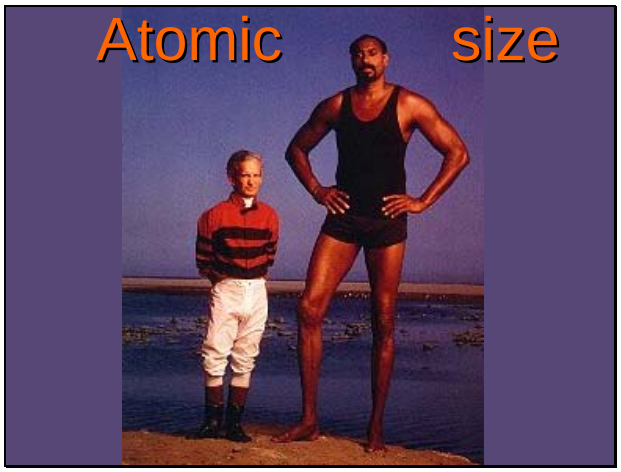
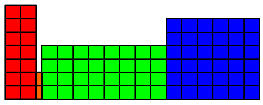


Patterns in Atomic Size

- Atomic size refers to the distance that the furthest (valence) electrons are from nucleus
- It can affect the properties of atoms & elements

Which Is Bigger
a) Ne or Ar
b) B or C
c) P or Ge

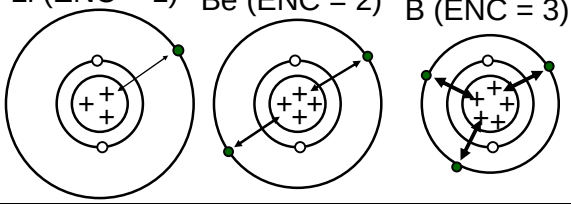


- Electron configuration for Br: $[Ar]4s^2 3d^{10} 4p^5$, valence electron configuration for Br: $4s^2 4p^5$.
- We often only focus on valence electrons: they are involved in reactions and determine size
- Two factors affect size: n and the balance between attractions & repulsions in the atom.

Decreasing Atomic Size Across a Period

- As the attraction between the +ve nucleus and the -ve valence electrons ↑, the atomic size ↓.
- From left to right, size decreases because there is an increase in nuclear charge and Effective Nuclear Charge (# protons – # core electrons)
- Each valence electron is pulled by the full ENC

Li (ENC = 1) Be (ENC = 2) B (ENC = 3)

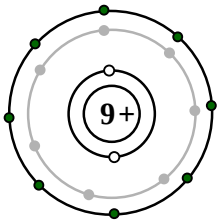


Sizes of ions

- Ions are atoms that have either gained or lost electrons (so that the # of electrons is not equal to the # of protons)
- The size of an atom can change dramatically if it becomes an ion (reference: pg. 214)
- E.g. when sodium loses its outer electron to become Na^+ it becomes much smaller. Why?
- Na^+ is smaller than Na because it has lost its 3s electron. Its valence shell is now $2s^2 2p^6$ (it has a smaller value of n)
- Changing n values is one explanation for the size of ions. The other is …

Sizes of ions: electron repulsion

- Valence electrons push each other away
- When an atom becomes a –ve ion (adds an electron to its valence shell) the repulsion between valence electrons increases without changing ENC



- Thus, F^- is larger than F
- Sort from largest to smallest: Mg, Mg^+, Mg^{2+}. Explain your answer. Pg. 215 PE 11
- pg. 221 6.79, 6.82, 6.83, (6.84), 6.85

Ionization energy

- Ionization energy is the energy required to remove an electron from a gaseous atom
- If n is small & ENC is large, electrons will be difficult to remove (i.e. the IE will be high)
- There are as many IEs as there are electrons
- Subsequent IEs are higher than the first because you are removing a -ve charge (electron) from an increasingly +ve atom/ion
- Subsequent IEs make a huge jump after the electrons in the outer shell are lost - it is not difficult for Mg to lose 12th and 11th electron, but very difficult for it to lose it's 10th electron.
- If you are asked for a trend in IE, talk about 1st

Electron Affinity

- Ionization energy: 6.12, electron affinity: 6.13
- Electron affinity is the energy related to adding an electron to a gaseous atom
- Represented as $X(g) + e^- \rightarrow X^-(g)$
- Whereas IE is: $X(g) \rightarrow X^+(g) + e^-$
- The trend for EA is the same as that for IE
- Imagine an atom with a high IE. It is difficult to remove an electron (due to a small size or high ENC); so, it will also be easy to add a new one
- Noble gases do not follow the trend in EA (a filled valence shell makes it energetically unfavorable to add an electron)
- PE 12 (pg 217), RE (pg 221) 6.86 – 6.90

Energy: exothermic, endothermic

- Energy can be described according to whether we are gaining or losing energy
- Endothermic: requires energy (given a + sign)
  E.g. lifting a book, removing an electron
- Exothermic: gives off energy (given a – sign)
  E.g. dropping a book.
- IE is positive (it takes energy to remove an e^-)
- 1st EA is negative (energy is given off – i.e. it is energetically favorable to add an electron)
- After 1st EA, energy may be required to add electrons to an increasingly negative atom/ion
- Note an EA of –200 is greater than –100