Chemical symbols

- Know names and symbols of elements #1 – 30, **plus**

Rb, Cs, Sr, Ba, Ag, Au, Cd, Hg, Pt, Ga, Ge, As, Sn, Pb, Se, Br, I, and U

Coulomb's Law

$$F = k\ \frac{Q_1 Q_2}{r^2}$$

F = attractive/repulsive force

Q_{1}, Q_2 = charges (size)

r = distance between charges

$$F = k\,\frac{Q_1 Q_2}{r^2}$$

- Electrons in which occupied energy level should be held most tightly by the nucleus? Most loosely?
- *“n=1” electrons should be held most tightly*
- ***Valence Electrons should be held most loosely***

$$F = k\,\frac{Q_1 Q_2}{r^2}$$

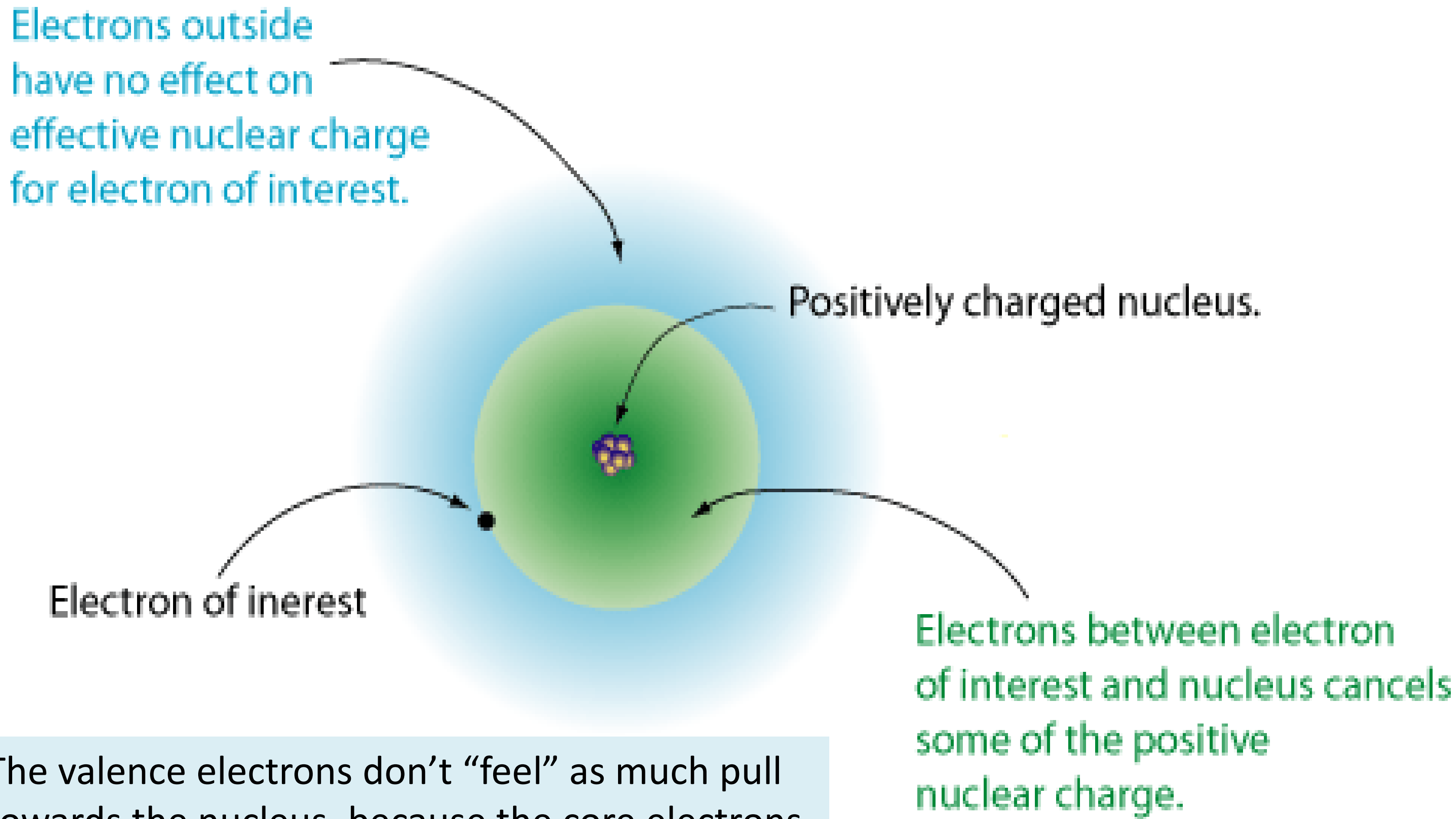
- Electrons in which orbital within an energy level should be held most tightly by the nucleus: s? p? d? Most loosely?
- *“s” electrons should be held most tightly*
- ***Electrons in the last orbital WITHIN AN ENERGY LEVEL being filled should be held most loosely***

Some

“periodic trends”

1) Effective nuclear charge (Z_{eff})

- Not all electrons in an atom can "feel" or experience the entire positive charge of the protons in the nucleus
- Electrons that are in between the outer electrons and the nucleus *"shield"* the outer electrons from feeling the entire positive nuclear charge
- Electrons in the same orbital set cannot "shield" each other



The valence electrons don’t “feel” as much pull towards the nucleus, because the core electrons “shield” them… but valence electrons cannot shield each other.

From wikipedia

Effective nuclear charge

- How to **estimate** the effective nuclear charge (Z_{eff}) on the outermost electrons in an atom:

Atomic number (Z) – electrons in filled orbitals below the orbital being filled

("shielding electrons")

Effective nuclear charge

- Example: what is the estimated Z_{eff} felt by the 3p electrons in **Al**?
- Al has a single 3p electron
- All 12 of the other electrons (1s thru 3s) shield this electron
- 13 – 12 = **+1**

Effective nuclear charge

- Example: what is the estimated Z_{eff} felt by the 3p electrons in **P**?
- P has a three 3p electrons
- All 12 of the other electrons (1s thru 3s) shield this electron
- 15 – 12 = **+3**
- **The 3p electrons in P are held more tightly than in Al!**

Effective nuclear charge

- Example: what is the estimated Z_{eff} felt by the 3p electrons in Cl?
- Cl has a five 3p electrons
- All 12 of the other electrons (1s thru 3s) shield this electron
- 17 – 12 = **+5**
- **The 3p e-'s in Cl are held more tightly than in P or in Al!**
- **Cl would attract additional e-'s more strongly as well!**

- Trend: effective nuclear charge ***increases*** moving ***left to right*** **within an orbital set** due to the shielding effect of core electrons

- Coulomb's Law – because the positive charge felt by the electron is larger (Q), the attractive force is also larger

Effective nuclear charge

- How about up and down?
- Biggest difference: moving down **increases the energy level** occupied by the outer electrons
- The outermost electrons are **further away** from the nucleus

- Example: how about Na vs K vs Rb?
- ***Estimated*** effective nuclear charge:

  Na: 11 – 10 = +1

  K: 19 – 18 = +1

  Rb: 37 – 36 = +1
- BUT: Rb's $5s^1$ valence electron is held more loosely than K's $4s^1$, which is held more loosely than Na's $3s^1$
- WHY?
- The valence electron is further away from the nucleus!

- Trend: the nuclear charge felt by the outer electrons ***decreases*** moving ***down a column*** on the periodic table due to the <u>increasing distance between the electrons and the nucleus</u>

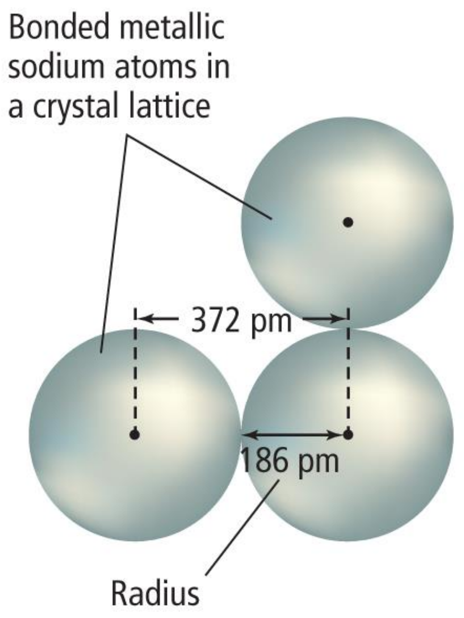
- Coulomb's Law – because the distance between the opposite charges is larger (r), the attractive force is smaller

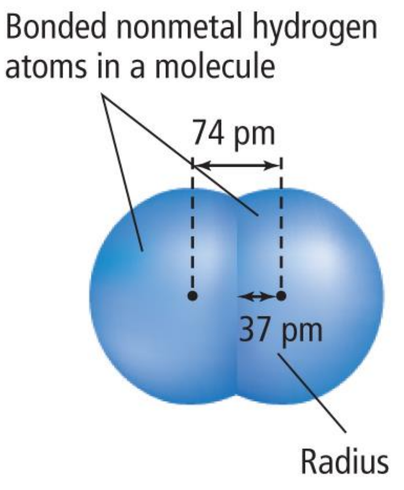
Do atoms have edges?

- No!
- Which orbitals does an atom “have”?
- **ALL ATOMS “HAVE” EVERY ORBITAL**
- There is a **BIG** difference between *“HAVING”* and orbital and having an orbital **OCCUPIED**
- **Ex:** every carbon atom “has” every orbital, but only the 1s, 2s and 2p are **occupied**

Do atoms have edges?

- No!
- Then… how do measure the size of an atom?
- We measure the distance between two identical bonded atoms, and “split the difference”

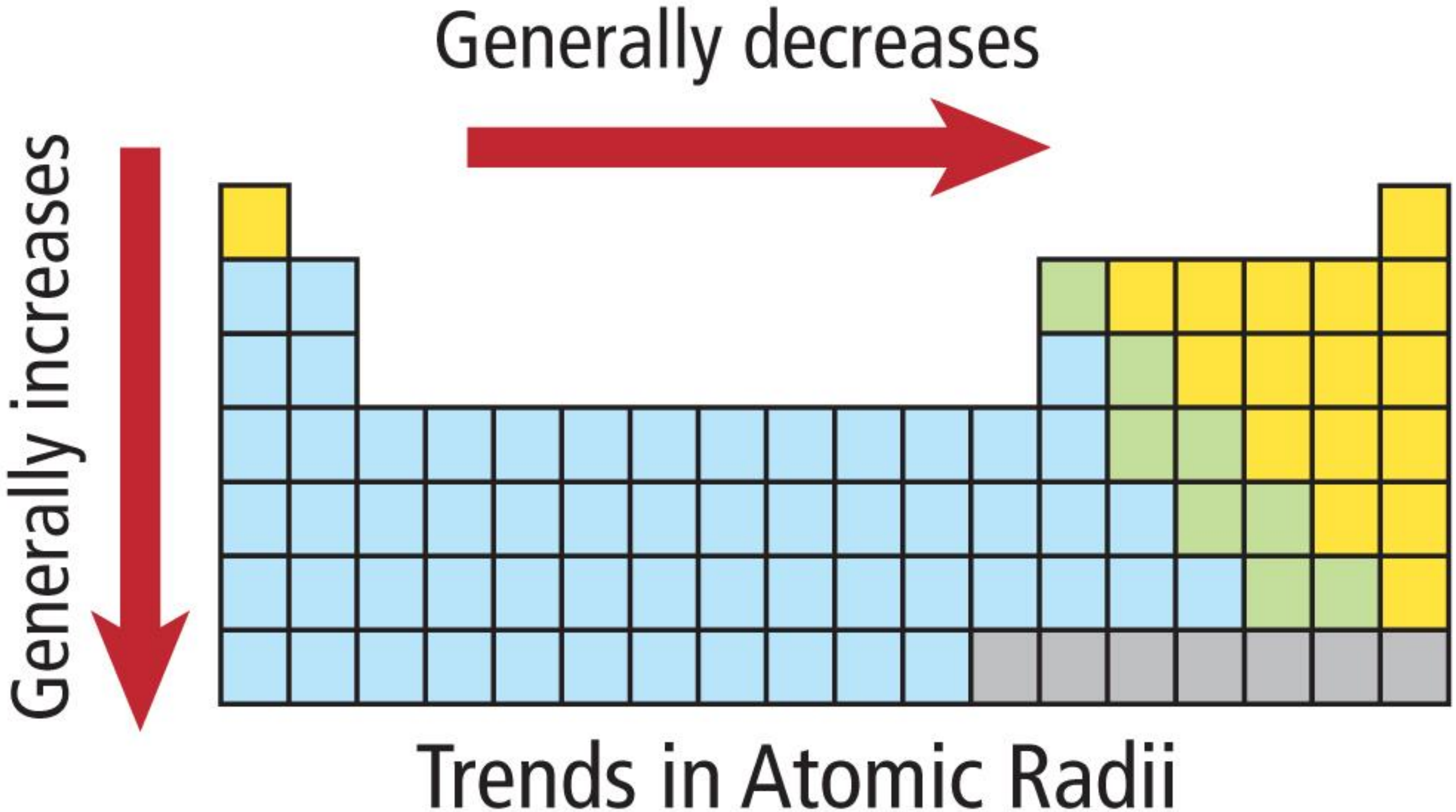




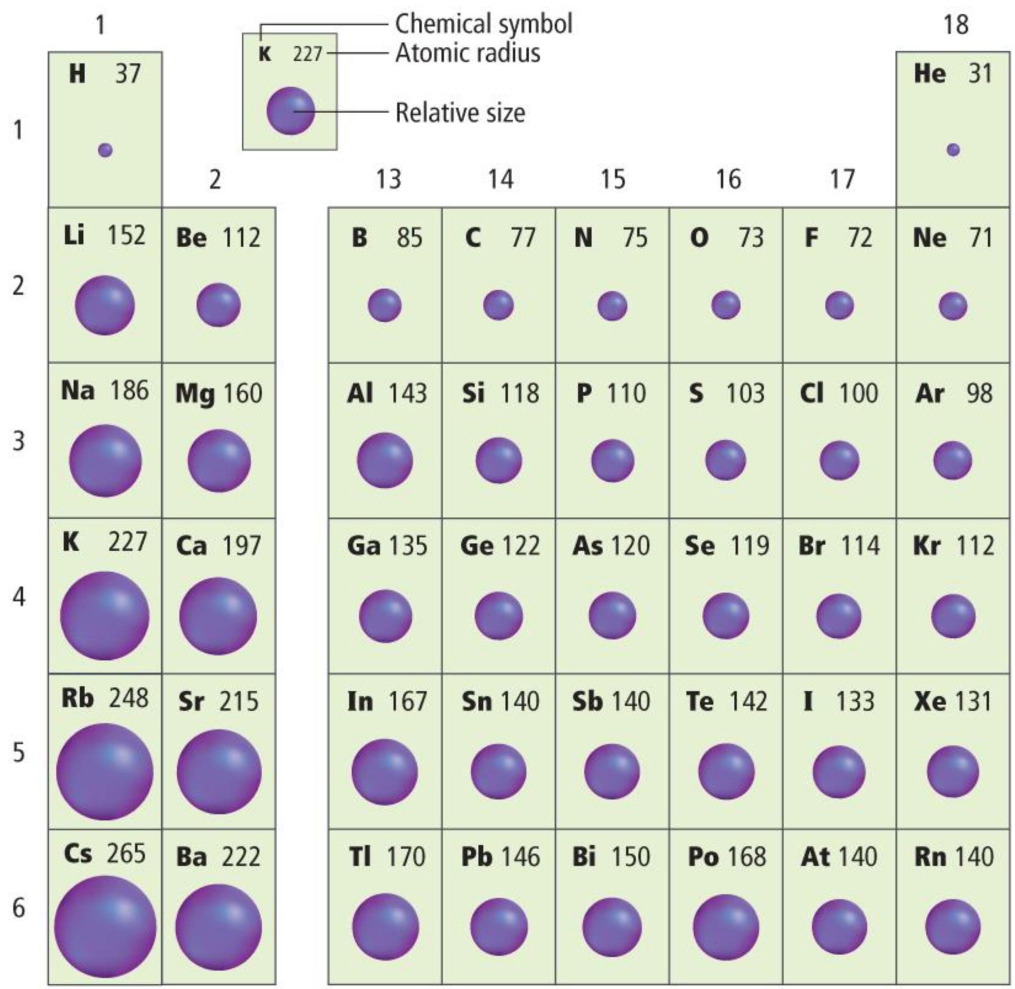
2. Sizes of atoms

- Periodic trend: atomic radii increase moving down a group
  - Increasing energy level
- Periodic trend: atomic radii decrease moving left to right in a period
  - The effective nuclear charge felt by the valence electrons becomes larger

Atomic Radius

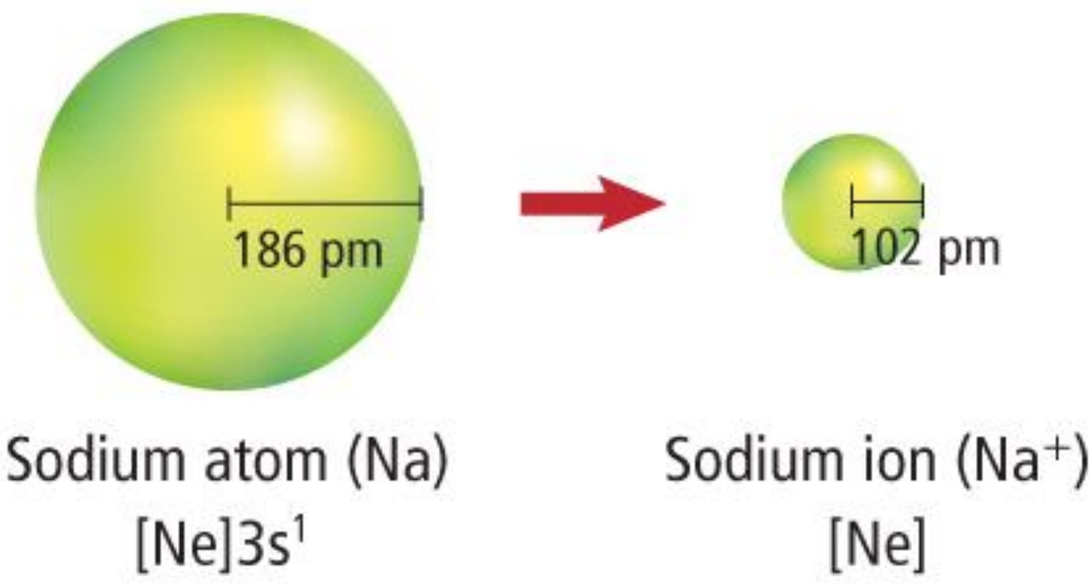


Atomic Radius



3. Sizes of ions

- <u>Periodic trend</u>: cations are always smaller than the atom they were formed from



Ionic Radius

- When atoms **lose electrons** and form *positively charged ions*, they always become **smaller** for two reasons:

1. The loss of a valence electron(s) can leave an empty outer energy level resulting in a small radius.

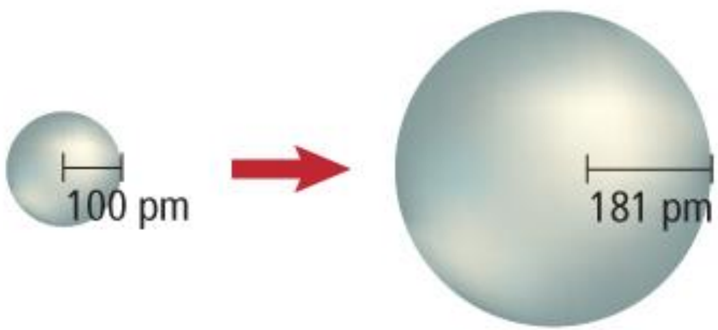
2. Electron/electron repulsion decreases allowing the electrons to be pulled closer to the nucleus.

3. Sizes of ions

- Periodic trend: anions are always larger than the atom they were formed from
  - Electrons repel each other

Ionic Radius

- When atoms gain electrons, they can become larger, because the addition of an electron increases e-/e- repulsion.

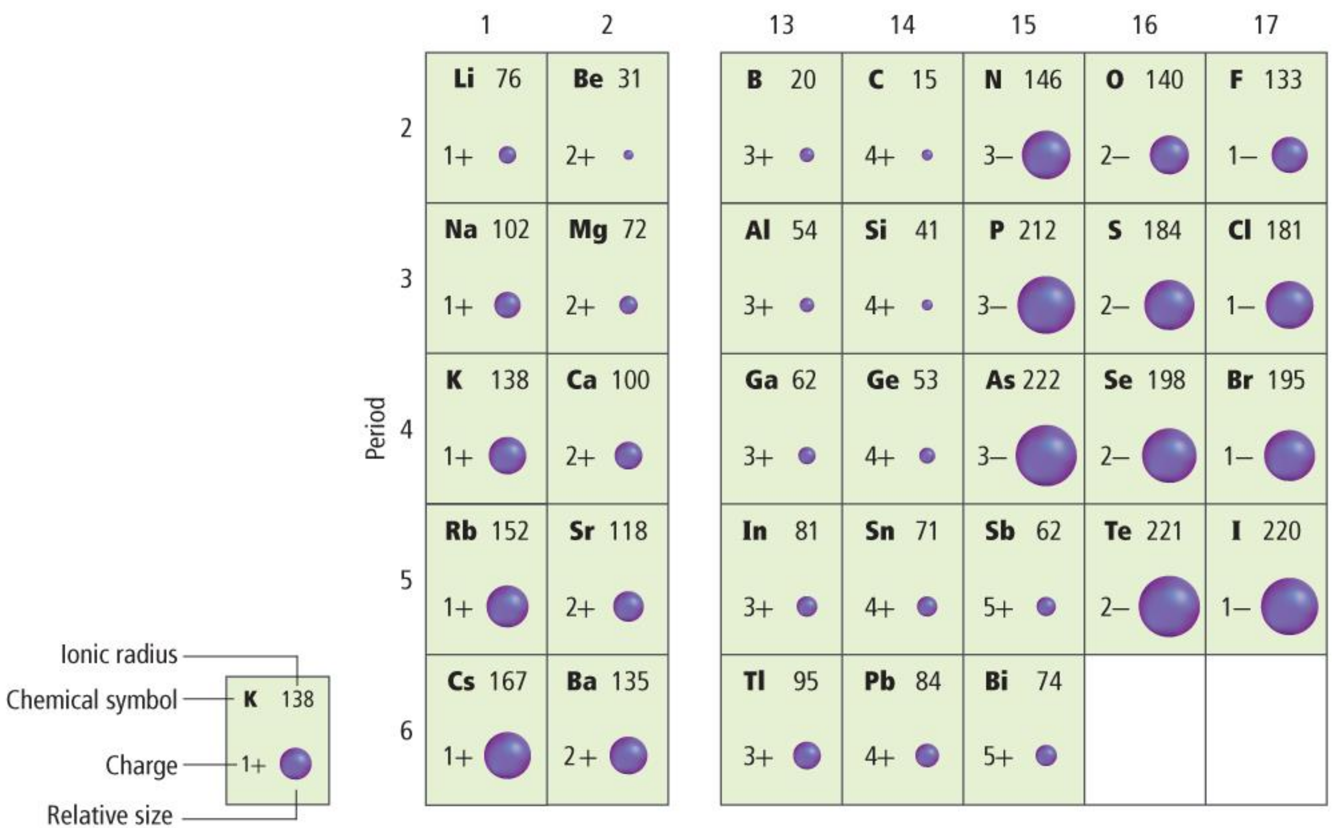


Chlorine atom (Cl)
[Ne]$3s^2 3p^5$

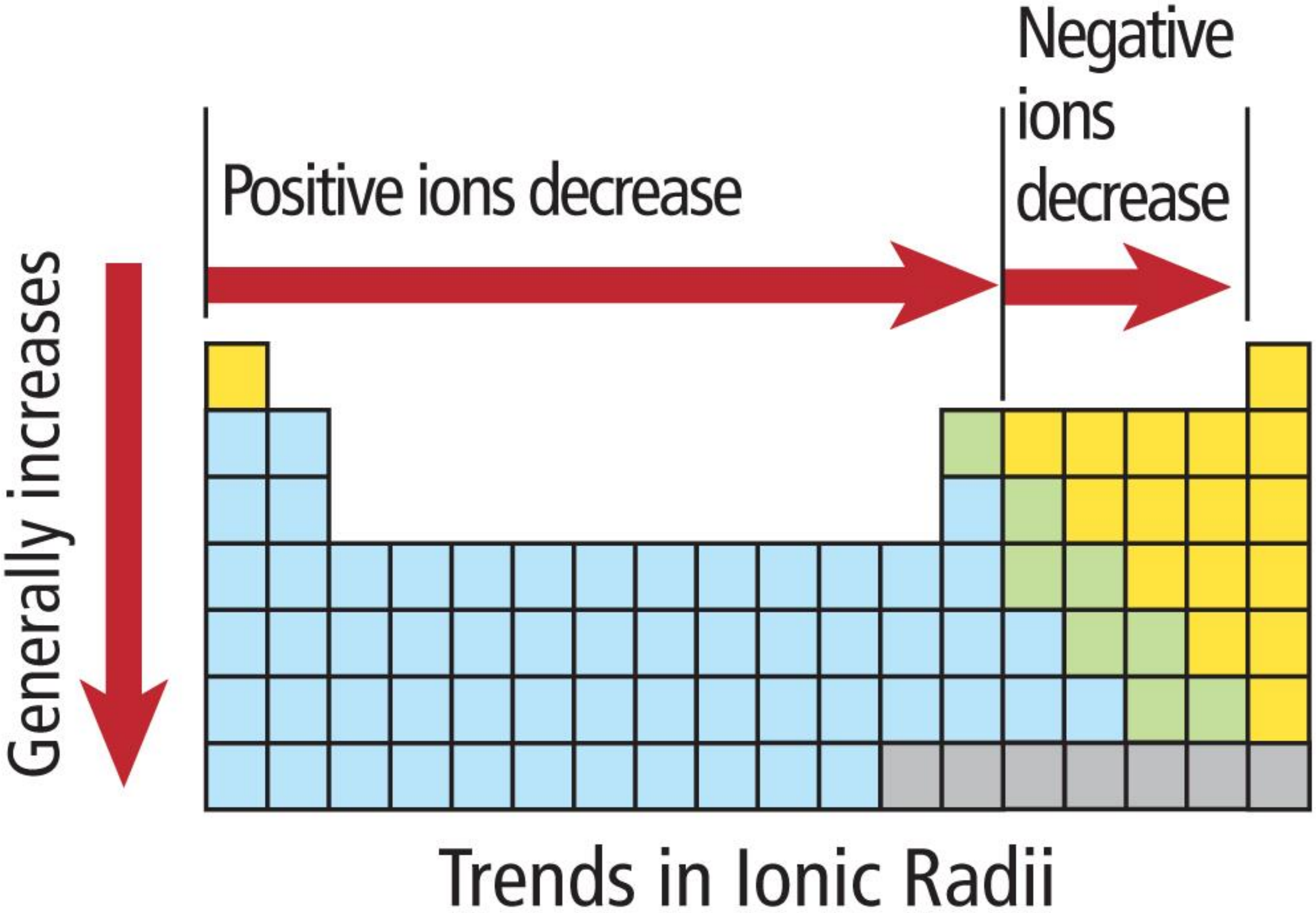
Chlorine ion (Cl^-)
[Ne]$3s^2 3p^6$ or [Ar]

Ionic Radius

- Both positive and negative ions increase in size moving down a group.



Ionic Radius



Some more "periodic trends"

4. Ionization energy

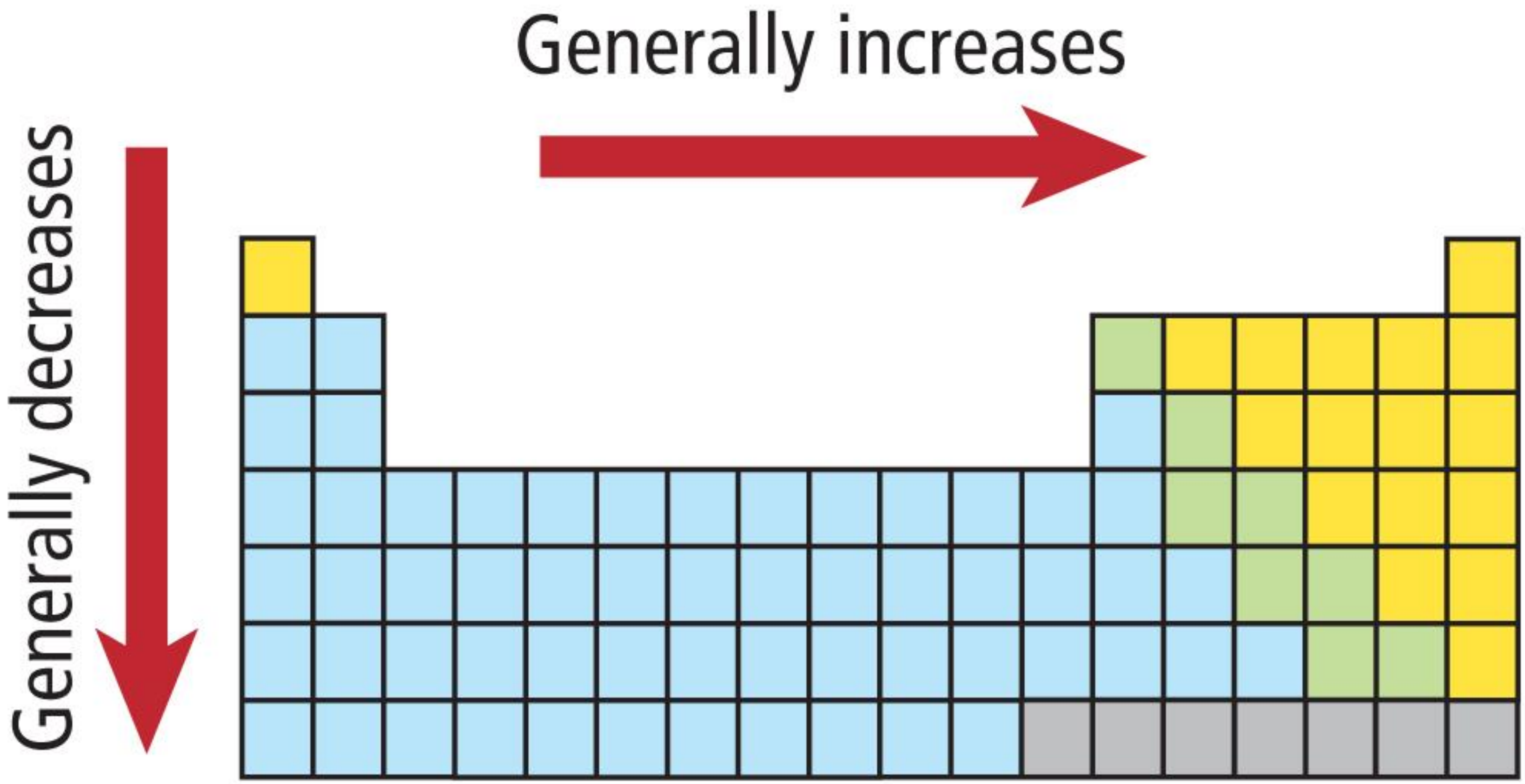
- The energy needed to remove an electron from an atom
- A measure of how tightly the electrons are being held
- **$M \rightarrow M^+ + e^-$**

4. Ionization energy

periodic trend:

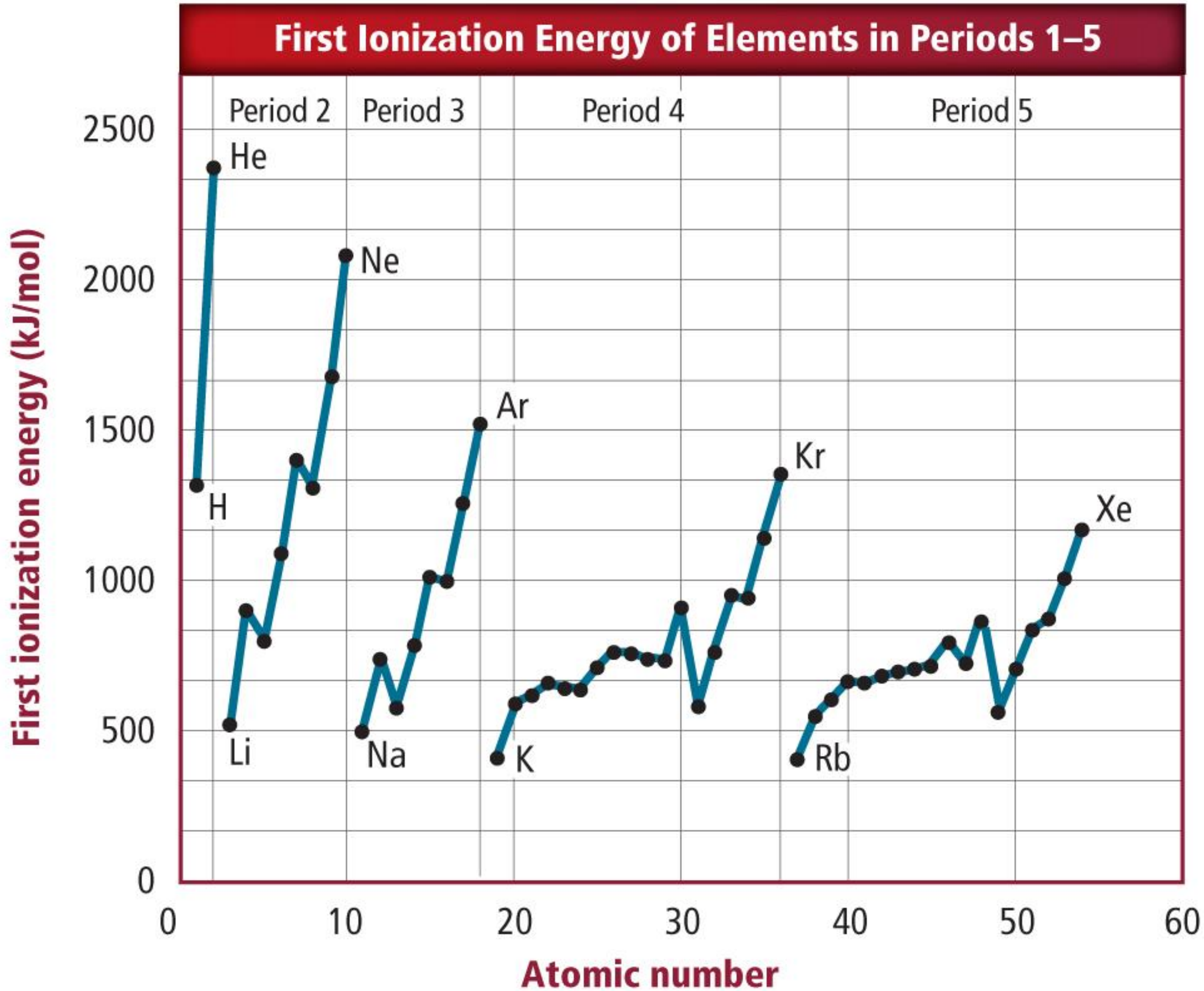
- decreases moving down a group
  - atomic size increases
  - electron farther from the nucleus.
- increases left to right
  - Greater effective nuclear charge
  - Electrons held more tightly

Ionization energy



Trends in First Ionization Energies

Ionization energy



- In general, metals have lower IE than nonmetals
  - alkali metals are the lowest IE family
  - noble gases are highest IE family

Ionization energy

- The energy required to remove the first electron is called the *first ionization energy*.

$$\mathbf{M \rightarrow M^{+} + e^{-}}$$

- Removing the second electron requires more energy, and is called the second ionization energy.

$$\mathbf{M^{+} \rightarrow M^{2+} + e^{-}}$$

- Each successive ionization requires more energy, but it is not a steady increase.

- The ionization at which the large increase in energy occurs is related to the number of valence electrons.

Ionization energy

Table 6.5 **Successive Ionization Energies for the Period 2 Elements**

Element	Valence Electrons	Ionization Energy (kJ/mol)*								
		1st	2nd	3rd	4th	5th	6th	7th	8th	9th
Li	1	520 ⇔	7300							
Be	2	900	1760 ⇔	14,850						
B	3	800	2430	3660 ⇔	25,020					
C	4	1090	2350	4620	6220	37,830				
N	5	1400	2860	4580	7480	9440	53,270			
O	6	1310	3390	5300	7470	10,980	13,330	71,330		
F	7	1680	3370	6050	8410	11,020	15,160	17,870	92,040	
Ne	8	2080	3950	6120	9370	12,180	15,240	20,000	23,070	115,380

* mol is an abbreviation for mole, a quantity of matter.

5. Electron affinity

- A measure of how strongly an element would like to gain an electron
- periodic trend
  - increases from the bottom up
  - increases left to right
  - ignore the noble gases

Periodic Trends…

- As you move from left to right along a period…
- Atoms get

…. *Smaller*

- Ionization energy goes

…. *Up*

- Electron affinity goes

…. *Up*

PERIODIC TRENDS…

- As you move down a group/family
- Atoms get

…. *Larger*

- Ionization energy goes

…. *Down*

- Electron affinity goes

…. *Down*