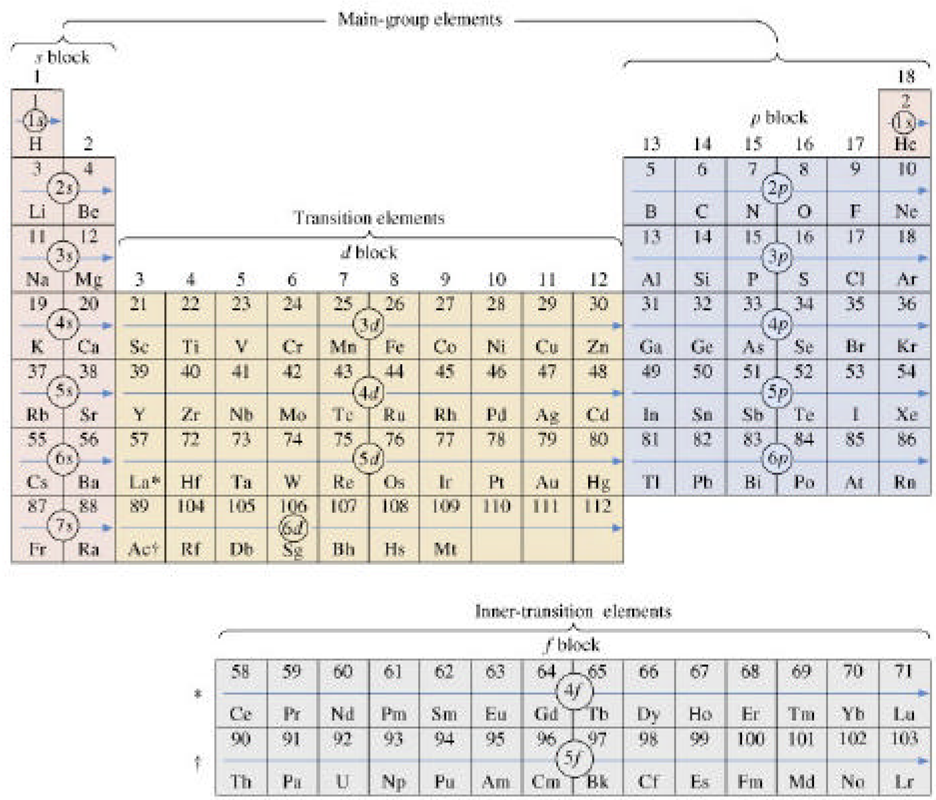
Electron Configurations continued:

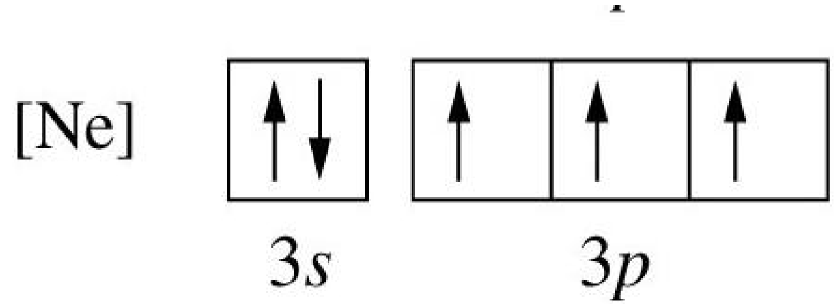
Electrons in the outermost shell are called ***valence electrons.***

- It is the ***valence electrons*** determine an atom's chemical properties.
- Electrons in the inner shells are inner electrons or ***core electrons***.
- Regions in periodic table are designated as the ***s-block***, ***p-block***, ***d-block*** and the ***f-block***.



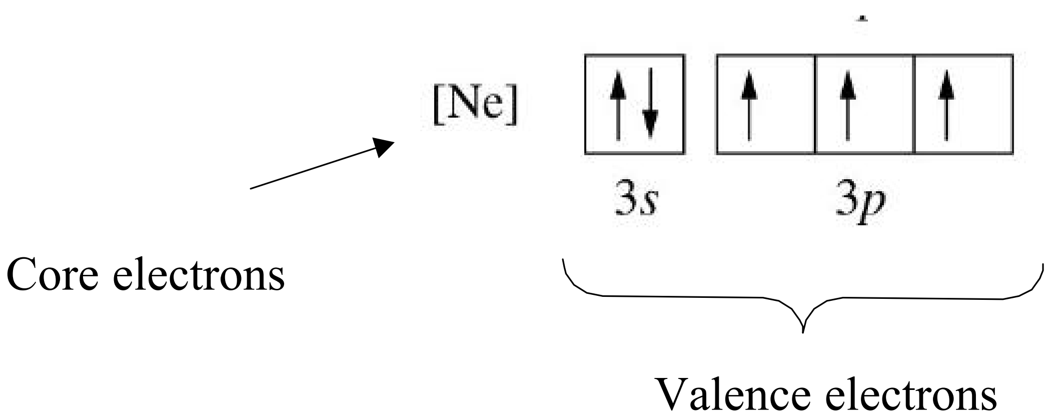
What is the ***electron configuration*** of the element as shown on the right?

Which electrons are the ***valence electrons*** and which are the ***core electrons***?



Answers:

$1s^2 2s^2 2p^6 3s^2 3p^6$ P   (***phosphorous***)



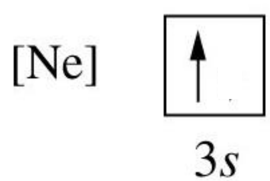
Electron Configurations of ions:

When an atom gains or loses electrons it becomes an ion.

Loss of electrons:	**Cation**	Gain of electrons:	**Anion**
	(+)		(-)

Atoms of group 1, 2 and the non-metals will lose or gain electrons to attain the electron configuration of a ***Nobel gas (ns^2np^6).***

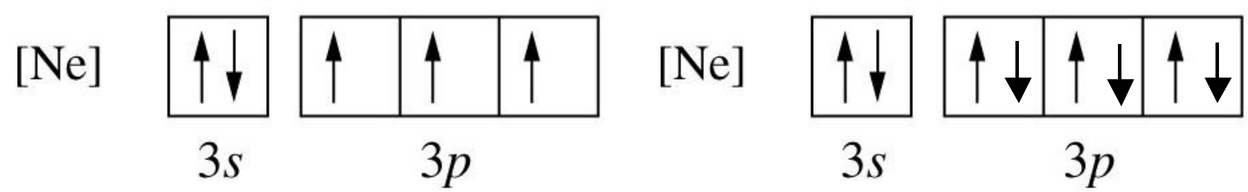
Example: **Sodium forming sodium ion:**

Na	**Na^+**
	[Ne]
$1s^22s^2sp^63s^1$	$1s^22s^2sp^6$

All group 1 cations (+1) will have the electron configuration of the previous Noble gas.

The formation of the phosphide ion:

P $(1s^2 2s^2 2p^6 3s^3 3p^3)$ P^{3-} $(1s^2 2s^2 2p^6 3s^3 3p^6) = [Ar]$



All group 5 (15) anions (-3) will have the electron configuration of the next Nobles gas.

Electron configurations of Transition Metal Ions:

The charge states on most transition metals cannot be equated to the electron configurations of the Noble gases. You will see why in chem. 2C (ch. 24 and 25)

The electron configurations of transitions metal cations must be determined from the charge state.

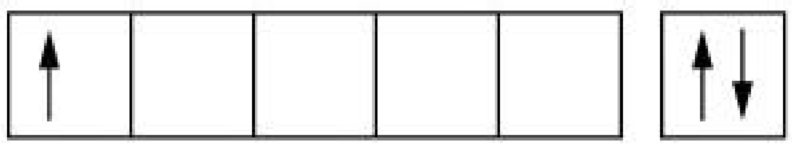
Fe^{3+}… 3 electrons lost Ni^{2+}… 2 electrons lost and so on.

- When ***transition metals*** ionize, they lose the s-electrons first.
- This may seem odd as it is opposite to the filling order.
- From a stability argument, the removal of the s-electrons first results in lower energy configuration when we take into account ionization energies of s vs. p and d electrons.

Examples:

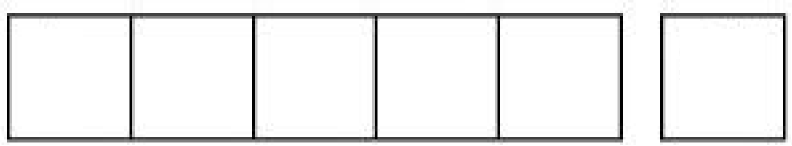
$3d^1$ $4s^2$

Sc: [Ar]

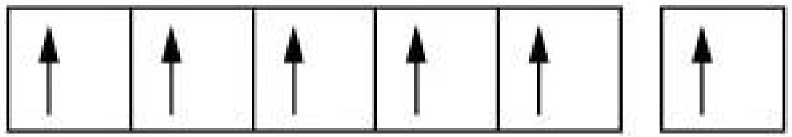


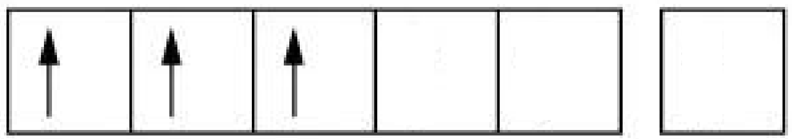
$[Ar]3d^1 4s^2$

Sc^{3+}: [Ar]



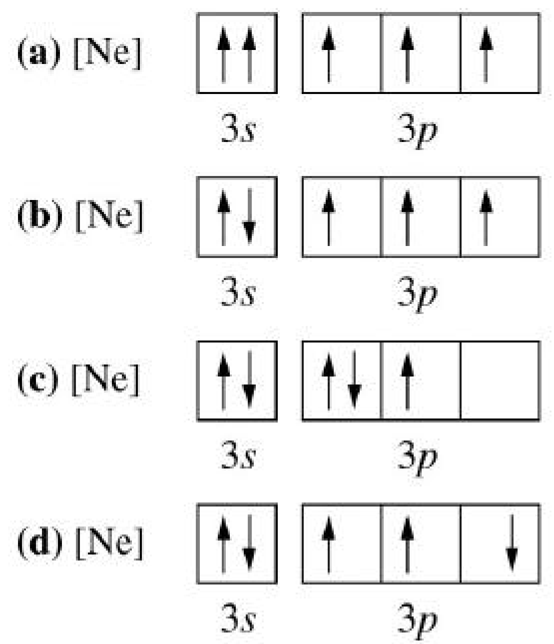
[Ar]

Cr: [Ar]  [Ar]$3d^5 4s^1$

Cr^{3+}: [Ar]  [Ar] $3d^3$

Electron Configurations: ***Excited States***

Which of the electron configurations to the right corresponds to phosphorous?



Answer: c

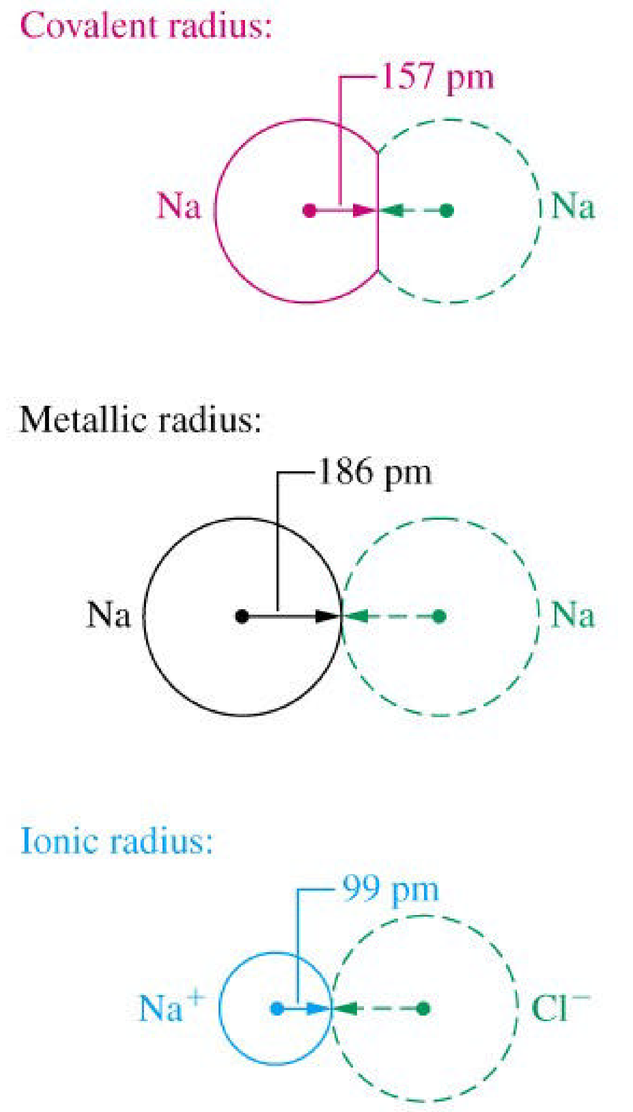
The rest are excited states of phosphorous.

11.

The Periodicity of Atomic Properties: Trends across the periodic table (Ch.10)

Half the distance between neighboring atoms defines an ***atomic radius***.

- In covalent compounds (molecules) the ***covalent radius*** is half the distance between the nuclei of two identical atoms held together by a single covalent bond.
- The ***metallic radius*** in a metal is half the distance between nuclei of two atoms in contact in a crystalline solid metal.
- An ***ionic radius*** is derived from the distance between nuclei of a cation and anion pair in an ionic bond.

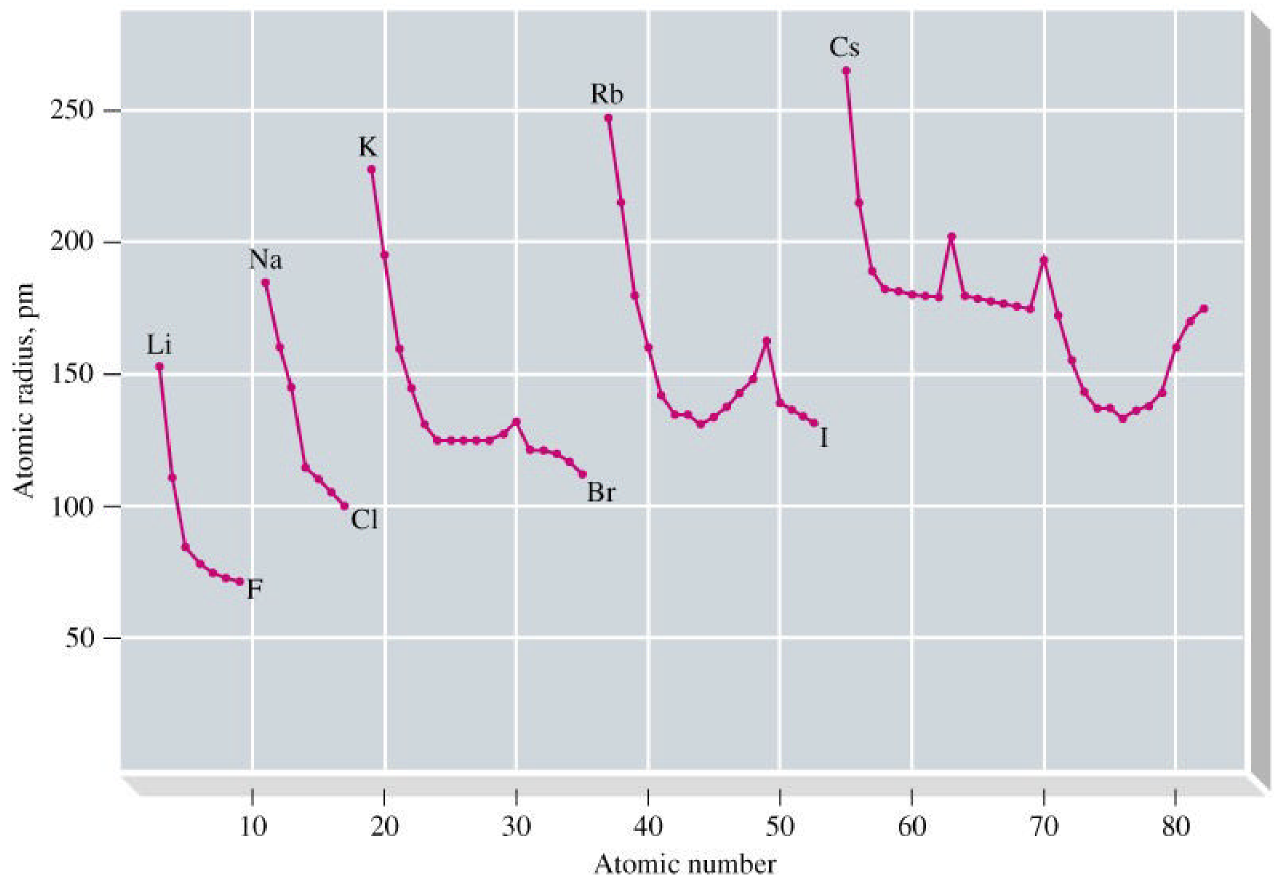


The Periodicity of Atomic Radii

The ***atomic radius*** of atoms decreases going from left-to-right across a period.

- Across a period, the effect of shielding decreases which allows the valence electron to be held tighter.

- Atomic radius increases down a group. Each step down involves adding a new shell, which increases size.



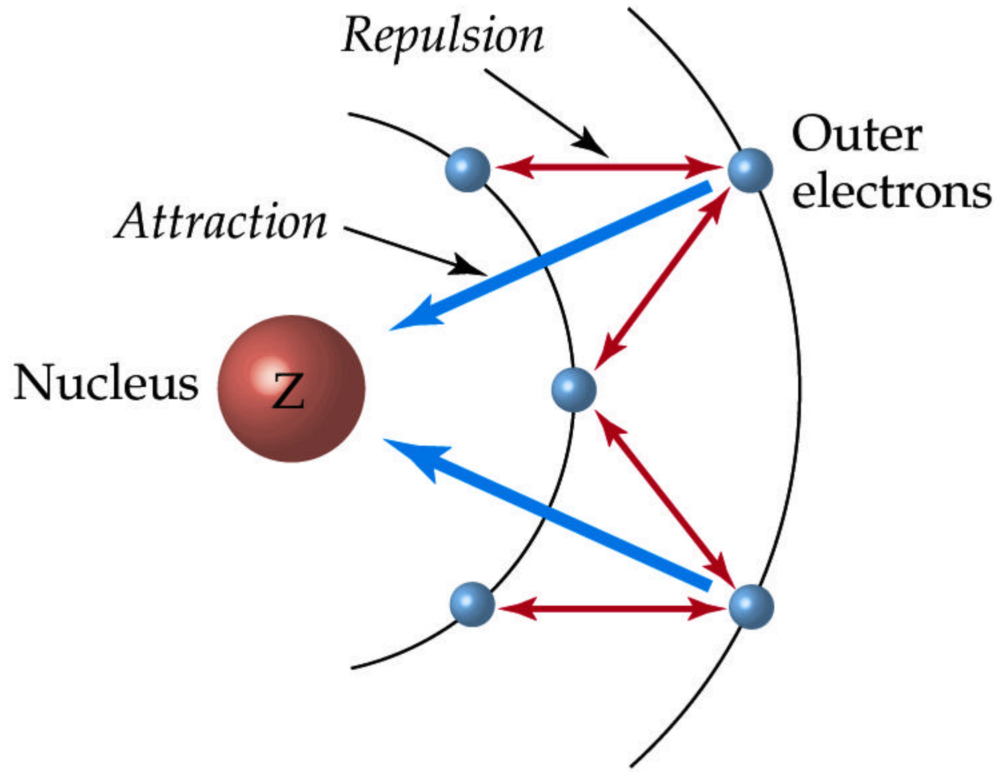
Effective Nuclear Charge:

The periodic trend of atomic radius is an example of the effect of shielding.

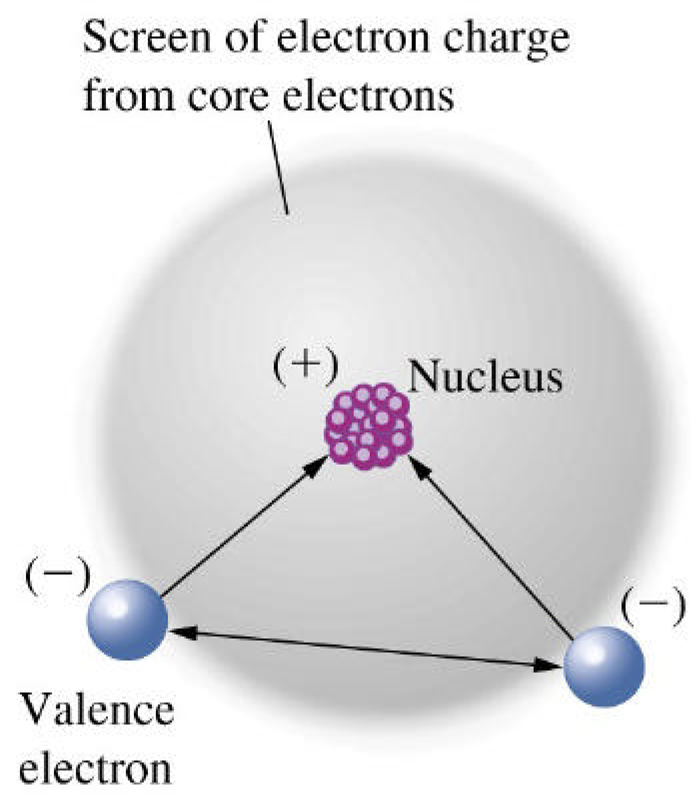
Shielding reduces the charge $\mathbf{Z}$ that the valence electrons feel to $\mathbf{Z_{eff}}$, which is called the ***effective nuclear charge***.

The innermost electrons (***core***) screen or "***shield***" the nucleus from the view of the outer (***valence***) electrons.

The valence electrons in effect see a charge (Z_{eff}) that is less that the actual charge on the nucleus, Z.



The interaction of the electrons reduces the attractive forces between the nucleus and the electrons.
The extent of this effect is a function of the screening ability of the core electrons:



$$Z_{eff} = Z - S$$

Z = atomic number

S = screening constant

The trend is:

s-electrons > d-electrons > p-electrons

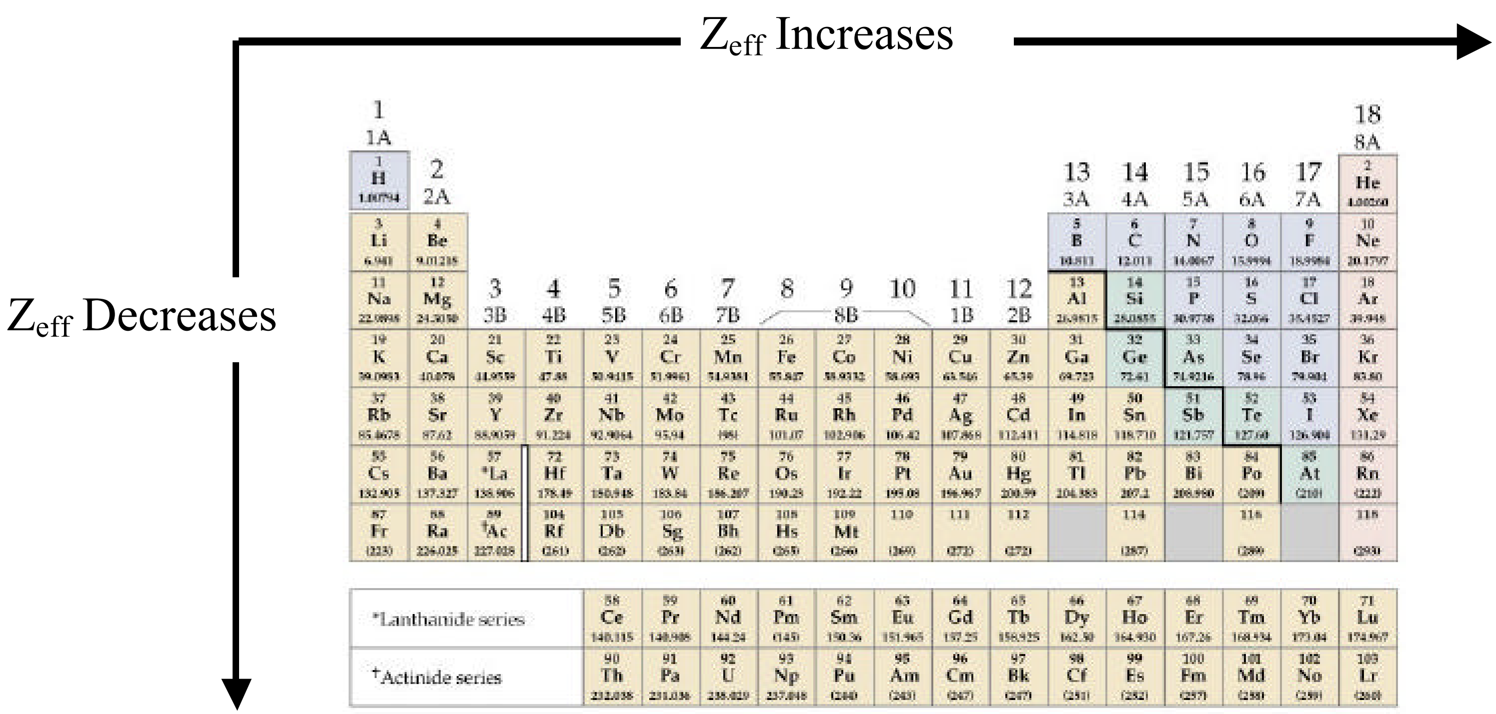
$\longleftarrow$

Increasing Screening ability

This indicates that Z_{eff} increases from left-to-right across a period.

As one moves down a group, Z_{eff} decreases as screening increases.

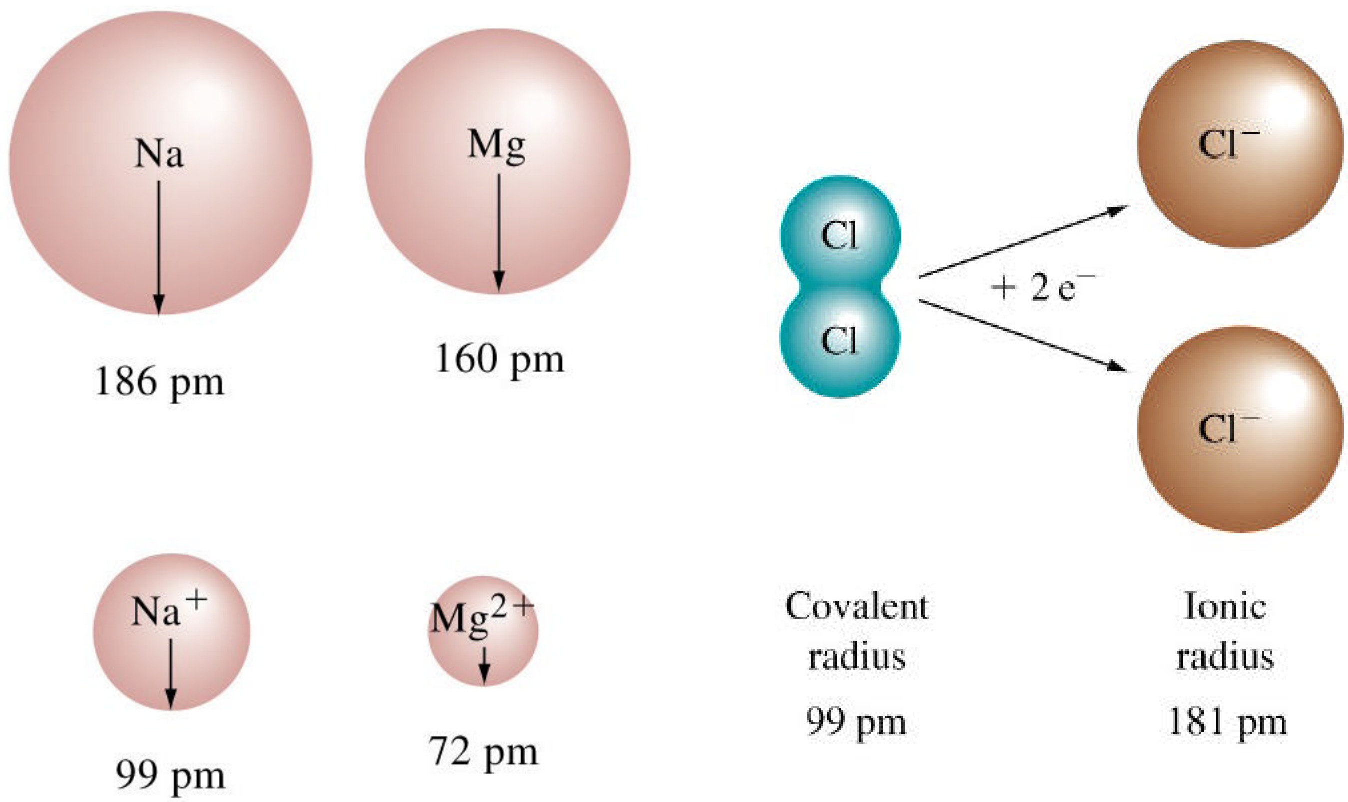
The over all trend is:



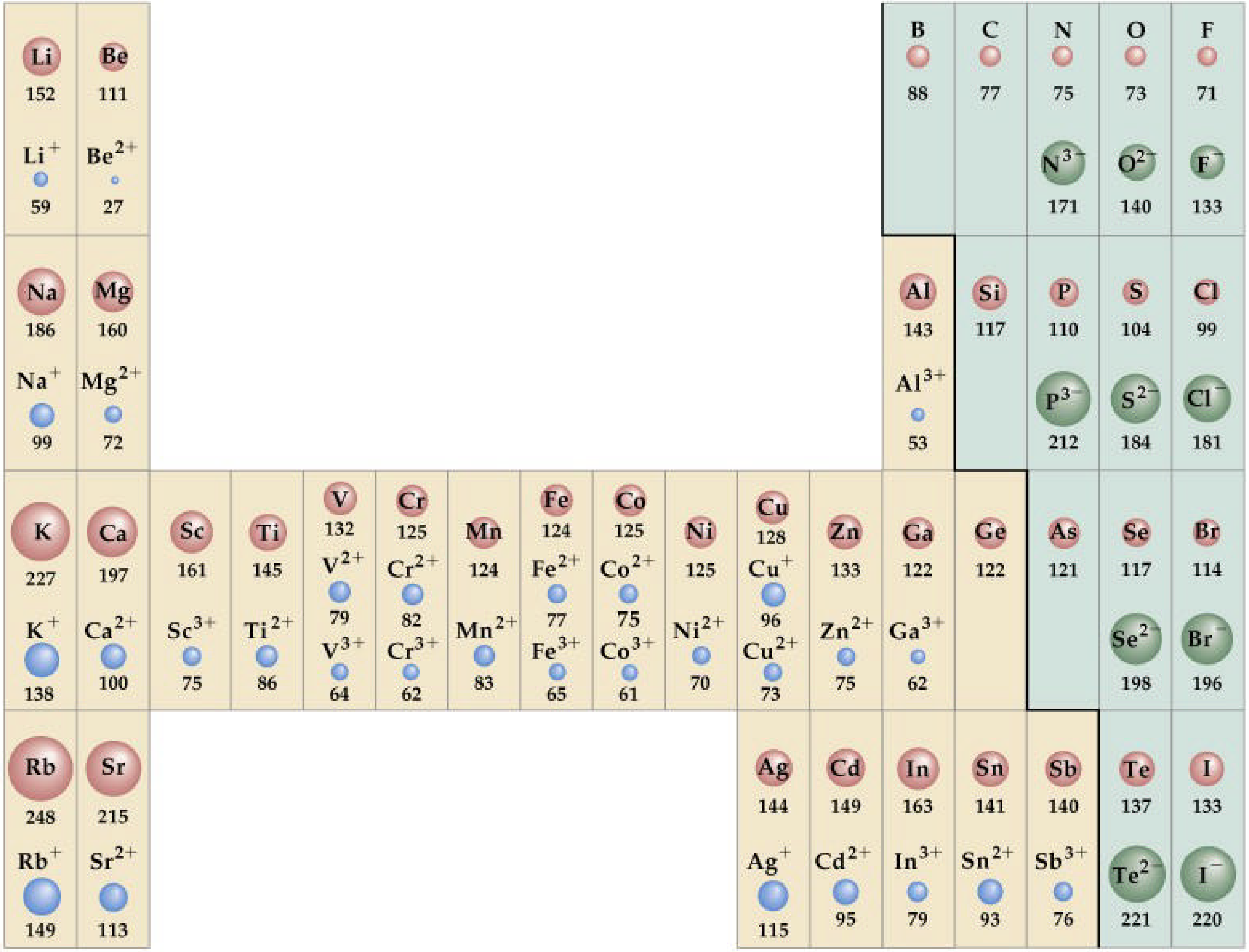
This explains the trend in atomic radius: Across, as Z_{eff} gets larger, the valence electrons are held more tightly. Down a group, the reverse is seen.

Trends in Ionic Radii

- The radii of cations are smaller than radii of the corresponding neutral atoms. (same Z, less electrons)
- The radii of anions are larger than radii of the corresponding neutral atoms. (same Z, more electrons)



The trends of ionic radii are a function of the charge state of the ion.



Rank the following ions in order of decreasing size?

Na^{+}, N^{3-}, Mg^{2+}, F^{-} and O^{2-}

ion	# protons	# of electrons	ne/np
Na^{+}	11	10	0.909
N^{3-}	7	10	1.43
Mg^{2+}	12	10	0.833
F^{-}	9	10	1.11
O^{2-}	8	10	1.25

- Notice that they all have 10 electrons: They are ***isoelectronic*** (same electron configuration) as Ne.
- Since N^{3-} has the highest ratio of electrons to protons, it must be the largest.
- Mg^{2+} must then be the smallest.

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