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Chapter 6 – The Periodic Table

Section 6.1 – Organizing the Elements

1. Early chemists used the properties of elements to sort them into groups.
   a. There were only 13 elements identified by 1700.
      i. Chemists suspected that other elements existed
      ii. Began using scientific methods to search for elements and between 1765 and 1775, five new elements were discovered (including 3 colorless gases: hydrogen, nitrogen, and oxygen)
   b. In 1829, J. W. Dobereiner published a classification system. In his system, the known elements were grouped into triads
      i. A triad is a set of three elements with similar properties
2. Mendeleev
   a. In 1869, Dmitri Mendeleev published a table of 60 elements.
      i. He was a chemist and teacher and wanted to have a table for his students
      ii. He wrote the properties of each element on a separate note card and then moved the cards around until he found a pattern that worked.
      iii. Arranged the elements in his periodic table in order of increasing atomic mass
         1. He developed his table before scientists knew about the structure of atoms
         2. He didn't know that the atoms of each element contain a unique number of protons (basis for the atomic number)
      iv. Left spaces in his table (marked with a ?) as he predicted that elements would be discovered to fill those spaces and he predicted what their properties would be based on their locations in the table.
         1. By 1886, those 3 elements were discovered and had properties very similar to the ones Mendeleev predicted they would have.
   b. Later that same year, Lothar Meyer published a periodic table that was nearly identical.
   c. Mendeleev was given more credit than Meyer because Mendeleev's table was published first and he was better able to explain its usefulness.

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3. Moseley
   a. In 1913, Henry Moseley determined an atomic number for each known element.
      i. He discovered that there should be a slightly different pattern than what Mendeleev had used. The reason for the pattern wasn't based on the atomic mass but instead on the number of protons in the nucleus.
      ii. His work led to the modern definition of the atomic number and the recognition that the atomic number (not atomic mass) is the basis for organization.
   b. In the modern periodic table, the elements are arranged in order of increasing atomic number.
4. Periodic Law: When the elements are arranged in order of increasing atomic number, elements with similar properties appear at regular intervals. The physical and chemical properties of the elements are functions of their atomic number.
5. Periodicity – the arrangement of the electrons around the nucleus explains the patterns we see.
6. International Union of Pure and Applied Chemistry (IUPAC) is an organization that sets standards for chemistry.
   a. They numbered the groups from left to right as Groups 1 through 18
7. The three classes of elements are metals, nonmetals, and metalloids
   a. Metals
      i. 80 % of the Periodic Table
      ii. Good conductors of heat and electric current
         1. Silver and Copper are the best conductors of electric current
      iii. Have high luster or sheen (ability to reflect light) when freshly cleaned or cut
      iv. All are Solids at room temperature (except Mercury)
      v. Many are ductile – able to be drawn into wires
      vi. Most are malleable – able to be hammered into thin sheets without breaking.

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b. Nonmetals

i. Upper right hand corner of the periodic table

ii. Great variation

1. Most are gases, a few are solids, one is a liquid

iii. Poor conductors of heat and electric current

1. One exception - Carbon, in the form of graphite, is able to conduct heat and electric current.

iv. Solid nonmetals tend to be brittle (will shatter if hit with a hammer)

c. Metalloids

i. Stair steps that separate the metals from nonmetals

ii. Have properties that are similar to those of metals and nonmetals

1. Under some conditions, a metalloid may behave like a metal. Under other conditions, a metalloid may behave like a nonmetal.

iii. Referred to as semi-Conductors

iv. One example

1. Silicon – poor conductor of electric current, but if it is mixed with a small amount of boron, the mixture is a good conductor of electric current.

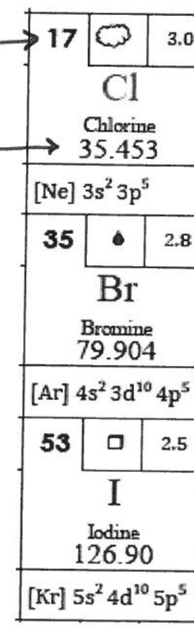
a. Glass items are made from silicon dioxide

b. Silicon can be cut into wafers and used to make computer chips

Section 6.2 – Classifying the Elements

Atomic #

Average atomic mass



8. Reading the Periodic Table

a. The periodic table usually displays the symbols and names of the elements, along with information about the structure of their atoms

i. Atomic number (top number) is above the symbol

ii. In the center is the symbol for the element, with the name underneath

iii. Average atomic mass (bottom number)

b. Most will show if the element is a solid, liquid, or gas at room temperature

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c. Some groups on the Periodic Table have special names
   i. Group 1 (1A) – Alkali metals
   ii. Group 2 (2A) – Alkaline earth metals
   iii. Group 17 (7A) – Halogens
   iv. Group 18 (8A) – Noble gases

d. Elements can be sorted into noble gases, representative elements, transition elements, or inner transition metals based on their electron configurations
   i. Noble Gases
      1. Group 18 (8A) – the noble gases are nonmetals that rarely take part in reactions. They are very unreactive. Sometimes called inert gases.
      2. All noble gases (except He) have 8 valence electrons. $\#s^2$ $\#p^6$
   ii. Representative Elements
      1. Labeled as Groups 1-17 (or 1A through 7A)
      2. Also known as "A groups"
      3. Called representative elements because they display a wide range of physical and chemical properties
      4. Have from 1 to 7 valence electrons
         a. s and p sublevels in the highest occupied energy level are partially filled
   iii. Transition Elements
      1. Labeled as "B groups"
      2. d block
      3. In atoms of a transition metal, the highest occupied s sublevel and a d sublevel contain electrons. Valence electrons are based on the number of electrons in the s sublevel.
   iv. Inner Transition Elements
      1. Elements that appear below the main body of the periodic table
      2. f block (lanthanides and actinides)
      3. Some of these are not found in nature. These elements were prepared in laboratories (using methods discussed in Chapter 25)

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e. Electron Configuration and the Periodic Table

i. The period an element is in tells you the highest occupied principal energy level (the number in front of the s or p).

ii. s block – Groups 1 and 2 (1A and 2A) and Helium

iii. p block – Groups 13-17 (3A – 7A) except for Helium

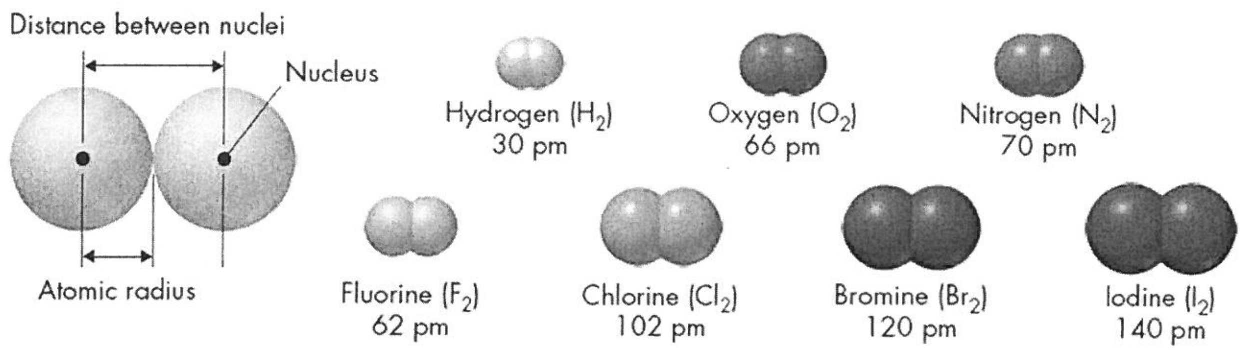
iv. d block – Groups 3-12 (B groups), transition elements

v. f block – two rows below the periodic table, inner transition elements

Section 6.3 – Periodic Trends

9. Atomic Radius – size of an atom

a. Atomic radius is one - half of the distance between the nuclei of two atoms of the same element when the atoms are joined.



b. The distances between atoms in a molecule are extremely small.

c. Often measured in picometers (pm)

i. One trillion picometers in a meter (1×10^{12} pm = 1m)

d. Periodic Trend of Atomic Radius

i. Group – atomic size increases from top to bottom in a group

1. Increases are there are more orbitals of electrons as you go down a group

ii. Period – atomic size decreases from left to right across a period

1. Decreases as the nucleus has more protons and is able to attract the electrons closer to the nucleus.

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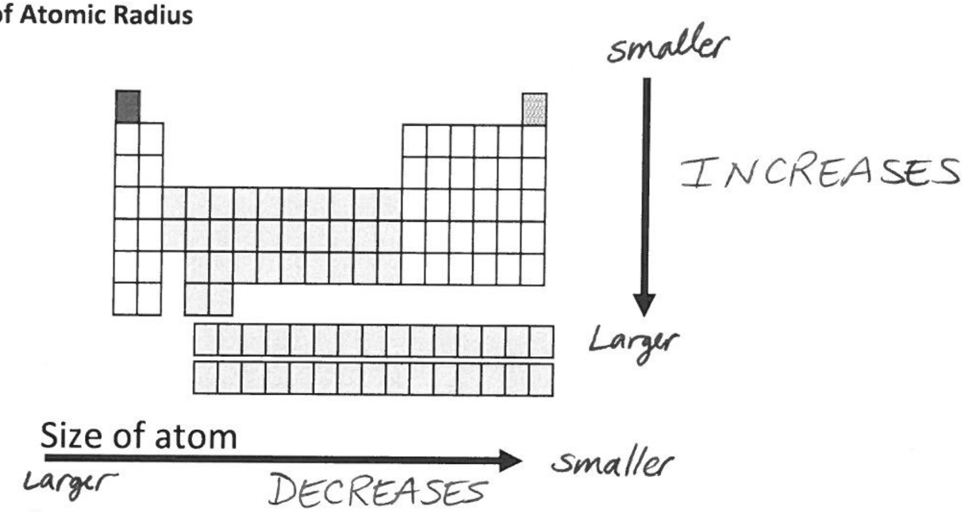
Periodic Trend of Atomic Radius



10. Review of Ions

a. Some compounds are composed of particles called ions

i. An ion is an atom or group of atoms that has a positive or negative charge

ii. This charge is due to a change in the number of electrons

1. Neutral atoms have an equal number of protons and electrons

b. Cations – ions with a positive charge

i. Result of losing an electron

ii. Group 1 (1A) elements will form cations with a +1 charge

iii. Group 2 (2A) elements will form cations with a +2 charge

iv. Group 13 (3A) elements will form cations with a +3 charge

v. Metals tend to form cations

c. Anions – ions with a negative charge

i. Group 15 (5A) elements will form anions with a -3 charge

ii. Group 16 (6A) elements will form anions with a -2 charge

iii. Group 17 (7A) elements will form anions with a -1 charge

iv. Nonmetals tend to form anions

11. Ionization Energy – amount of energy required to remove an electron from an atom

a. First ionization energy – the energy required to remove the first electron from an atom

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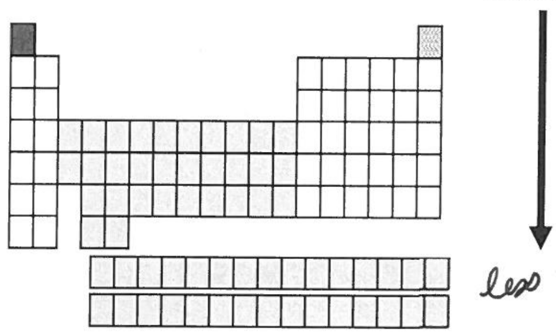
b. Periodic Trend of First Ionization Energy
   i. Group – first ionization energy tends to __decrease__ from top to bottom within a group
      1. Atoms have more __total__ electrons as you go down a group
   ii. Period – first ionization energy tends to __increase__ from left to right across a period
      1. Atoms want to __gain__ electrons as you approach the nonmetals

c. For each __successive__ electron that is removed, it takes __more__ energy
   i. With each electron removed, it takes more __energy__ as the more positive nucleus has a __stronger__ pull on the remaining electrons

Periodic Trend of Ionization Energy

more energy



DECREASES

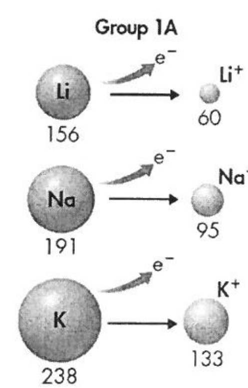
less energy

Energy needed to lose e-

less energy → more energy

INCREASES

12. __Ionic__ __Radius__ – size of ions



a. __Cations__ are always __smaller__ than the atoms from which they form
   i. __Lose__ electrons to have a positive charge, so they are __Smaller__ than the neutral atom

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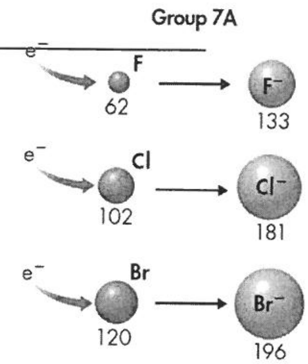


b. Anions are always larger than the atoms from which they form

i. Gain electrons to have a negative charge, so they are larger than the neutral atom

c. Cations < Neutral atoms < Anions

d. Periodic Trend of Ionic Radius

i. Group – ionic size increases from top to bottom within a group

1. Increases as there are more orbitals of electrons as you go down a group

ii. Period – ionic size decreases from left to right across a period

1. Elements on the left hand side (metals) tend to form cations. As they get rid of 1, 2 or 3 electrons, their size decreases.

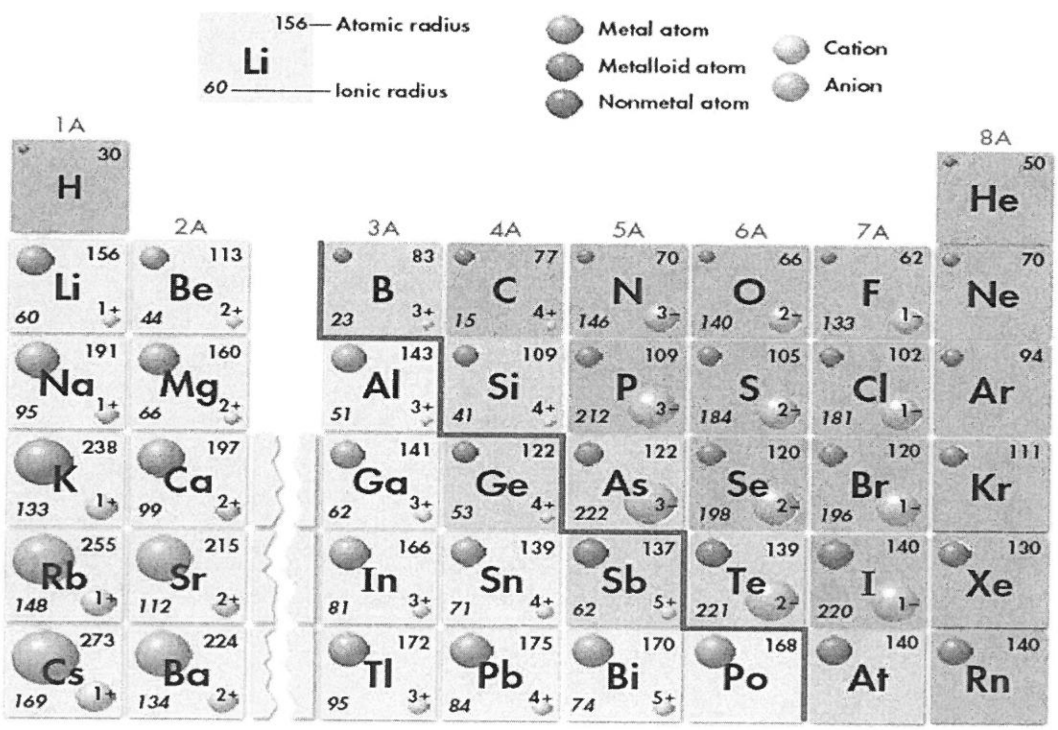
2. Elements on the right hand side (nonmetals) tend to form anions.

a. Sizes decrease from left to right within nonmetals as they are gaining fewer electrons

i. Group 15 (5A) gain 3 electrons

ii. Group 16 (6A) gain 2 electrons

iii. Group 17 (7A) gain 1 electron



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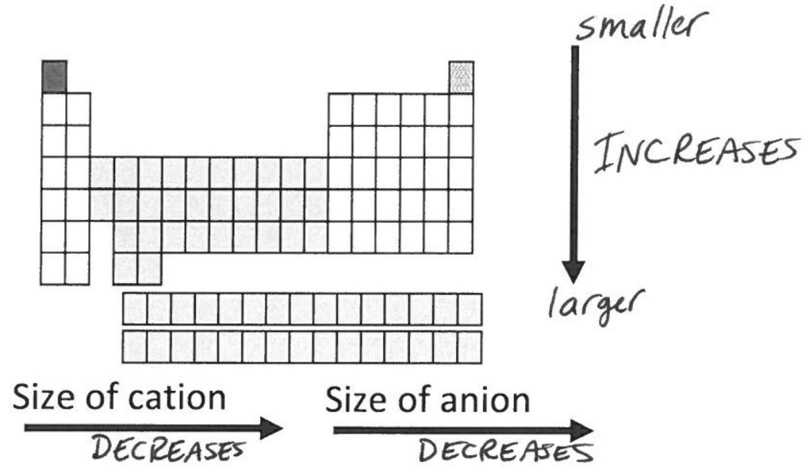
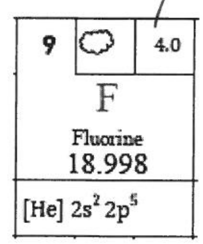
Periodic Trend of Ionic Radius



13. Electronegativity – ability of an atom in a Compound to attract electrons

a. Differences in electronegativity values are used to predict what type of bonds will form between elements. Listed in the upper right hand corner of your colorful periodic table.

electronegativity value



b. Periodic Trend of Electronegativity

   i. Group – electronegativity decreases from top to bottom within a group

      1. More orbitals of electrons

   ii. Period – electronegativity increases from left to right across a period

      1. Metals tend to have low values
      2. Nonmetals (except for Noble gases) tend to have high values
      3. Noble gases are excluded
      4. Transition metals can vary

c. Least electronegative elements – Cesium and Francium (0.7)

   i. Bottom of Group 1
   ii. Have the least tendency to attract electrons. When they react with other elements, they tend to lose electrons and form cations.

d. Most electronegative element – Fluorine (4.0)

   i. Top of Group 17 (7A)

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Periodic Trend of Electronegativity