1. Dmitri Mendeleev
	1. Russian
	2. Invented periodic table
	3. Organized elements by properties
	4. Arranged elements by atomic mass
	5. Predicted the existence of several unknown elements
	6. Element 101 named in his honor
2. Henry G.J. Moseley - 1914
	1. Created the modern periodic table using Mendeleev’s periodic table
	2. Determined the atomic numbers of elements from their X-ray spectra
	3. Arranged elements by increasing atomic number
3. Introduction to the Periodic Table
	1. Elements are arranged in seven horizontal rows, in order of increasing atomic number from left to right and from top to bottom
	2. Rows are called periods and numbered from 1 to 7
	3. Elements with similar chemical properties form vertical columns, called groups or families, which are numbered from 1 to 18
	4. Groups 1,2, and 13-18 are the main group elements (Representative Elements)
		1. s and p orbitals
	5. Groups 3 -12 are in the middle of the periodic table and are known as the transitions elements.
		1. d orbital
	6. The two rows of 14 elements at the bottom of the periodic table are the lanthanides and actinides and together are known as the inner-transition metals.
		1. f orbital
4. Identify different parts of the periodic table:
	1. 
5. Chemistry of the groups – Elements with similar chemical behavior are in the same groups.
	1. The alkali metals (Group 1)
		1. The alkali metals are:
		2. Hydrogen is placed in Group 1 but is not a metal
		3. The alkali metals react readily with nonmetal to give ions with a +1 charge
		4. Compounds of alkali metals are common in nature and in daily life
	2. The alkaline earth metals (Group 2)
		1. The alkaline earth metals are:
		2. Are metals that react readily with nonmetals to give ions with a +2 charge
	3. The noble gases (Group 18)
		1. The noble gases are:
		2. Are monatomic
		3. Are unreactive gases at room temperature and pressure
		4. Are call inert gases
	4. The halogens (Group 17)
		1. The halogens are:
		2. They react readily with metals to form ions with a -1 charge.
		3. They are all diatomic
6. Periodic Trends within the Periodic Table - Within the periodic table, there are trends that follow certain patterns.
	1. Atomic Number and Atomic Mass
		1. Atomic Number goes up going from left to right and top to bottom
		2. Atomic Mass follows the trend for the most part
	2. Electronegativity – the ability of an atom in a molecule to attract shared electrons to itself
		1. Like electron affinity – ability of an atom to attract an electron to itself to form an ion
		2. As we go down the periodic table, electronegativity \_\_\_\_\_\_\_\_\_\_\_\_\_.
		3. As we go right in the periodic table, electronegativity \_\_\_\_\_\_\_\_\_\_\_\_\_.
	3. Atomic Radii - the distance from the center of an atom to the outer-most electrons
		1. As we go down the periodic table, atomic radii \_\_\_\_\_\_\_\_\_\_\_\_\_.
		2. As we go right in the periodic table, atomic radii \_\_\_\_\_\_\_\_\_\_\_\_\_. (Why?????)
			1. Coulombic Attraction – attraction between positive and negative charges within an atom
				1. Depends upon:

Amount of charge – more charge = more attraction

Distance between charges – more distance = less attraction

* + - * 1. As we go right in the periodic table = more attraction

Due to no new energy levels of electrons

More charges = more pull

Therefore, smaller size

* + - 1. Shielding Effect – Core electrons block the attractive force of the nucleus from the valence electrons
				1. Therefore, distance from positive and negative charges is greater causing coulombic attraction to be less causing larger atomic radius
	1. Ion Radii – the distance from the center of an ion to the outer-most electrons
		1. Cations are smaller than their neutral counterparts
		2. Anions are larger than their neutral counterparts
		3. Why?
	2. Ionization Energy – Energy required to remove an electron from an atom to form a cation.
		1. Ionization energy is always positive
		2. Larger values mean that the electron is more tightly bound to the atom and harder to remove
		3. As we go down the periodic table, ionization energy \_\_\_\_\_\_\_\_\_\_\_\_\_.
		4. As we go right in the periodic table, ionization energy \_\_\_\_\_\_\_\_\_\_\_\_\_. (Why???)
			1. Nuclear charge – larger the charge, greater the ionization energy
			2. Shielding effect – greater the shielding effect, less ionization energy
			3. Radius – greater the distance between the nucleus and the valence electrons, less ionization energy
			4. Sublevel – electrons from a full or half-full sublevel requires additional energy to be removed
		5. As more electrons are added to an atom, they have the potential to be lost.
			1. The first electron to be lost always requires energy and is known as the first ionization energy
			2. The second electron to be lost always requires more energy to be removed and is known as the second ionization energy
			3. The pattern continues until all electrons are lost

