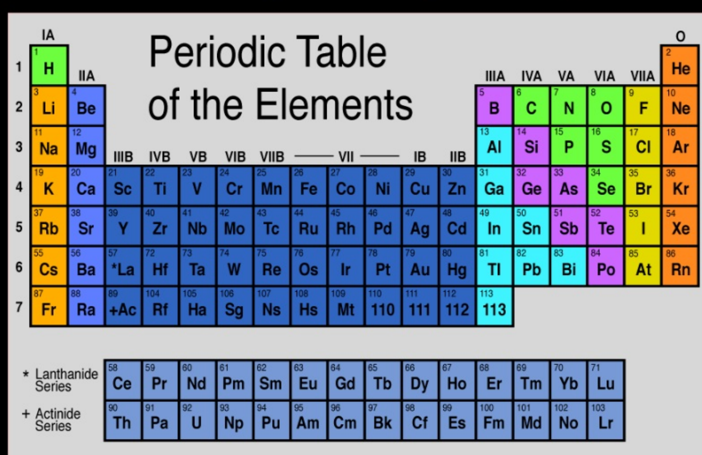
# Topics 3 & 13 - Periodicity

## How are the elements in the periodic table arranged? (3.1.1)



*Period - a horizontal row in the periodic table
*Group - a vertical column in the periodic table (3.1.2)

## Periodic Electron Configurations (3.1.3 & 3.1.4)

| 1 | 2 | 3 | 4 | 5 | 6 | 7 | 8 | 9 | 10 | 11 | 12 | 13 | 14 | 15 | 16 | 17 | 18 |
|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|---|
| | | | | | | | | | | | | | | | | | |

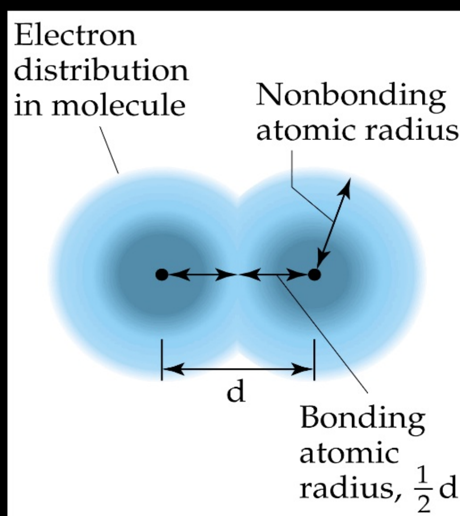
### Effective Nuclear Charge

- Effective nuclear charge is the charge experienced by an electron on a many-electron atom.
- The effective nuclear charge is not the same as the charge on the nucleus because of the effect of the inner electrons.
- Screening – the effect, due to electron-electron repulsion, by which an electron is repelled from the nucleus by other electrons that are closer to the nucleus.
- Electrons are attracted to the nucleus, but repelled by the electrons that screen it from the nuclear charge.

- The nuclear charge experienced by an electron depends on its distance from the nucleus and the number of core electrons.
- As the average number of screening electrons ($S$) increases, the effective nuclear charge ($Z_{eff}$) decreases.
- As the distance from the nucleus increases, $S$ increases and $Z_{eff}$ decreases.

# ATOMIC AND IONIC RADIUS

- Consider a simple diatomic molecule.
- The distance between the two nuclei is called the bond distance.
- If the two atoms which make up the molecule are the same, then half the bond distance is called the covalent atomic radius of the atom.



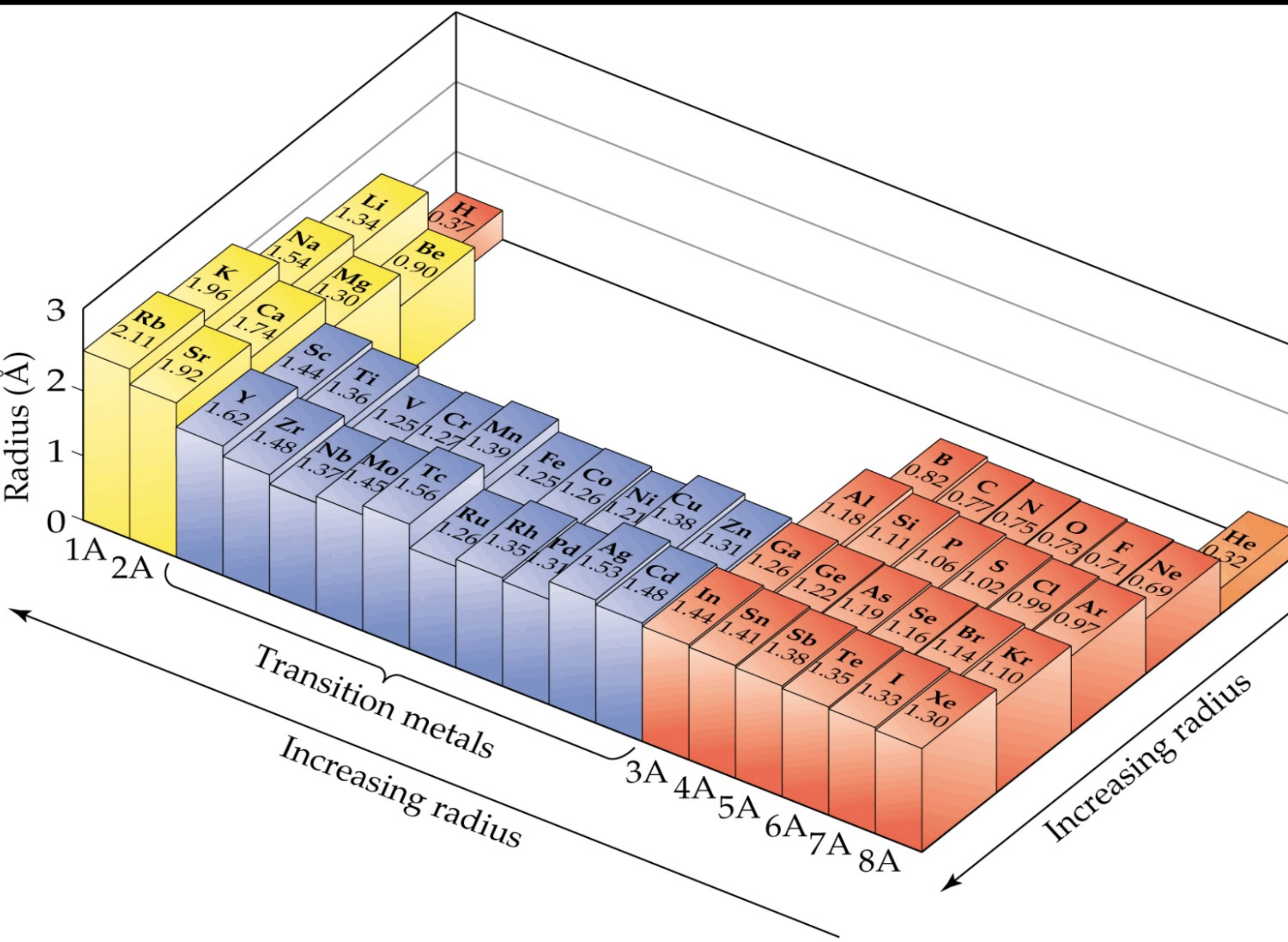
- As the principal quantum number increases, the size of the orbital increases.
- Consider the *s* orbitals.
- All *s* orbitals are spherical and increase in size as *n* increases.

## Periodic Trends in Atomic Radii (3.2.2 & 3.2.3)

- As a consequence of the ordering in the periodic table, properties of elements vary periodically.
- Atomic size varies consistently through the periodic table.
- As we move down a group, the atoms become larger.
- As we move across a period, atoms become smaller.

There are two factors at work:

- principal quantum number, *n*, and
- the effective nuclear charge, $Z_{eff}$.



Radius (Å)
3
2
1
0
1A
2A
3A
4A
5A
6A
7A
8A
H 0.37
Li 1.34
Be 0.90
Na 1.54
Mg 1.30
K 1.96
Ca 1.74
Rb 2.11
Sr 1.92
Sc 1.44
Ti 1.36
V 1.25
Cr 1.27
Mn 1.39
Fe 1.25
Co 1.26
Ni 1.21
Cu 1.38
Zn 1.31
Y 1.62
Zr 1.48
Nb 1.37
Mo 1.45
Tc 1.56
Ru 1.26
Rh 1.35
Pd 1.31
Ag 1.53
Cd 1.48
B 0.82
C 0.77
N 0.75
O 0.73
F 0.71
Ne 0.69
He 0.32
Al 1.18
Si 1.11
P 1.06
S 1.02
Cl 0.99
Ar 0.97
Ga 1.26
Ge 1.22
As 1.19
Se 1.16
Br 1.14
Kr 1.10
In 1.44
Sn 1.41
Sb 1.38
Te 1.35
I 1.33
Xe 1.30
Transition metals
Increasing radius
Increasing radius

**As the energy level increases (i.e., we move down a group), the distance of the outermost electron from the nucleus becomes larger. There is an increasing amount of screening on the outermost electrons. Hence, the atomic radius increases.**

**As we move across the periodic table, the number of core electrons remains constant. However, the nuclear charge increases. Therefore, there is an increased attraction between the nucleus and the outermost electrons. This attraction causes the atomic radius to decrease.**

Examples Place each group of elements in order of increasing atomic radius:

1. S, Al, Cl, Mg, Ar, Na

2. K, Li, Cs, Na, H

3. Ca, As, F, Rb, O, K, S, Ga

## Electron Configurations of Ions (3.2.3)

- **Cations**: electrons removed from orbital with highest principle quantum number, $n$, first:

$$\text{Li } (1s^2\ 2s^1)\quad \text{Li}^+\ (1s^2)$$

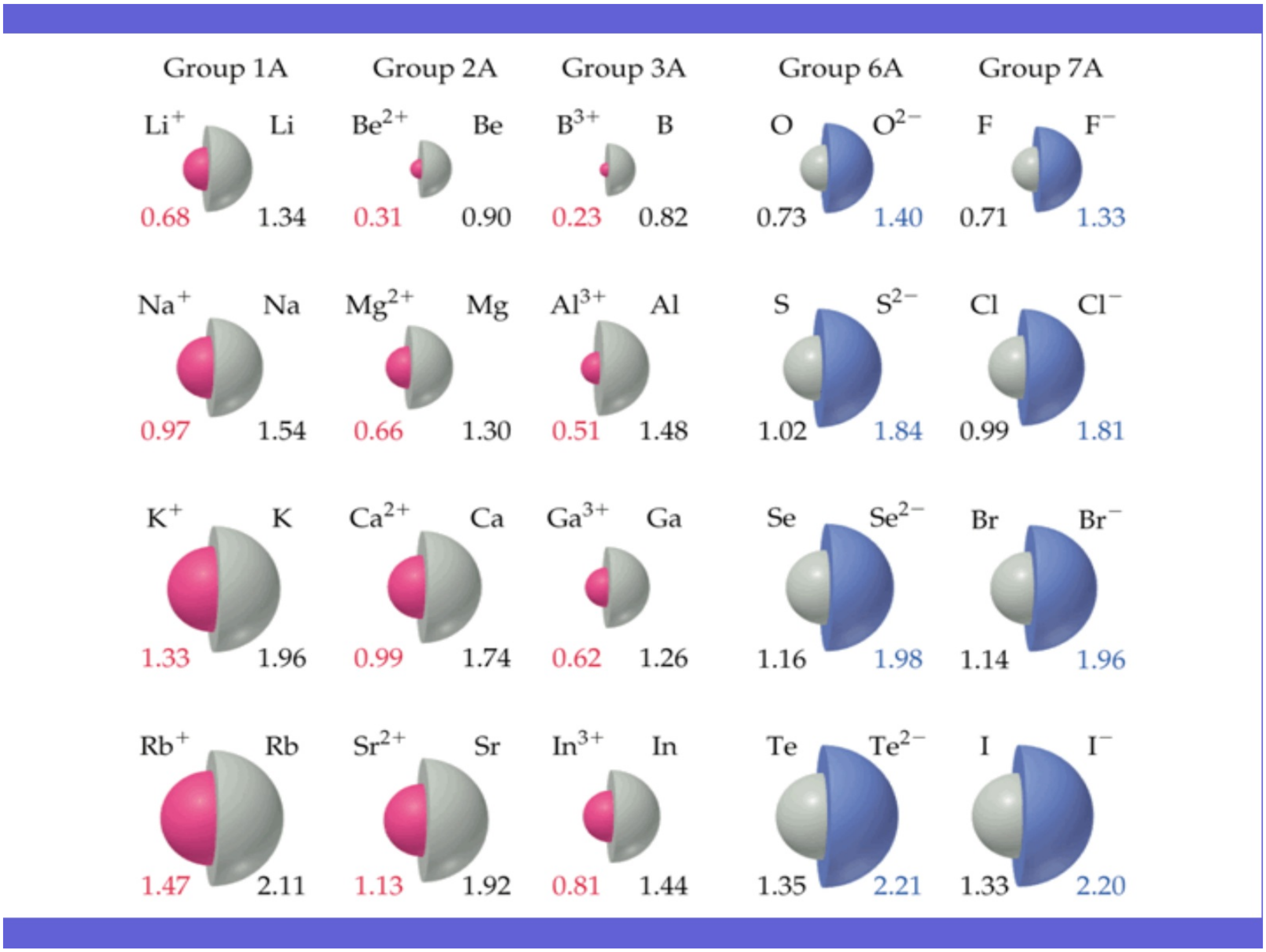
$$\text{Fe } ([\text{Ar}]3d^6\ 4s^2)\quad \text{Fe}^{3+}\ ([\text{Ar}]3d^5)$$

- **Anions**: electrons added to the orbital with highest $n$:

$$\text{F } (1s^2\ 2s^2\ 2p^5)\quad \text{F}^-\ (1s^2\ 2s^2\ 2p^6)$$

## Trends in the Sizes of Ions

•Ion size is the distance between ions in an ionic compound.
•Ion size also depends on nuclear charge, number of electrons, and orbitals that contain the valence electrons.
•Cations vacate the most spatially extended orbital and are **smaller than the parent atom**.
•Anions add electrons to the most spatially extended orbital and are **larger than the parent atom**.



Group 1A
Group 2A
Group 3A
Group 6A
Group 7A
$Li^+$ Li
0.68 1.34
$Be^{2+}$ Be
0.31 0.90
$B^{3+}$ B
0.23 0.82
O $O^{2-}$
0.73 1.40
F $F^-$
0.71 1.33
$Na^+$ Na
0.97 1.54
$Mg^{2+}$ Mg
0.66 1.30
$Al^{3+}$ Al
0.51 1.48
S $S^{2-}$
1.02 1.84
Cl $Cl^-$
0.99 1.81
$K^+$ K
1.33 1.96
$Ca^{2+}$ Ca
0.99 1.74
$Ga^{3+}$ Ga
0.62 1.26
Se $Se^{2-}$
1.16 1.98
Br $Br^-$
1.14 1.96
$Rb^+$ Rb
1.47 2.11
$Sr^{2+}$ Sr
1.13 1.92
$In^{3+}$ In
0.81 1.44
Te $Te^{2-}$
1.35 2.21
I $I^-$
1.33 2.20

- For ions of the same charge, ion size increases down a group.
- All the members of an **isoelectronic series** have the same number of electrons.
- As nuclear charge increases in an isoelectronic series the ions become smaller:

$$O^{2-} > F^{-} > Na^{+} > Mg^{2+} > Al^{3+}$$

Examples – Choose the larger species in each case:

1. $Na$ or $Na^+$
2. $Br$ or $Br^-$
3. $N$ or $N^{3-}$
4. $O^-$ or $O^{2-}$
5. $Mg^{2+}$ or $Sr^{2+}$
6. $Mg^{2+}$ or $O^{2-}$
7. $Fe^{2+}$ or $Fe^{3+}$