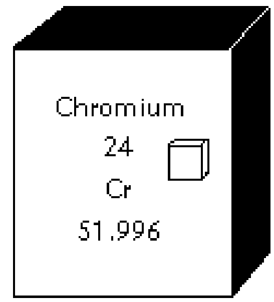
UNIT 4: The Periodic Table

Aristotle — Four element theory: earth, air, fire, water	Circa 300 BC
Antoine Lavoisier — Known as Father of Modern Chemistry; compiled list of 23 to 33elements; some elements later shown to be compounds and mixtures	1770–1789
Jöns Jakob Berzelius — Developed table of atomic weights; introduced element symbols; experiments discovered new elements (Ce, Se, Th)	1828
Johann Döbereiner — Classified groups of elements into triads: three elements with physical and chemical similar properties; led to idea of groups (columns in Periodic Table)	1829
John Newlands — Arranged about 60 known elements by increasing atomic weight; proposed Law of Octaves based on observed similarities between elements; led to idea of periods (rows)	1864
Julius Lothar Meyer — Compiled periodic table of 56 elements based on periodicity of properties when arranged in order of atomic weight	1869-1870
Dmitri Mendeleev — Produced periodic table based on atomic weights and arranged elements with similar properties under each other; known as Father of the Periodic Table; used his table to predict physical properties of three unknown elements	1869-1870
William Ramsay — Discovered the noble gases	1894
Henry Moseley — Determined atomic numbers of elements and modified periodic law to read that properties of elements vary periodically with atomic number	1913
Concluded 92 elements existed up to and including uranium	1914
Glenn T. Seaborg — Synthesized ten transuranium elements (94 thru 102 and 106); developed actinide series	1940s

Periodic Law: there is a periodic repetition of chemical and physical properties of elements when they are arranged by increasing atomic number

MODERN PERIODIC TABLE



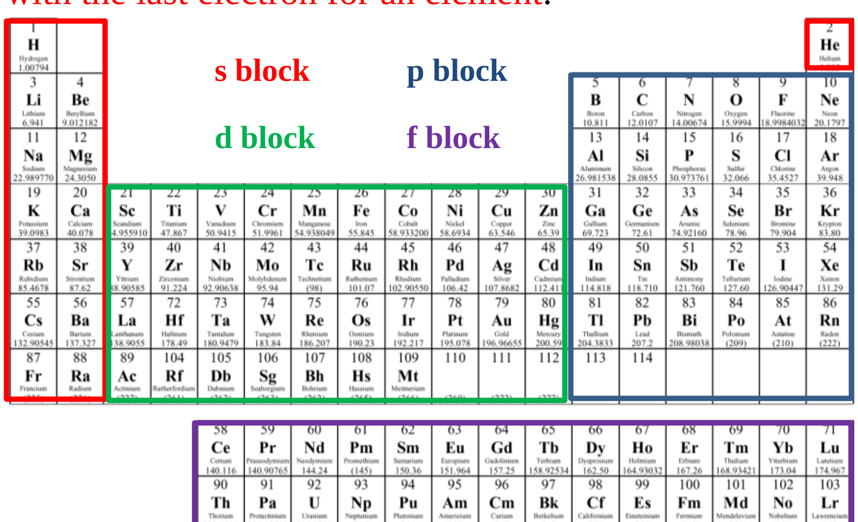
The modern table maintains Moseley's arrangement and clearly shows periodicity, which refers to trends or recurring variations in element properties caused by regular and predictable variations in atomic structure. The periodic table consists of boxes for elements arranged in order of increasing atomic number. Each box contains the element's name and symbol, atomic number, and atomic mass.

The 118 element boxes are arranged in rows called periods and columns called groups or families.

- Total of seven (7) periods, numbered 1 through 7
  - ☞ Correspond to the energy levels for electron configuration
  - ☞ Referred to as principal quantum number, n
  - ☞ Number of electrons in an atom determines period placement
- Two numbering systems (left to right) are in effect for the groups:
  1. Numbered 1 through 18.
  2. Numbered 1 through 8 followed by the letter A or B.
     - ☞ Groups designated with the letter A are known as representative elements. They display a wide range of chemical and physical properties.
     - ☞ Groups designated with the letter B are transition elements.

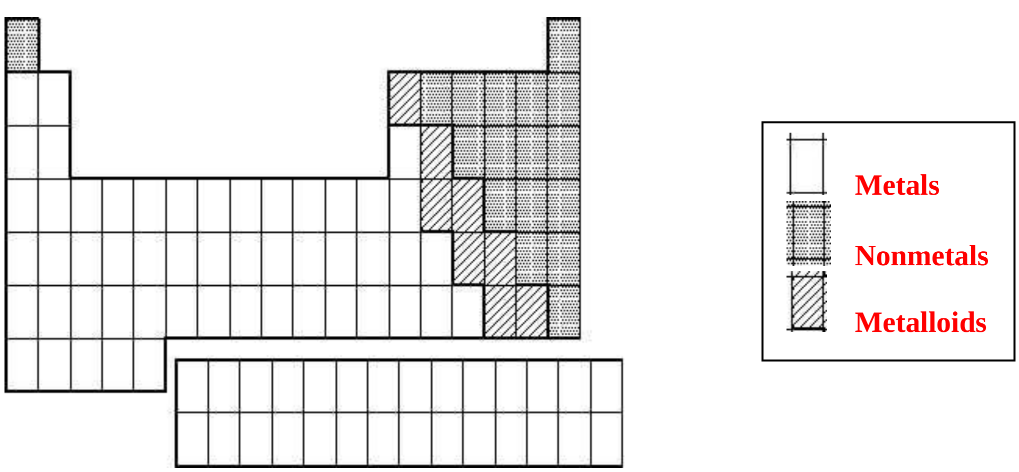
THE s-, p-, d-, AND f-BLOCKS

The periodic table is divided into four blocks representing the energy sublevel being filled with the last electron for an element.



Classification of Elements

	Metals	**Nonmetals**	**Metalloids**
Characteristics	Lustrous (shiny) Solid at room temperature Tend to form cations ♦*Exception*: mercury is liquid metal Good conductors of heat and electricity Ductile, malleable High tensile strength React with acids	Gases or dull, brittle solids at room temperature Tend to form anions ♦*Exception*: bromine is liquid nonmetal Poor conductors of heat and electricity (insulators) Do not react with acids	Display properties of both metals and nonmetals Solid at room temperature High tensile strength Semi-conductors: conduct better than insulators but not as well as metals
Location on PT	Left of the stair step line Most Group A elements ♦*Exception*: hydrogen All Group B elements ♦Transition metals in d block ♦Inner transition metals: lanthanide and actinide series in f block	Right of the stair step line	Border the stair step line ♦*Exception*: aluminum



Groups of Elements

Group	Group Name	Group Properties	Valence e^-	Oxidation Number	Block	e- Conf End
Hydrogen			1	+1	s	s^1
1 or 1A	Alkali Metals	Most reactive metals	1	+1	s	s^1
2 or 2A	Alkaline Earth Metals	Very reactive metals	2	+2	s	s^2
13 or 3A	Boron Group		3	+3	p	p^1
14 or 4A	Carbon Group	↑ in metallic character, ↓group	4	n/a	p	p^2
15 or 5A	Nitrogen Group	↑ in metallic character, ↓group	5	–3	p	p^3
16 or 6A	Oxygen Group		6	–2	p	p^4
17 or 7A	Halogens	Most reactive nonmetals	7	–1	p	p^5
18 or 8A	Noble Gases	Inert, nonmetal gases	8 (2 for He)	0	p (2 for He)	p^6 (s^2 for He)
3 – 12 or B Groups	Transition Metals Periods 4–7 Final electron enters d sublevel	Less reactive metals; typically hard solids with high melting and boiling points	Varies	Varies	d	d^x x = 1 to 10
	Inner Transition Metals ♦Lanthanide Series ♦Actinide Series	♦Rare earth metals; silvery, high melting points ♦Radioactive elements *Three occur naturally* ♦Transuranium elements (synthetic)	Varies	Varies	f	f^x x = 1 to 14

Valence electrons: electrons in an atom's highest principal energy level that determine the chemical properties and behavior of an element

- Atoms in the same group have similar properties because they have the same number of valence electrons
- Energy level of valence electrons is indicated by the period
- For representative elements, the last digit of the group number indicates the number of valence electrons. Transition elements have different numbers of valence electrons under different conditions.

Element	# Valence Electrons	Energy Level of Valence Electrons	Configuration of Valence Electrons
1. Hydrogen	1	1	$1s^1$
2. Nitrogen	5	2	$2s^2 2p^3$
3. Magnesium	2	3	$3s^2$
4. Silicon	4	3	$3s^2 3p^2$
5. Sulfur	6	3	$3s^2 3p^4$
6. Krypton	8	4	$4s^2 4p^6$

Ions

- Neutral atoms have no overall electrical charge because they have equal numbers of positively charged protons and negatively charged electrons.
- Noble gases have stable configurations because the s and p orbitals of their highest energy level are filled, forming a stable octet.
  - Exception: helium has only two s e^- in its highest energy level (duet)
- Atoms gain or lose electrons to increase stability by attaining electron configurations similar to that of the noble gases.
- Such an atom is no longer neutral but has become a charged particle known as an ion.
  - Metals: lose electrons to become cations (positive ions)
  - Nonmetals: gain electrons to become anions (negative ions)

✐Write the electron configuration for a neutral potassium atom. Circle valence e^-.

$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

✐Write the electron configuration for a potassium ion.

$1s^2 2s^2 2p^6 3s^2 3p^6$

✐Write the electron configuration for a neutral argon atom.

$1s^2 2s^2 2p^6 3s^2 3p^6$

OCTET RULE

Atoms tend to gain, lose, or share electrons to acquire a full set of eight valence electrons. First energy level is complete with only two electrons (duet).

Oxidation number (or **state**): the positive or negative charge of a monatomic ion. It equals the number of electrons transferred when an atom forms its ion.

- Predicted by the group or family of an element
- Positive when electrons are lost; negative when electrons are gained
- Written above group number on periodic table
- Noble gases have oxidation number of 0; they do not transfer electrons and do not form ions.
- Elements in the carbon group have no oxidation number; they do not typically form ions.

✐

Element Name	Element Symbol	Oxidation Number	Symbol for Ion	Configuration of Ion (*noble gas configuration*)
1. Hydrogen	H	+1	H^+	N/A
2. Magnesium	Mg	+2	Mg^{2+}	$1s^2 2s^2 2p^6$
3. Oxygen	O	–2	O^{2-}	$1s^2 2s^2 2p^6$
4. Aluminum	Al	+3	Al^{3+}	$1s^2 2s^2 2p^6$
5. Nitrogen	N	–3	N^{3-}	$1s^2 2s^2 2p^6$
6. Sodium	Na	+1	Na^+	$1s^2 2s^2 2p^6$
7. Sulfur	S	–2	S^{2-}	$1s^2 2s^2 2p^6 3s^2 3p^6$
8. Calcium	Ca	+2	Ca^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6$
9. Chlorine	Cl	–1	Cl^-	$1s^2 2s^2 2p^6 3s^2 3p^6$
10. Bromine	Br	–1	Br^-	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$

✏ Practice

1. Elements #110 through #118 are the most recently discovered elements. For these nine elements, complete the table below.

Atomic #	Element Name	Element Symbol	Discovered in	Group	# ve^-	Most closely resembles
110	Darmstadtium	Ds	1994			
111	Roentgenium	Rg	1994			
112	Copernicium	Cn	1996			
113	Ununtrium	Uut	2003			
114	Flerovium	Fl	1998			
115	Ununpentium	Uup	2004			
116	Livermorium	Lv	2000			
117	Ununseptium	Uus	2009			
118	Ununoctium	Uuo	2002			

1. Explain the term *representative elements*.

   What is the primary difference between the representative elements and the transition elements? *(Think: electron configuration.)*

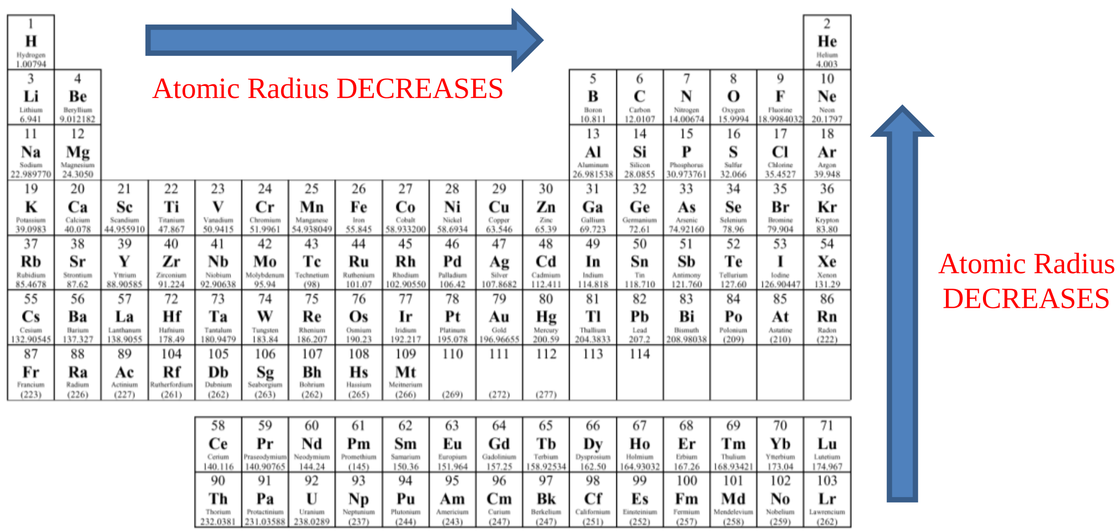
2. The numbers and locations of valence electrons determine the ________________ ________________ of elements.
3. Oxygen is a gas. Sulfur is a solid. What is it about their electron structures that cause them to be grouped in the same chemical family?
4. Identify the element fitting each of the following descriptions:
   a. The metalloid in group 3A: ____________________________________
   b. The halogen in period 5: ______________________________________
   c. The alkali metal in period 4: ___________________________________
   d. The nonmetal that is a liquid at room temperature: ________________
5. Why about zinc, cadmium, and mercury cause them to be in the same chemical family?

CP/Honors Chemistry

PERIODIC TRENDS: ATOMIC RADIUS

Atomic radius: one half the distance between the nuclei of two atoms of the same element that are bonded together

Trends within periods (L to R)	Trends within groups (bottom to top)
Atomic radius DECREASES Why? ↑electrostatic attraction (greater positive charge in nucleus pulls orbitals closer); e^- added to same principal energy level within period (no shielding of valence e^-)	Atomic radius DECREASES Why? As positive nuclear charge increases, e^- are added to higher energy levels, and outer orbital increases in size; valence e^- are further from nucleus and are shielded by inner e^-



Atoms with largest atomic radii: Group 1 and atoms in last period of other groups

Practice

1. Which has the largest atomic radius: lithium (Li), beryllium (Be), nitrogen (N), or neon (Ne)? _______________ The smallest? _______________
2. Which has the largest atomic radius: sodium (Na), potassium (K), rubidium (Rb) or cesium (Cs)? _______________ The smallest? _______________
3. Circle the element in each pair with the largest atomic radius.

   Na or K   Na or Mg   O or F   Br or I   Ti or Zr
4. Rank the following elements from smallest to largest atomic radius: Na, Mg, Cl, K, Rb. ____
5. Which has the largest atomic radius: potassium (K) or magnesium (Mg)?
6. Which has the smallest atomic radius: aluminum (Al) or barium (Ba)?

PERIODIC TRENDS: IONIC RADIUS

Ionic radius: one half the distance between the nuclei of two adjacent ions of the same element

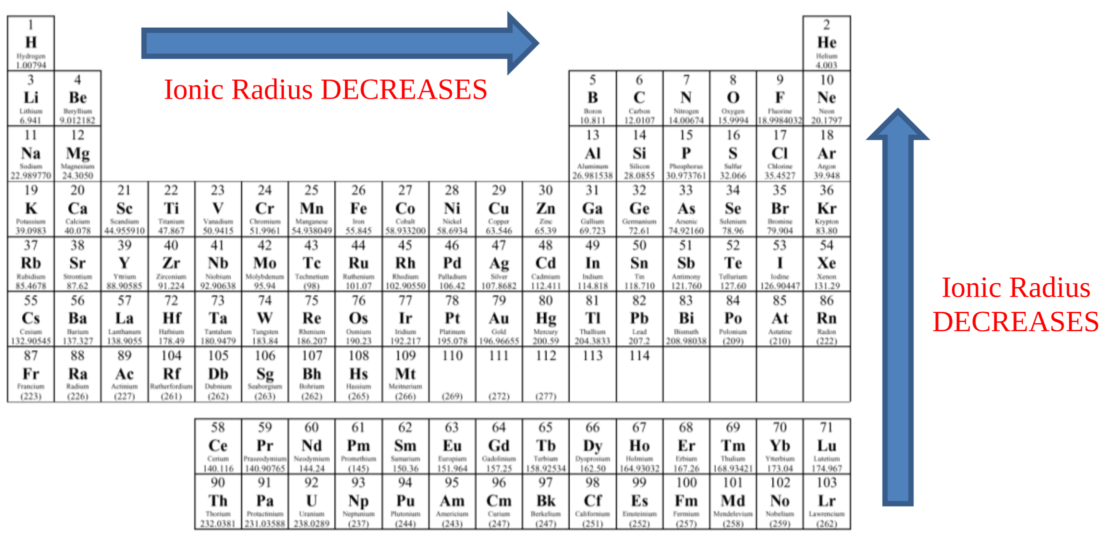
When atoms form cations, they become smaller. Why? Loss of e- decreases number of valence e^- and electrostatic repulsion; more p^+ than e^- holds e^- more tightly to nucleus

When atoms form anions, they become larger. Why? Gain of e-; added e^- increase electrostatic repulsion forcing e^- to spread apart; more e^- than p^+ decreases electrostatic attraction

Trends within periods (L to R)
Ionic radius DECREASES
Anions begin to appear in carbon group. See Figure 6-14, p. 166.

Trends within groups (bottom to top)
Ionic radius DECREASES
When ion's outer electrons are in higher principal energy levels, size of ion increases

Anions are larger than cations



Practice

Underline the one in each pair with the larger radius.

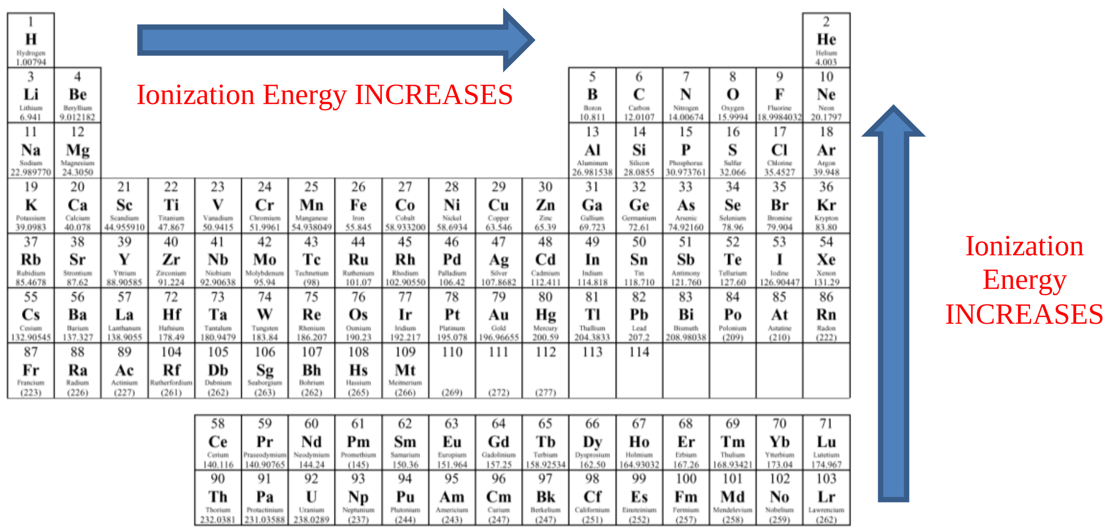
1. a calcium atom or a calcium ion
2. a chlorine atom or a chloride ion
3. a magnesium ion or an aluminum ion
4. a sodium atom or a silicon atom
5. a potassium ion or a bromide ion
6. a potassium atom or a potassium ion
7. a sodium ion or a chloride ion
8. a strontium atom or an iodine atom
9. a rubidium ion or a strontium ion

PERIODIC TRENDS: IONIZATION ENERGY (IE)

Ionization energy: the minimum amount of energy required to remove an electron from a gaseous atom; must overcome the attraction between the positive charge in the nucleus and the negative charge of the electron

First ionization energy: required to remove the first electron from an atom
Second ionization energy: required to remove the second electron from 1+ ion
Third ionization energy: required to remove a third electron from a 2+ ion, *etc*.

Trends within periods (L to R)	Trends within groups (bottom to top)
Ionization energy INCREASES Why? Increased nuclear charge increases hold on valence e^-, making them more difficult to remove	Ionization energy INCREASES Why? Higher the energy level of valence e-, further from nucleus, and less energy is required to remove them



Practice

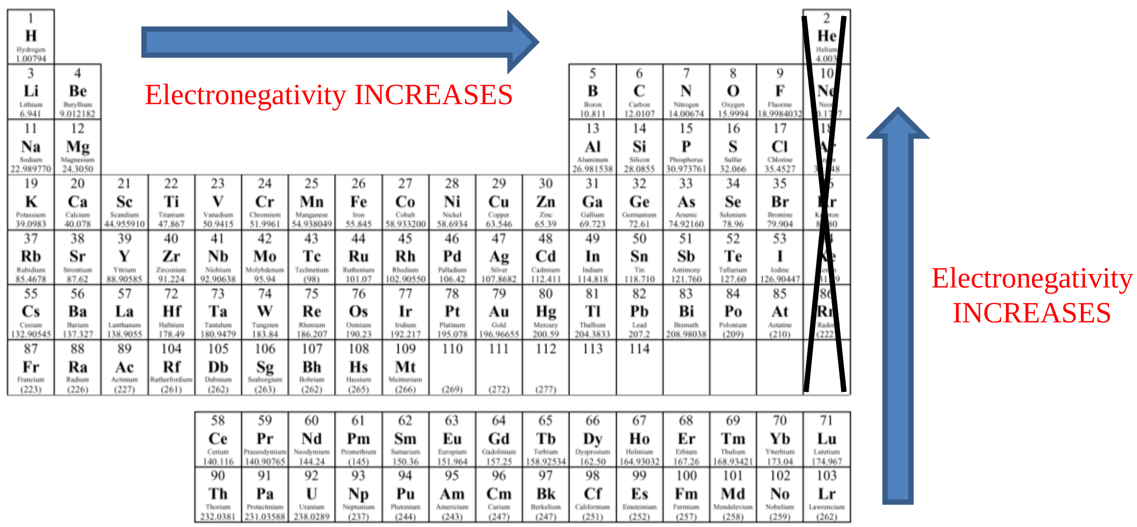
1. Does sodium (Na) or potassium (K) have a higher first ionization energy?
2. Does magnesium (Mg) or argon (Ar) have a higher first ionization energy?
3. Explain why much more ionization energy is required to remove the first electron from neon than from sodium.
4. Why does barium (Ba) have a lower ionization energy than beryllium (Be)?

PERIODIC TRENDS: ELECTRONEGATIVITY

Electronegativity: the ability of atoms to attract electrons in a chemical bond – expressed as number value of 4.0 or less with the unit Paulings

- Differences in electronegativity determine the types of bonds formed
- Electronegativity of noble gases is not usually included

Trends within periods (L to R) Electronegativity INCREASES	Trends within groups (bottom to top) Electronegativity INCREASES



✐Shade the box of the element with the *highest* electronegativity. (fluorine)

✐Outline the box of the element with the *lowest* electronegativity. (francium)

✐ Practice

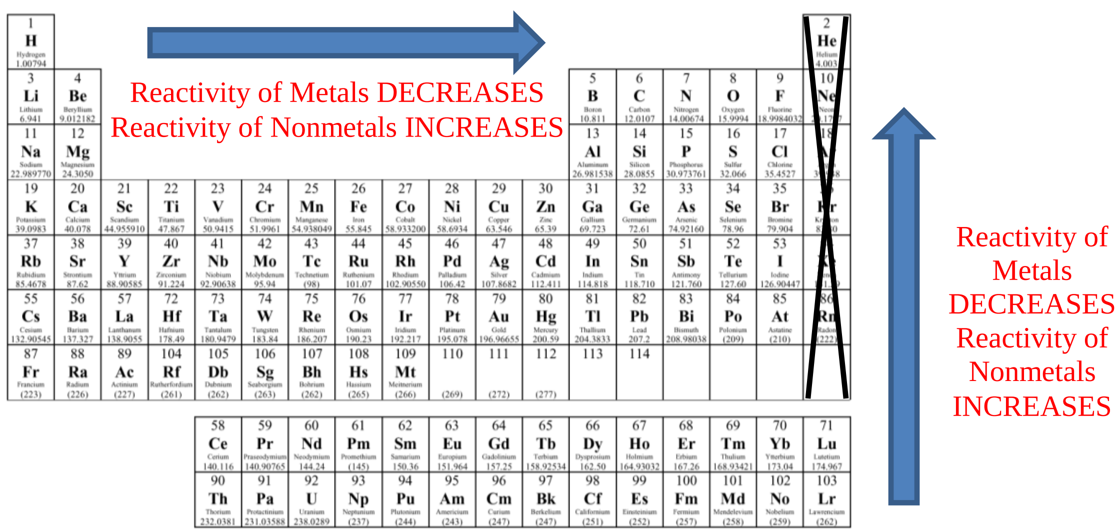
1. Does magnesium or aluminum have a higher electronegativity value?
2. Does nitrogen or phosphorous have a higher electronegativity value?
3. Does calcium (Ca) or bromine (Br) have a higher electronegativity value?
4. Does sodium (Na) or potassium (K) have a higher electronegativity value?
5. Which atom is more electronegative: hydrogen (H) or oxygen (O)?
6. Which atom is more electronegative: carbon (C) or chlorine (Cl)?
7. Which atom is more electronegative: magnesium (Mg) or oxygen (O)?
8. Which atom is more electronegative: sodium (Na) or chlorine (Cl)?

PERIODIC TRENDS: REACTIVITY

Reactivity: refers to how readily chemical substances undergo chemical reaction – related to several factors, including the number of valence electrons, ionization energy, and electronegativity

- Metals: more reactive if have low number of valence electrons and low ionization energies
- Nonmetals: more reactive if have larger numbers of valence electrons and high electronegativity values

Trends within periods (L to R)	Trends within groups (bottom to top)
Metals: Reactivity DECREASES	Metals: Reactivity DECREASES
Nonmetals: Reactivity INCREASES	Nonmetals: Reactivity INCREASES



✎Shade the element box of the most reactive metal: francium

✎Outline the box of the most reactive nonmetal: fluorine

✎ Practice

1. Which metal is more reactive: sodium (Na) or rubidium (Rb)?
2. Which nonmetal is more reactive: oxygen (O) or sulfur (S)?
3. Which element is more reactive: magnesium (Mg) or aluminum (Al)?
4. Which element is more reactive: phosphorous (P) or chlorine (Cl)?