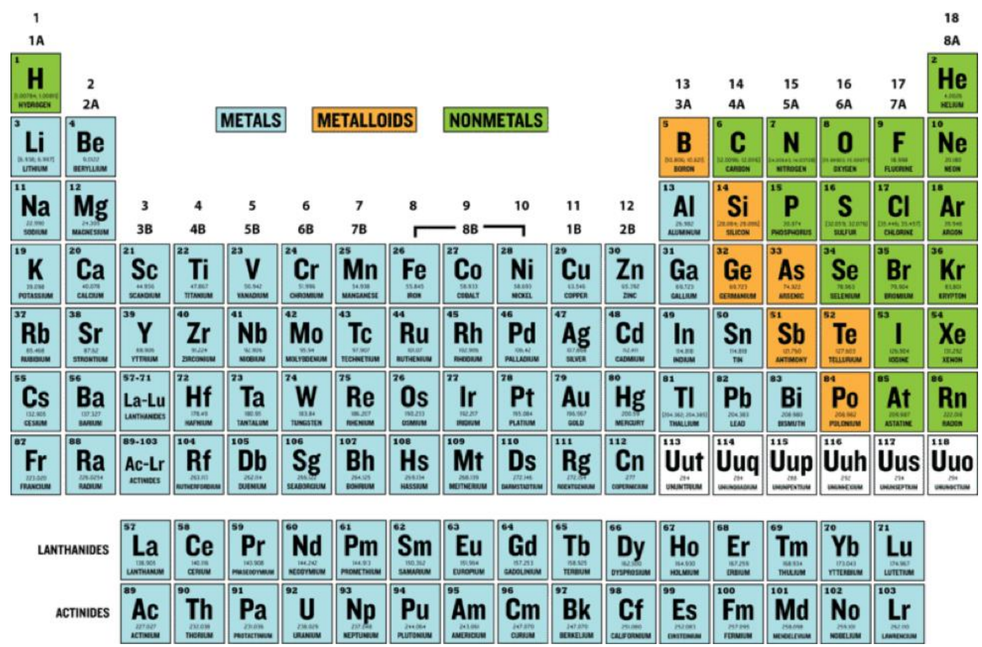
The Periodic Table and Periodic Law

Chapter 6

The Modern Periodic Table

- Know these locations on the Periodic Table
  - Alkali metals
  - Alkaline earth metals
  - Halogens
  - Representative elements
    - Groups 1, 2, 13-18
  - Transition elements
  - Metalloids
  - Noble gases
  - Inner-transition metals (names of each series)



The Modern Periodic Table

- Vertical columns are **groups** or **families**
- Horizontal rows are **periods**
- Know the locations where the sublevels are filled with electrons
- Short hand way of writing electron configuration
  - Start with the Noble-gas notation [noble gas] in the previous period and the electron configuration of the additional orbitals being filled
  - Example: Strontium is [Kr] $5s^2$

Valence Electrons

- Electrons in the highest principal energy level are called valence electrons
  - Be able to find the number of valence electrons for a group

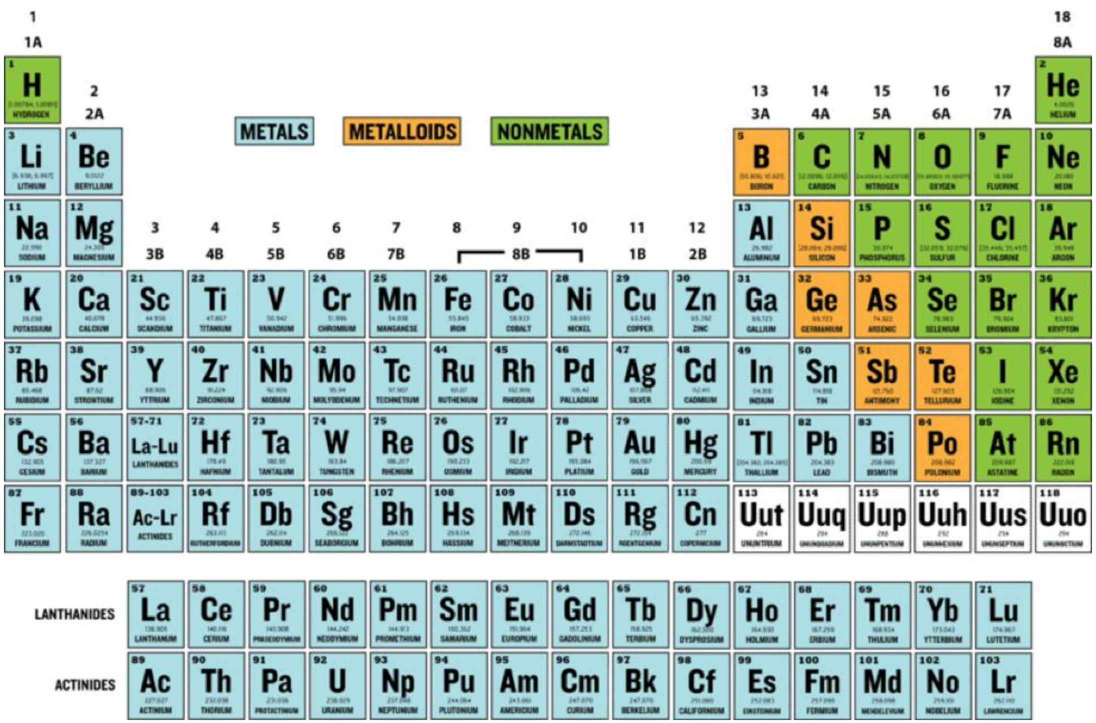


Table 6.3	Electron Configuration for the Group 1 Elements		
Period 1	hydrogen	$1s^1$	$1s^1$
Period 2	lithium	$1s^22s^1$	$[He]2s^1$
Period 3	sodium	$1s^22s^22p^63s^1$	$[Ne]3s^1$
Period 4	potassium	$1s^22s^22p^63s^23p^64s^1$	$[Ar]4s^1$

Periodic Trends

On Tests you have to EXPLAIN

- A trend is NOT an explanation
- Use the following in your explanations
  - Effective nuclear charge Zeff
  - Distance
  - Shielding
  - Minimize e/e repulsion
  - Electrons are negative and thus, attracted by the positive nucleus

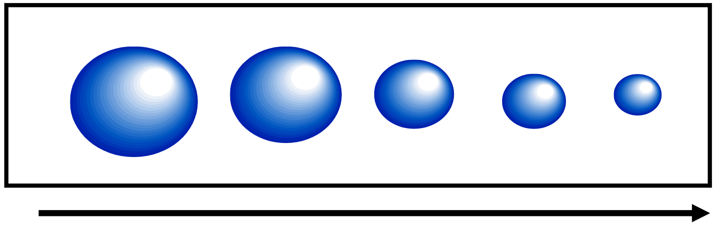
Periodic Trends - Terms

- **Valence electrons-** electrons in the highest principle energy level

- **Core electrons** – all other electrons

- **Effective nuclear charge (Zeff)**
  - the net charge experienced by a particular electron in an atom
  - depends on the number of electrons that **screen** the electron of interest
    - **Screening** – inner-core electrons act to offset the positive charge of the nucleus as seen by an electron further away; they screen a portion of the nuclear charge from valence electrons

Atomic Radius



- **Atomic radius** is a periodic trend influenced by electron configuration.
- In general as we go across a **period** from left to right, the atomic radius **decreases**.
  - No additional energy levels
  - The **nuclear charge (Z_{eff}) increases** across the period as protons increase. (more protons in the same energy level) therefore the valence electrons are drawn closer to the nucleus, decreasing the size of the atom

Atomic Radius

- Atomic radius generally **increases** as you move down a group.

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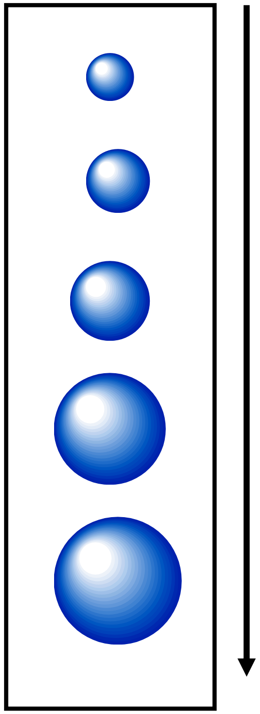
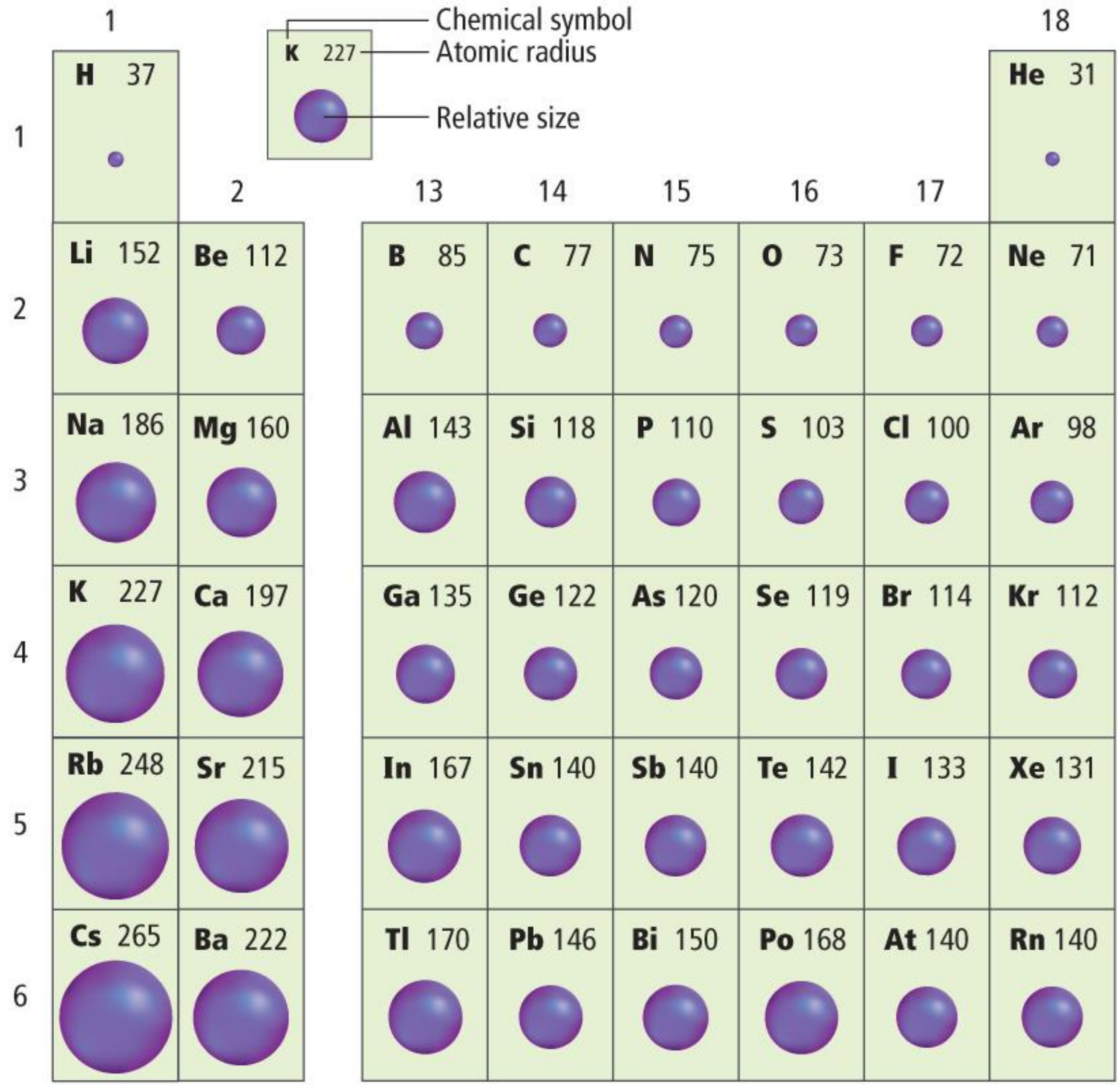
  - As we move down, energy levels are added to the atom, this pushes electrons further from the nucleus

  - The **attractive force of the nucleus** has less effect on outer electrons.
    - Due to keep adding shells




Chemical symbol
Atomic radius
Relative size
K 227
1
2
13
14
15
16
17
18
1
2
3
4
5
6
H 37
He 31
Li 152
Be 112
B 85
C 77
N 75
O 73
F 72
Ne 71
Na 186
Mg 160
Al 143
Si 118
P 110
S 103
Cl 100
Ar 98
K 227
Ca 197
Ga 135
Ge 122
As 120
Se 119
Br 114
Kr 112
Rb 248
Sr 215
In 167
Sn 140
Sb 140
Te 142
I 133
Xe 131
Cs 265
Ba 222
Tl 170
Pb 146
Bi 150
Po 168
At 140
Rn 140

Ionic Radius

- An **ion** is an atom or bonded group of atoms with a positive or negative charge

- When atoms are **negative ions** (gain electrons), they can become **larger**
  - Gaining electrons in the same energy level. Electrostatic repulsion in that level increases

- When atoms are **positively ions** (lose electrons), they **become smaller** for two reasons:
  - The loss of a valence electron can leave an empty outer orbital resulting in a small radius.
  - Electrostatic repulsion decreases allowing the electrons to be pulled closer to the radius.

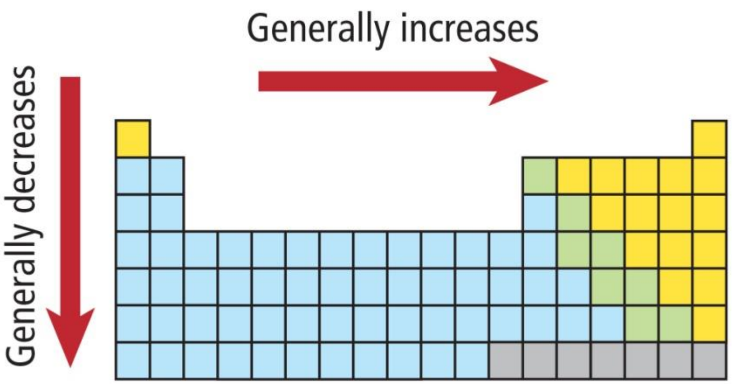
Ionization Energy

- **Ionization energy** is the energy required to remove an electron from a gaseous atom
  - Ionization energy indicates how strongly an atom has hold of its valance electrons
  - The energy required to remove the first electron is called the first ionization energy
  - Removing the second electron requires more energy, and is called the second ionization energy.
  - Each successive ionization requires more energy, but it is not a steady increase.

Ionization Energy

- First ionization energy increases from left to right across a period.
  - Reason: atomic radius smaller, electrons being held tighter (higher Zeff energy)

- First ionization energy decreases down a group
  - Reason: atomic size is increasing, so less energy is required to remove an electron farther from the nucleus

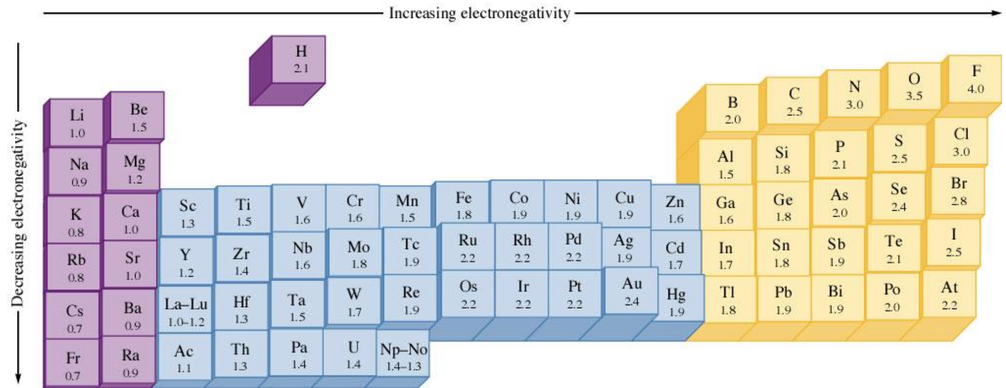


Trends in First Ionization Energies

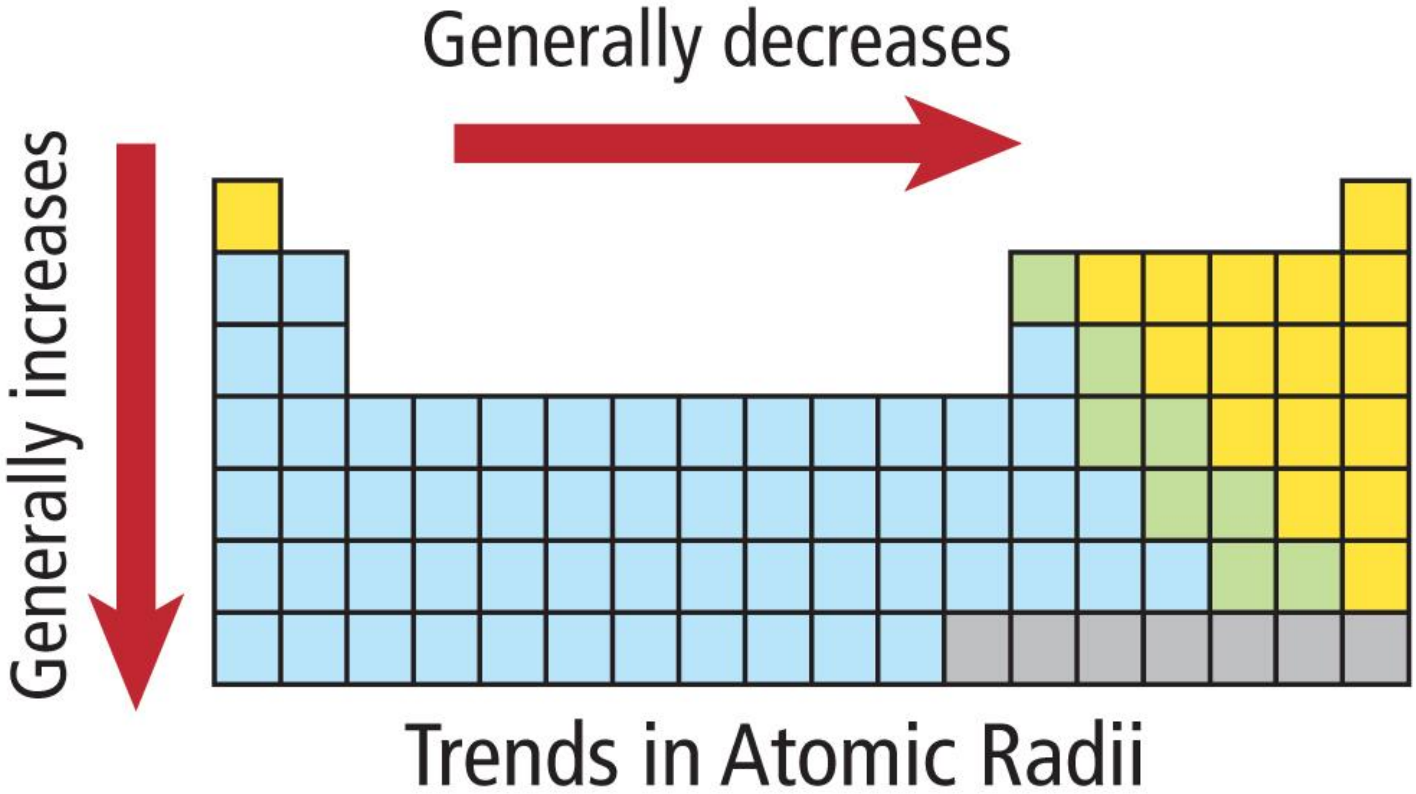
Electronegativity

- The **electronegativity (En)** : The ability of an atom IN A MOLECULE [meaning it’s participating in a BOND] to attract shared electrons to itself. Think “tug of war”.

- Electronegativity decreases down a group and increases left to right across a period
- Fluorine is the most En and Francium is the least En
- Why is F the most? Highest Zeff and smallest so that the nucleus is closest to the valence “action”.
- Why is Fr the least? Lowest Zeff and largest so that the nucleus is farthest from the“action”.



Review


Generally decreases
Generally increases
Trends in Atomic Radii

Key concepts

- The elements were first organized by increasing atomic mass, which led to inconsistencies. Later, they were organized by increasing atomic number
- The periodic law states that when the elements are arranged by increasing atomic number, there is a periodic repetition of their chemical and physical properties
- The periodic table organizes the elements into periods (rows) and groups (columns); elements with similar properties are in the same group
- Elements within a group have similar chemical properties.
- The group number for elements in groups 1 and 2 equals the element’s number of valence electrons..

Key Concepts

- The energy level of an atom’s valence electrons equals its period number.
- Atomic and ionic radii decrease from left to right across a period, and increase as you move down a group
- Ionization energies generally increase from left to right across a period, and decrease as you move down a group.
- The octet rule states that atoms gain, lose, or share electrons to acquire a full set of eight valence electrons.
- Electronegativity generally increases from left to right across a period, and decreases as you move down a group.

Key Concepts

- **Electronegativity** is a chemical property that says how well an atom can attract **electrons** towards itself. ... The **electron affinity** of an atom or molecule is defined as the amount of energy released when an **electron** is added to a neutral atom or molecule in the gaseous state to form a negative ion.