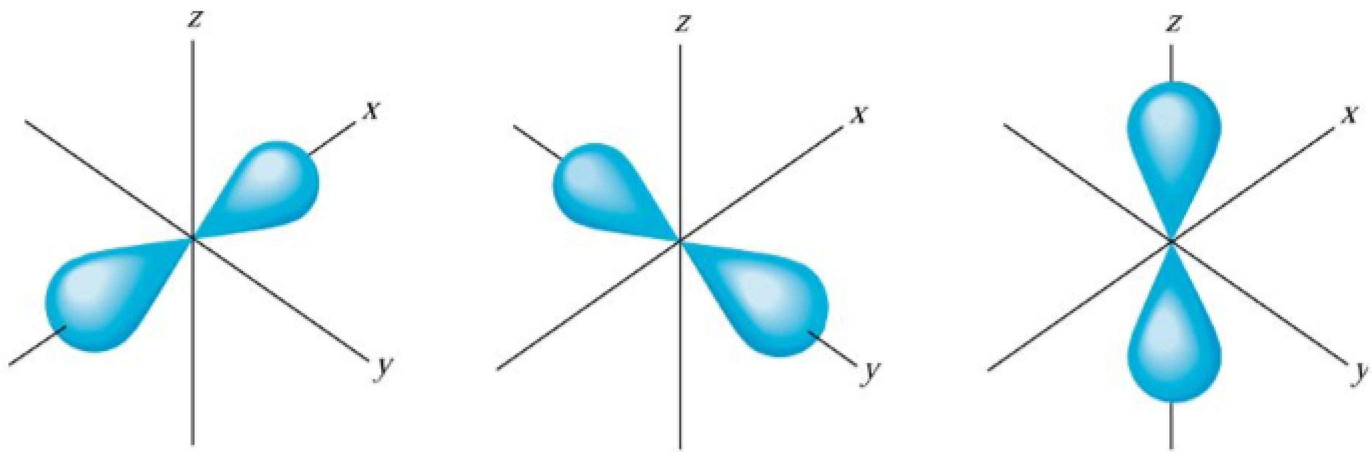
- Orbitals within the same subshell are usually of similar shape, but point in different directions.
- Electrons within orbitals also possess spin properties. Two electrons in the same orbital will spin in opposed directions.

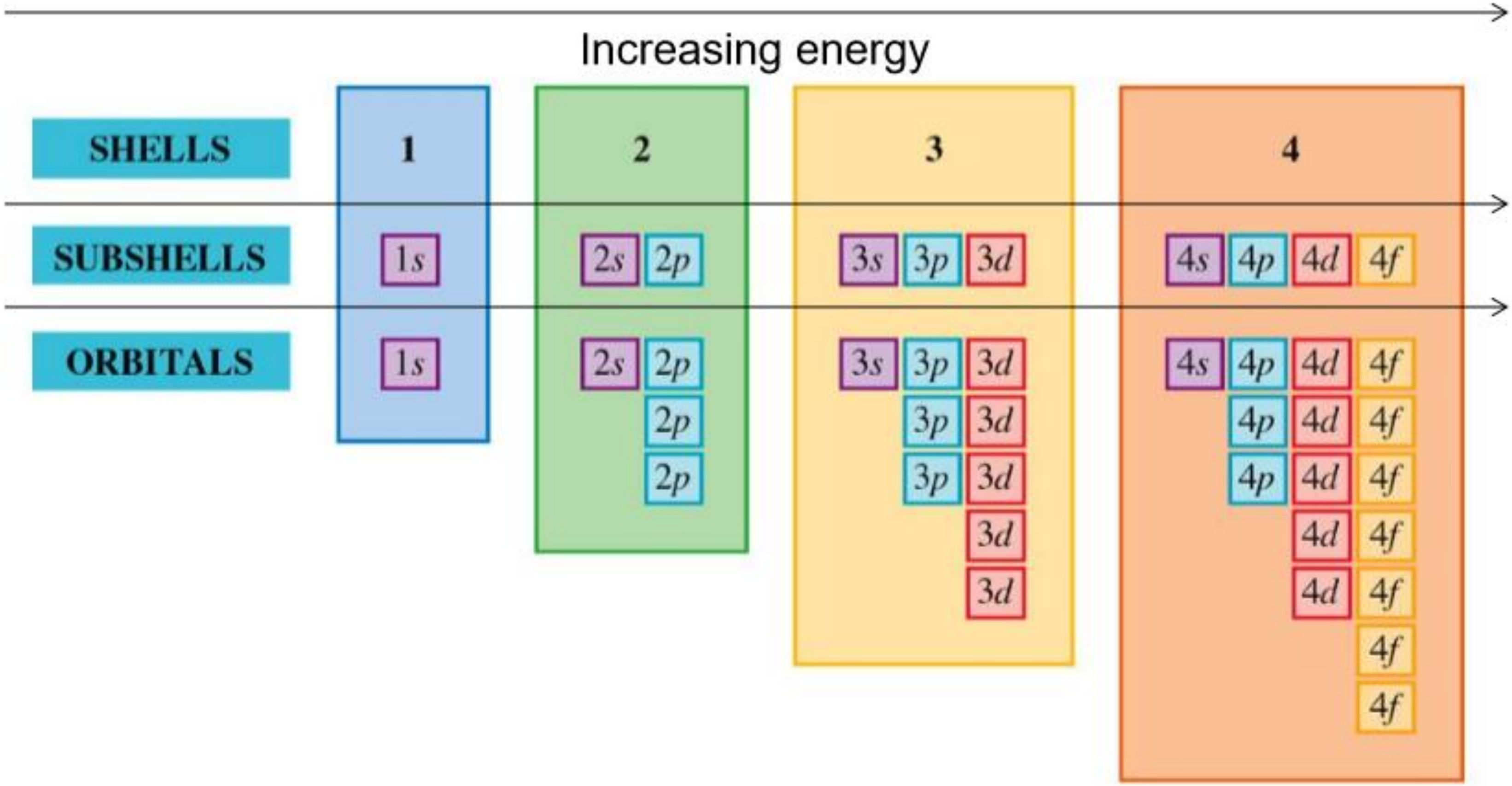
“spin-up” ↑

“pair” ↑↓

One electron is described as “spin-up” and the other “spin-down”

“spin-down” ↓





IMPORTANT NUMERICAL RELATIONSHIPS

- Subshells within a shell = shell number
- Orbitals within a subshell depends on shell type:
  1 for *s* 3 for *p* 5 for *d* 7 for *f*
- Electrons within an orbital = 2

Beginning with shell 5, not all subshells are needed to accommodate electrons. Those needed are

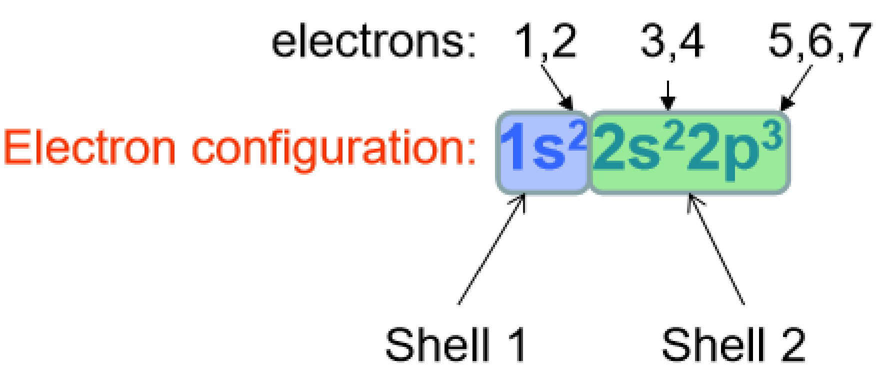
5*s*	5*p*	5*d*	5*f*	—		
6*s*	6*p*	6*d*	—	—	—	
7*s*	7*p*	—	—	—	—	—

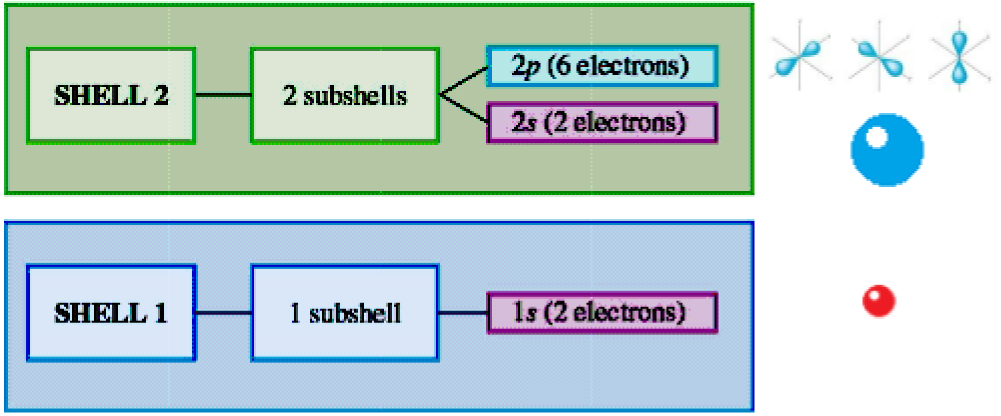
Electron configurations

- The arrangement of electrons within shells, subshells, and orbitals is governed by three rules:
  - Electron shells and subshells are filled in order of increasing energy
  - Orbitals can hold two electrons, at most, and these electrons must have opposed spins.
  - Electrons occupy orbitals of a subshell in a way that each orbital acquires one electron before any of them contains a pair of electrons. Also, these *singly occupied orbitals* must possess electrons which have the same spin.

Electron configurations

- An electron configuration is a statement of how many electrons each atom has in each of its subshells. The configuration lists an atom's electrons and their energies.
- The electron configuration for nitrogen (Z = 7) would show $2e^-$ in the first shell (1s) and the remaining $5e^-$ in the second shell. Because only $2e^-$ can fit into an s-subshell, the second shell would look have $2e^-$ in its 2s-subshell, and three in its 2p-subshell:



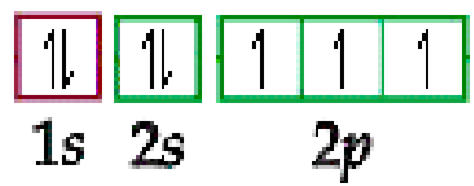


Orbital diagrams

- Orbital diagrams express information similar to electron configurations, but also express orbital occupation (how many electrons are in each of its occupied orbitals:

For nitrogen (Z = 7)

"Orbital diagram"

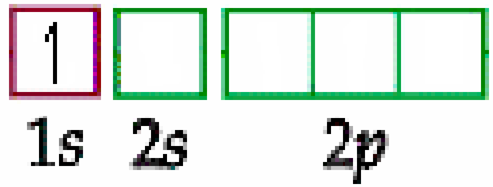


Electron configuration $1s^2 2s^2 2p^3$

Orbital diagrams and electron configurations for the first 10 elements (first, Z = 1 through 5)

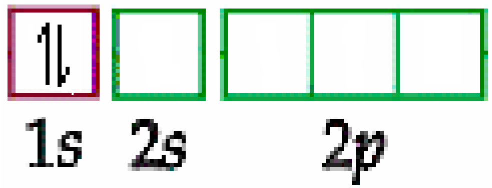
Orbital Diagram — **Electron Configuration**

H (Z = 1)



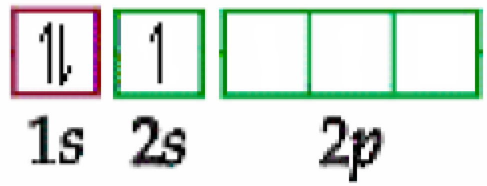
$1s^1$

He (Z = 2)



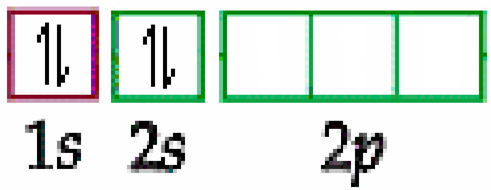
$1s^2$

Li (Z = 3)



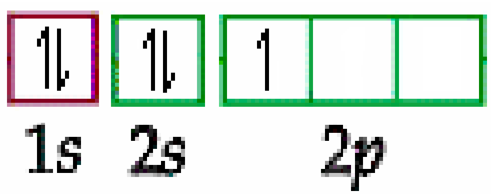
$1s^2 2s^1$

Be (Z = 4)



$1s^2 2s^2$

B (Z = 5)



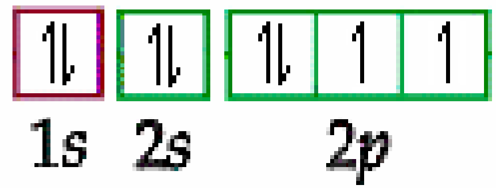
$1s^2 2s^2 2p^1$

Orbital diagrams and electron configurations for the first 10 elements (next, Z = 6 through 10)

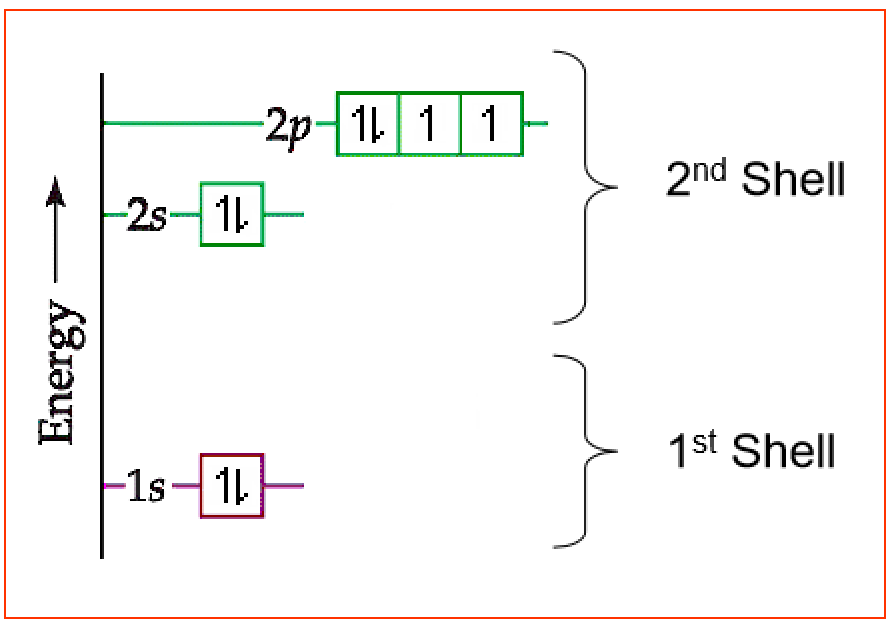
	Orbital Diagram					Electron Configuration
	1s	2s	2p			
C (Z = 6)	↑↓	↑↓	↑	↑		$1s^2 2s^2 2p^2$
N (Z = 7)	↑↓	↑↓	↑	↑	↑	$1s^2 2s^2 2p^3$
O (Z = 8)	↑↓	↑↓	↑↓	↑	↑	$1s^2 2s^2 2p^4$
F (Z = 9)	↑↓	↑↓	↑↓	↑↓	↑	$1s^2 2s^2 2p^5$
Ne (Z = 10)	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2 2s^2 2p^6$

Another way to draw orbital diagrams

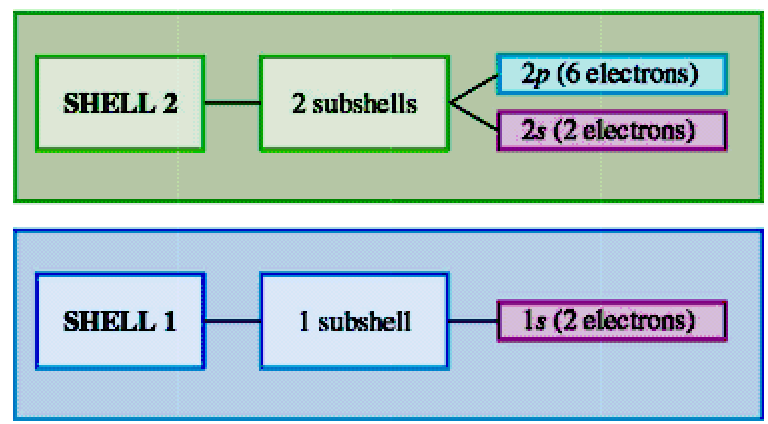
O (Z = 8)



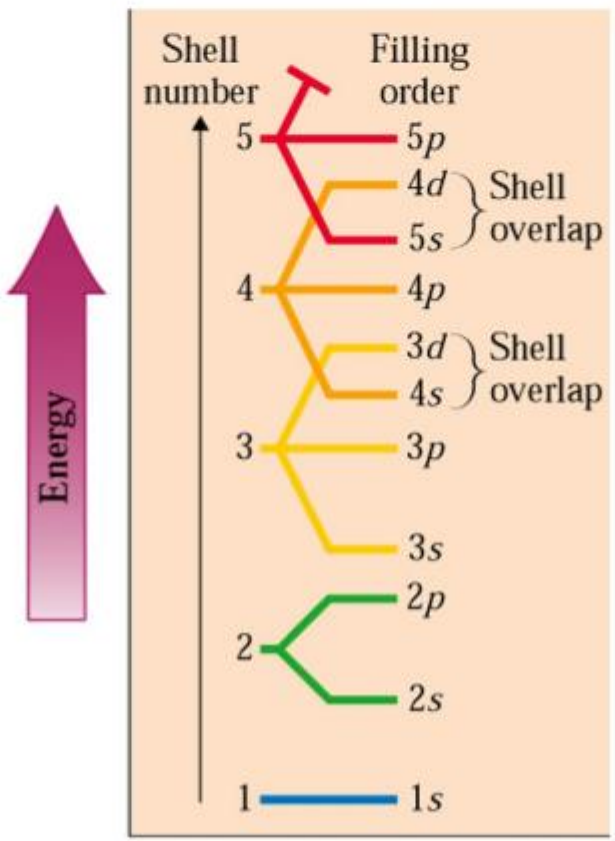
$1s^2 2s^2 2p^4$

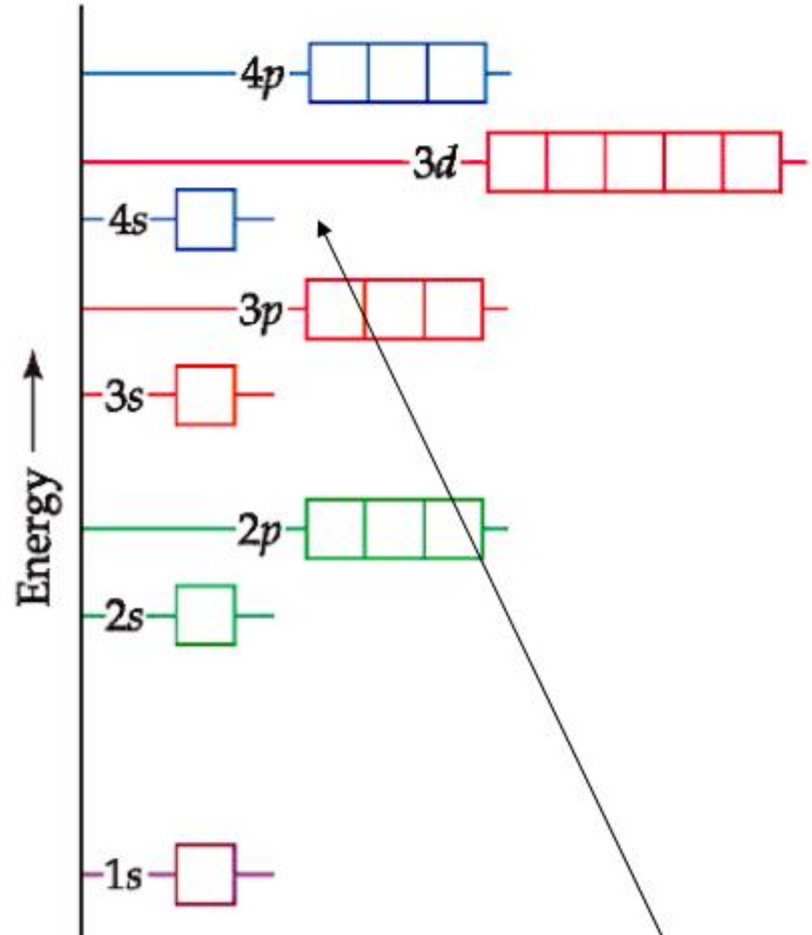


different way of drawing an orbital diagram



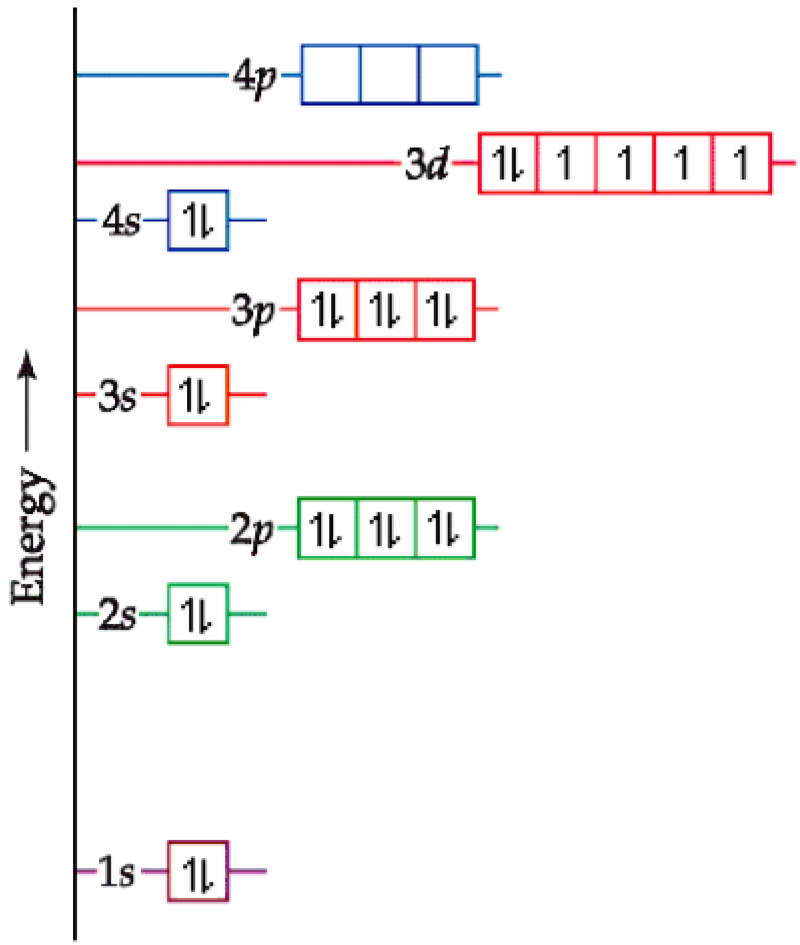
This is easy….no…wait!





4s orbital is lower in energy than 3d

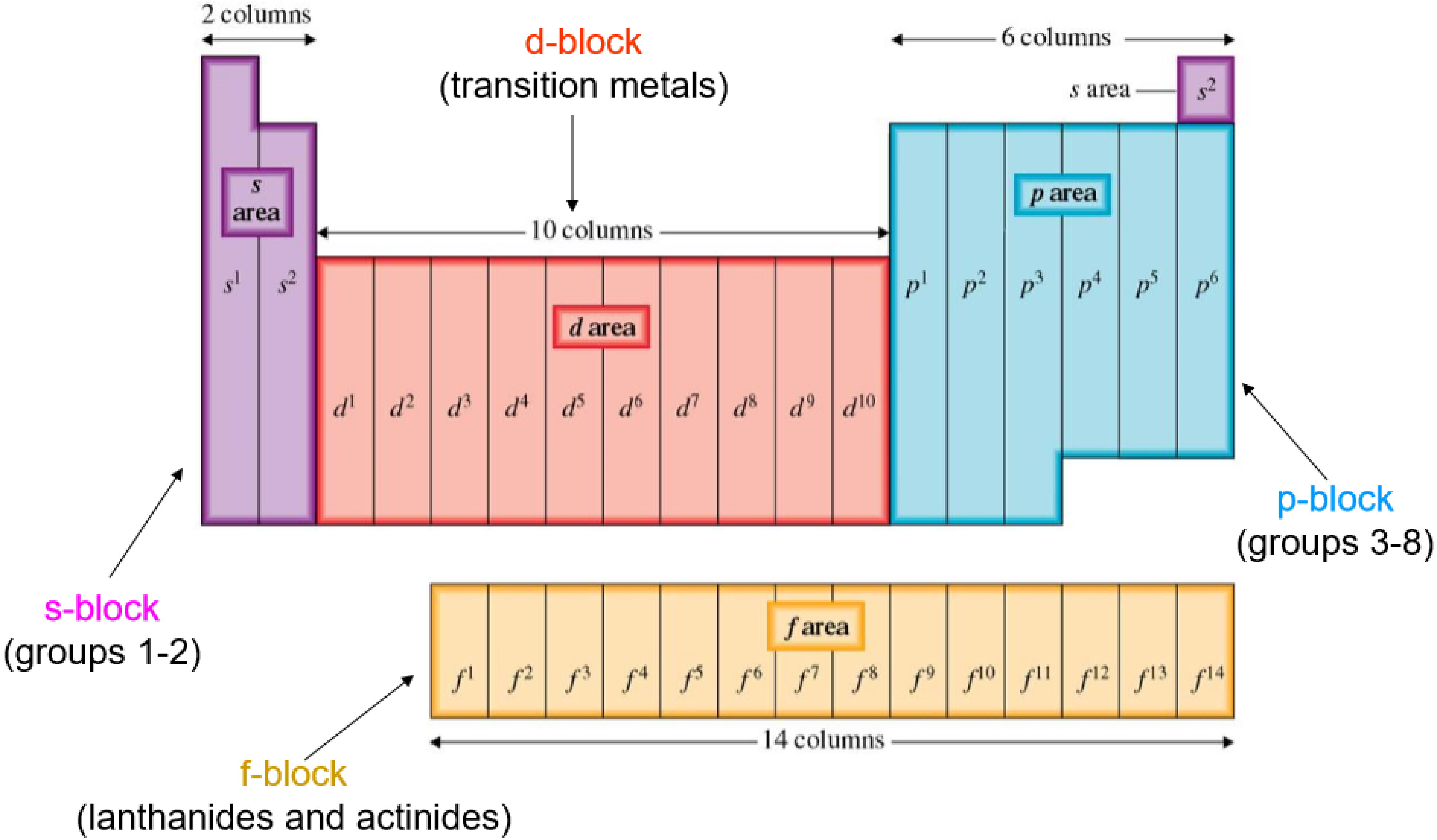
An example of a heavier element's configuration (Fe, Z = 26)

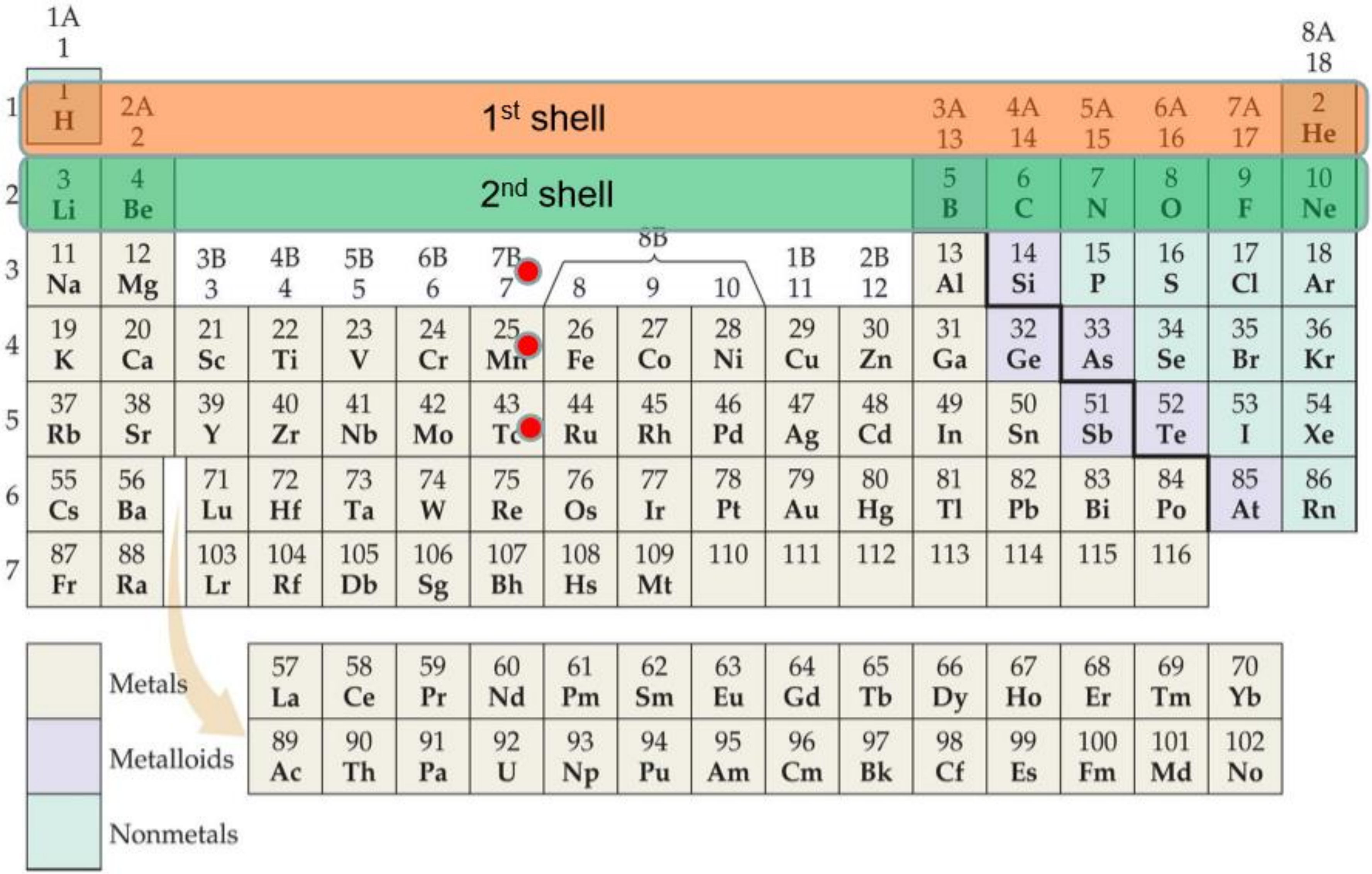


Electron configuration for Fe

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

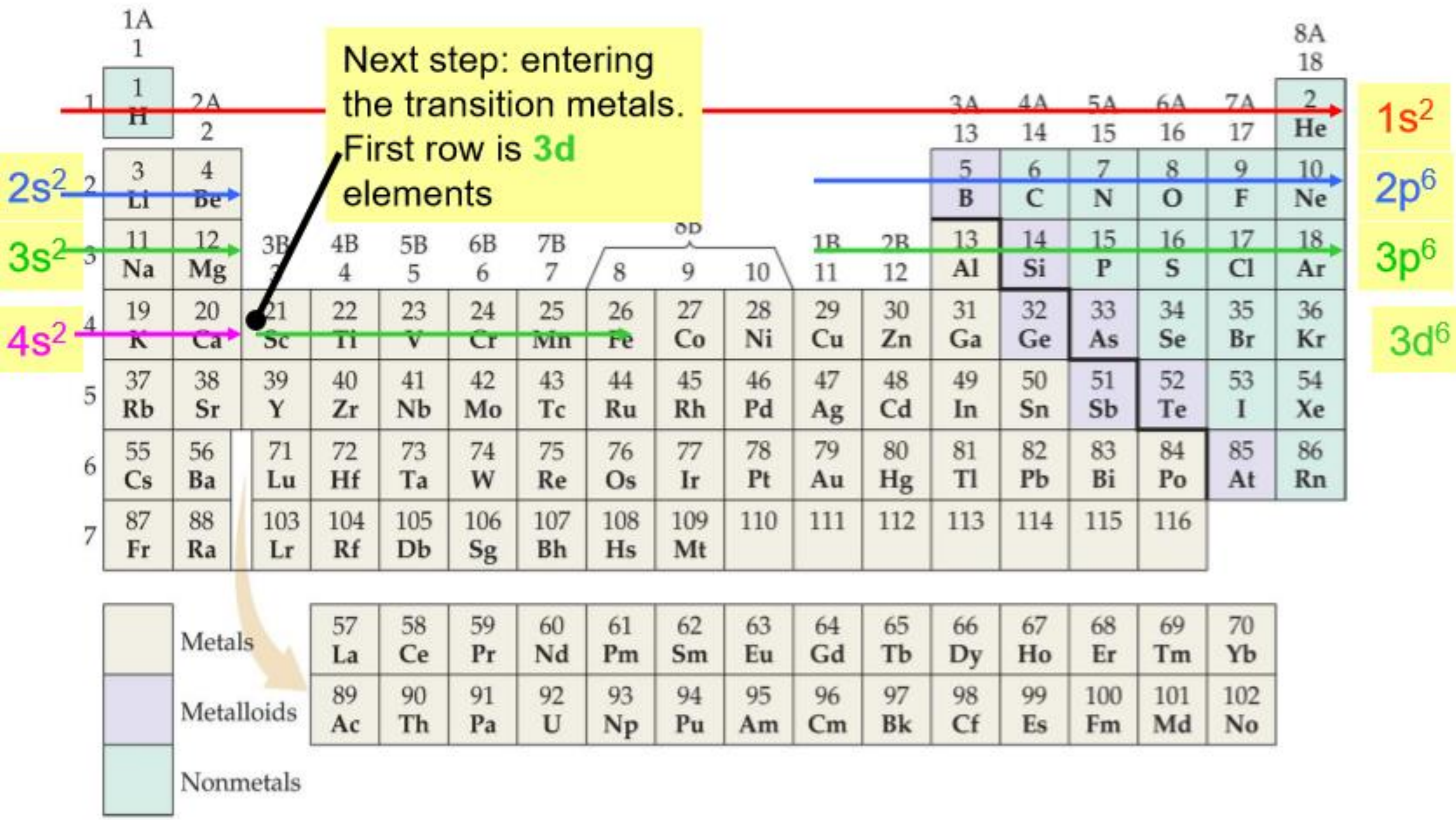
How do I remember the order of orbital filling when writing electron configurations?





1A 1
8A 18
1st shell
2nd shell
2A 2
3A 13
4A 14
5A 15
6A 16
7A 17
3B 3
4B 4
5B 5
6B 6
7B 7
8B 8 9 10
1B 11
2B 12
1 H
2 He
3 Li 4 Be 5 B 6 C 7 N 8 O 9 F 10 Ne
11 Na 12 Mg 13 Al 14 Si 15 P 16 S 17 Cl 18 Ar
19 K 20 Ca 21 Sc 22 Ti 23 V 24 Cr 25 Mn 26 Fe 27 Co 28 Ni 29 Cu 30 Zn 31 Ga 32 Ge 33 As 34 Se 35 Br 36 Kr
37 Rb 38 Sr 39 Y 40 Zr 41 Nb 42 Mo 43 Tc 44 Ru 45 Rh 46 Pd 47 Ag 48 Cd 49 In 50 Sn 51 Sb 52 Te 53 I 54 Xe
55 Cs 56 Ba 71 Lu 72 Hf 73 Ta 74 W 75 Re 76 Os 77 Ir 78 Pt 79 Au 80 Hg 81 Tl 82 Pb 83 Bi 84 Po 85 At 86 Rn
87 Fr 88 Ra 103 Lr 104 Rf 105 Db 106 Sg 107 Bh 108 Hs 109 Mt 110 111 112 113 114 115 116
57 La 58 Ce 59 Pr 60 Nd 61 Pm 62 Sm 63 Eu 64 Gd 65 Tb 66 Dy 67 Ho 68 Er 69 Tm 70 Yb
89 Ac 90 Th 91 Pa 92 U 93 Np 94 Pu 95 Am 96 Cm 97 Bk 98 Cf 99 Es 100 Fm 101 Md 102 No
Metals
Metalloids
Nonmetals

Use periodic table to determine electron configuration for Fe



Electron configuration: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^6$