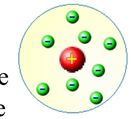
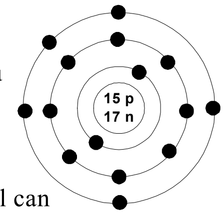
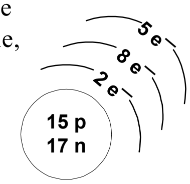
The Bohr-Rutherford Atom



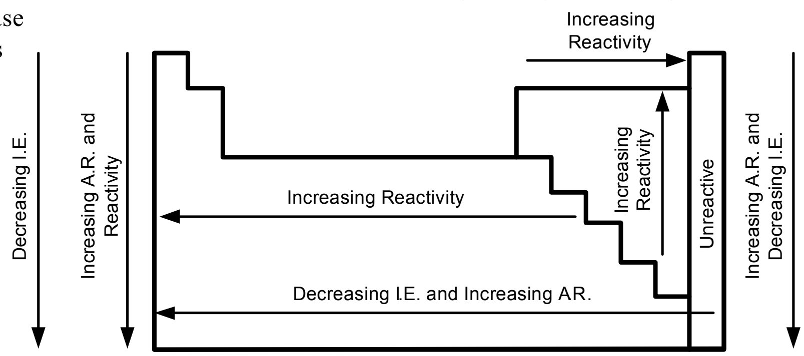
The Bohr-Rutherford arose out of a need to reconcile why moving electrons did not lose energy and eventually, due to electrostatic attraction to the positively charged protons in the nucleus, collapse and adhere to the nucleus. Also, since some electrons orbit the nucleus near the outer reach of the attractive force, there needed to be a reason why, occasionally, an electron did not escape the atom. The breakthrough came when it was discovered that after energized electrons lose a small amount of energy, they return to their original orbital distance from the nucleus or ground state and as they do so, they emit specific wavelengths of light which indicated that electrons were only capable of moving discrete jumps in distance from the nucleus. These jumps were called energy levels.



Further data analysis revealed that electrons always occupy lowest possible energy level (Aufbau principle) and that the maximum number of electrons for each energy level was defined by a unique formula, $2n^2$ where n is level number. To draw a representation of a Bohr-Rutherford atom, the number of protons and neutrons is indicated within the nucleus and the number of electrons found within each successively higher energy levels is indicated in some fashion. By convention, the electrons within a level can be paired (as electrons do exist in pairs) or separated (due to repulsion) or simply stated as a number because this is just a representation; all that is needed is the number of each sub-atomic particle present. For example, consider two acceptable versions of a neutral atom of phosphorus with 17 neutrons (see adjacent diagrams).



More importantly, this consistent sub-atomic structure also revealed that atoms differed internally in a predictable pattern; that is, trends. This is largely the result of the interplay between the two electromagnetic forces, attractions between oppositely charged electrons and protons and repulsion between negatively charged electrons (protons are locked within the nucleus). The force of attraction generally is more significant as proton number increases, but electrons only increase within the same level; whereas, repulsion becomes more significant as whole new levels of electrons are created.



Typically, the first trend investigated is **atomic size**. The size of an atom can be determined many ways. One way is to measure the volume of a sample with a known number of atoms and calculate the volume of a single atom and subsequently the radius of that atom. Another way involves determining the volume of a known number of molecules and calculating the volume occupied by a single molecule to subsequently determine the radius of one atom. Another way is to determine the minimum pore diameter an atom can fit through without significantly altering electron distribution within the atom. As a result answers don't always agree, but the trend in atomic size is relatively identical depending upon the method of measurement used. Atoms tend to get larger moving down the families/groups in the periodic table as outermost electrons become substantially further from nucleus. Atoms tend to get smaller moving right in the periodic table as protons as distance to the same outermost level decreases. Finally, additions of electrons to an atom (**anion** formation) increases repulsion for the same attraction so atomic size increases while removal of electrons from an atom (**cation** formation) decreases repulsion for the same attraction resulting in a decrease in size.

Another common periodic trend studied is **ionization energy** which is the minimal amount of energy required to remove the most loosely held electron in a gaseous atom. The easier it is to remove an electron, the lower the ionization energy. Thus the further the electron is from the nucleus, the easier it is to remove and; therefore, this trend relates directly to atomic size. Larger atoms have lower ionization energies (moving down or from right to left in a period in the periodic table).

Positive electron affinity is amount of energy released when a neutral atom gains an electron (positive because there is a gain in energy to the surroundings). Atoms further down a family/group have less positive electron affinities due to a diminishing force of attraction and increasing electron-electron repulsion, so there is more resistance to additional electrons. Atoms further to the right in a period also have more positive electron affinities because of increasing force of attraction for the almost the same amount of electron-electron repulsion.

The sum of all the periodic trends is expressed via **atomic reactivity**. When ionization energy and positive electron affinity are very low, atoms tend to react by losing electrons characteristic of metals. The lower the ionization energy, the more reactive the metal. When ionization energy and positive electron affinity become very high, atoms tend to react by gaining electrons characteristic of non-metals. The higher positive electron affinity the more reactive the non-metal. The dividing line between losing and gaining is marked by the staircase line in the periodic table. The noble gases are not reactive (to any significant extent). These elements have very high ionization energies and negative (extremely non-positive) electron affinities (energy must be used to force them to hold an extra electron). Thus, they do not lose or gain electrons and hence they are not reactive.

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