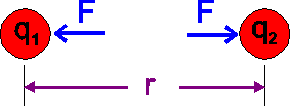
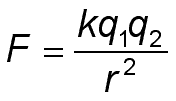
Electrons in Atoms

1. Coulomb’s Law Any two charged objects will create a force on each other.

-Opposite charges will produce an attractive force while similar charges will produce a force.

-The greater the charges, the greater the force.  The greater the distance between them, the smaller the force.

For two spherically shaped charges the formula would look like

In chemistry this means:

* For like charges, the potential energy (E) is positive and decreases as the particles get farther apart as *r* increases.
* For opposite charges, the potential energy is negative and becomes more negative as the particles get closer together.
* The strength of the interaction increases as the size of the charges increases.

Electrons are more strongly attracted to a nucleus with a 2+ charge than a nucleus with a 1+ charge



Compare the relative energy necessary to separate positive and negative

electrical charges in the following situations? Compare a and b, then compare

a and c.

2. Attraction/Repulsion- In a many-electron atom, each electron is simultaneously:

* ***attracted*** to the protons in the nucleus
* ***repelled*** by other electrons (like-charge repulsion)

*a.***Shielding**-Shielding electrons (core) are the electrons in the energy levels **between** the nucleus and the valence electrons. They are called "shielding" electrons because they "shield" the valence electrons from the force of attraction exerted by the positive charge in the nucleus.

b. **Effective nuclear charge**-Zeff the net positive charge experienced by an [electron](http://en.wikipedia.org/wiki/Electron) in a multi-electron [atom](http://en.wikipedia.org/wiki/Atom).

* Any core electrons (not valence electrons) will reduce the nuclear charge acting on valence electrons.
* The effective nuclear charge (***Z*eff**) equals the number of protons in the nucleus (***Z***), minus the average number of electrons (***S***) that are between the electron in question and the nucleus

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* The positive charge "felt" by the outer electrons is ***always less than the full nuclear charge*** (inner electrons "screen" the nuclear charge).

Example: Determine the ***Z*eff** for the valence electrons of Na and S.

**Na** Zeff = 11- 10 = +1 Na feels a small net positive charge and readily gives up its electron.

**S** Zeff = 16 – 10 = +6 S feels a big net positive charge and not only keeps its electrons, but recruits others.

3. Penetration- The closer an electron is to the nucleus, the more attraction it experiences.

The better an outer electron is at penetrating through the electron cloud of inner electrons, the more attraction it will have for the nucleus.

Therefore, for the same shell value (n) the penetrating power of an electron follows this trend in subshells:

s>p>d>f

And for different values of shell (n) and subshell (l), penetrating power of an electron follows this trend:

1s>2s>2p>3s>3p>4s>3d>4p>5s>4d>5p>6s>4f....

and the energy of an electron for each shell and subshell goes as follows...

1s<2s<2p<3s<3p<4s<3d<4p....

1. which atom has a stronger effective nuclear charge and why? (assuming S is the same in both cases)

Li, or N

2) why does the Hydrogen electron experiences the full charge of the nucleus without any shielding?

3) Which atom has a smaller radii?

Be or F?

4) Which electron has higher energy level? 2s or 2p? and why?

1) nitrogen atom has a stronger effective nuclear charge than lithium due to its greater number of protons in the nucleus holding the electrons tighter.

2)Hydrogen atom has only one electron total, therefore there are no other, lower energy (more penetrating), electrons available to help shield this electron from the nucleus.

3) Fluorine has a smaller radii than Beryllium due to its greater number of protons providing a greater effective nuclear charge on the outer electrons and therefore pulling them in tighter and providing a smaller atomic radii.

4)2p has higher energy level because the negatively charged electron experiences less of an effective nuclear charge than the 2s electron.