May 16, 2023

Hello,

You are receiving this because you are enrolled in AP Chemistry for the 2023-2024 school year.

A good foundation of first year chemistry concepts is crucial to success in AP Chemistry.

You are expected to have mastered the following BEFORE taking the course.

- Classification of matter
- Certain scientific laws like the Laws of Conservation, Multiple Proportions, and Definite Proportions
- SI units and their prefixes
- Significant digit rules for measurements and calculations
- Dimensional Analysis
- Atomic structure
- Periodic table organization
- Chemical nomenclature
- Calculation of empirical and molecular formulas
- Stoichiometry, including limiting reagent, excess yield, and percent yield

The textbook for the course is Chemistry: The Central Science, 15th edition, by Brown and LeMay. We have not received the new textbooks yet, so I am including access to a PDF file of the current textbook.

Chemistry_The_Central_Science-_13th_Edition_ebook3000.pdf

The AP Chemistry course is designed to be the equivalent of the general chemistry course usually taken during the first college year. The course content is organized into nine units. Your assignment for the summer is to work through Unit 1, Atomic Structure and Properties. While some of this will be a review of material covered in honors chemistry, there will be new content as well. To cover the material in unit 1, you will be reading specified pages from the textbook PDF, completing a notes packet, and watching videos to check the work in your notes packet. I recommend that you work through the packet answering as much as you can on your own first. Then watch the video lessons to check your work and learn about those topics that are new to you.

The notes packet videos are divided into 3 parts. I have listed the relevant textbook pages and links for the topics in each part. Some pages cover more than one topic section.

Part one: Topics 1.1 – 1.3 https://www.youtube.com/watch?v=Wgho7cLgoco

1.1: Moles and molar mass, pages 92-98
1.2: Mass spectroscopy, pages 50-52
1.3: Elemental composition of pure substances, pages 89-101

Part two: Topics 1.4-1.6 https://www.youtube.com/watch?v=Hp1cKGd6PR4

1.4 Composition of mixtures, pages 10-11, 91
1.5: Atomic structure and electron configuration, pages 49, 214-219, 223-224, 230-245
1.6: Photoelectron spectroscopy, pages 217-218

Part three: Topics 1.7-1.8 https://www.youtube.com/watch?v=yXLVe3igS-c

1.7: Periodic Trends, pages 259-272 (groups of the periodic table are considered prior knowledge from honors chemistry)
1.8" Valence electrons and ionic compounds, 239-243, 60-65

We will discuss and practice with the Unit 1 Concepts beginning on the first day of class. If you have any questions, please feel free to email me at ewatson@caschools.us

Polyatomic ions that you really should be familiar with, but I will not be giving a quiz on them.

Per____ate	**_____ate**	**____ite**	**hypo____ite**	**Monatomic anions For REFERENCE**
ClO_4^- perchlorate	ClO_3^- chlorate	ClO_2^- chlorite	ClO^- hypochlorite	Cl^- chloride
BrO_4^- perbromate	BrO_3^- bromate	BrO_2^- bromite	BrO^- hypobromite	Br^- bromide
IO_4^- periodate	IO_3^- iodate	IO_2^- iodite	IO^- hypoiodite	I^- iodide
xxx	NO_3^- nitrate	NO_2^- nitrite	xxx	N^{3-} nitride
xxx	SO_4^{2-} sulfate	SO_3^{2-} sulfite	xxx	S^{2-} sulfide
xxx	PO_4^{3-} phosphate	PO_3^{3-} phosphite	xxx	P^{3-} phosphide
xxx	CO_3^{2-} carbonate	xxx	xxx	C^{4-} carbide
xxx	CrO_4^{2-} chromate	xxx	xxx	xxx

per___ate: has one more oxygen than “ATE”

___ate: most common form

___ite: one less oxygen than “ATE”

hypo__ite: two less oxygens than “ATE”

Video:

https://youtu.be/69ZbHNNcfz0

OTHERS:

MnO_4^-	permanganate	$Cr_2O_7^{2-}$	dichromate
$C_2H_3O_2^-$	Acetate (some times written as CH_3COO^-	OH^-	hydroxide
HCO_3^-	hydrogen carbonate (bicarbonate)	CN^-	cyanide
NH_4^+	ammonium	(note positive charge)	

1.1 Moles and Molar Mass

Essential knowledge statements from the AP Chemistry CED:

- One cannot count particles directly while performing laboratory work. Thus, there must be a connection between the masses of substances reacting and the actual number of particles undergoing chemical changes.
- Avogadro's number ($N_A = 6.02 \times 10^{23}$ mol^{-1}) provides the connection between the number of moles in a pure sample of a substance and the number of constituent particles (or formula units) of that substance.
- Expressing the mass of an individual atom or molecule in atomic mass units (amu) is useful because the average mass in amu of one particle (atom or molecule) or formula unit of a substance will always be numerically equal to the molar mass of that substance in grams. Thus, there is a quantitative connection between the mass of a substance and the number of particles that the substance contains.

The particles of a substance can be described as atoms, molecules, or formula units, as shown in the following examples. The molar mass of a substance can be determined or calculated from the atomic mass values on the periodic table.

1 mol Mg = 24.30 g Mg = 6.02×10^{23} atoms Mg

1 mol CO_2 = 44.01 g CO_2 = 6.02×10^{23} molecules CO_2

1 mol NaCl = 58.44 g NaCl = 6.02×10^{23} formula units NaCl

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1. Calculate the mass, in grams, of 0.0850 mol $Ba(OH)_2$.

2. Calculate the number of moles of C_4H_{10} present in 2.00 g C_4H_{10}.

3. Calculate the number of atoms of Si present in 35.0 mol Si.

4. Calculate the number of moles of O_3 present in 4.3×10^{24} molecules of O_3.

5. Calculate the mass, in grams, of 8.2×10^{22} molecules of $CHCl_3$.

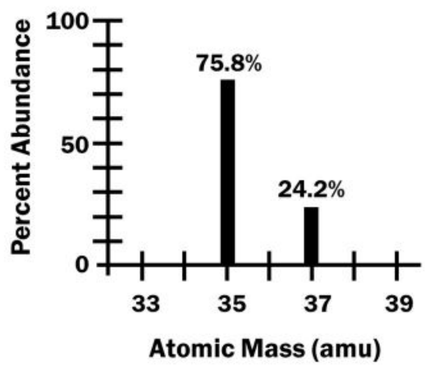
6. Calculate the number of formula units of Na_2SO_4 present in 0.248 g Na_2SO_4.

1.2 Mass Spectroscopy of Elements

Essential knowledge statements from the AP Chemistry CED:

- The mass spectrum of a sample containing a single element can be used to determine the identity of the isotopes of that element and the relative abundance of each isotope in nature.
- The average atomic mass of an element can be estimated from the weighted average of the isotopic masses using the mass of each isotope and its relative abundance.

Isotope	Abundance
Cl-35	75.8%
Cl-37	24.2%

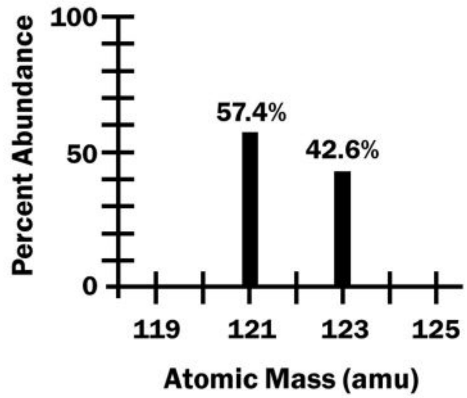


7. Based on the information shown above,

(a) calculate the average atomic mass of Cl.

(b) Fill in the table below.

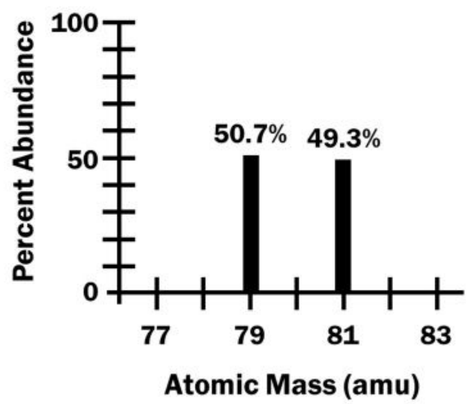
Isotope	Protons	Neutrons
Cl-35		
Cl-37		



8. Based on the information shown above,

   (a) calculate the average atomic mass of the element.

   (b) What is the most likely identity of this element? ______________________



9. Based on the information shown above,

   (a) what is the most likely identity of this element? ______________________

   (b) Fill in the table below.

Mass Number	Protons	Neutrons
79		
81		

10. A certain element has two naturally occurring isotopes with mass numbers of 63 and 65.

(a) What is the most likely identity of this element? ________________________________

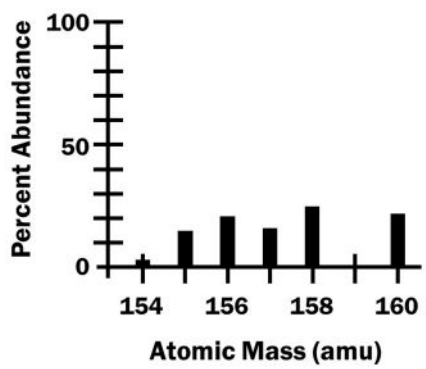
(b) Fill in the table below.

Mass Number	Protons	Neutrons
63		
65		

(c) Which isotope of this element, mass number = 63 or mass number = 65, is more abundant in nature? Justify your answer.

11. If an element has several naturally occurring isotopes, the calculation of the average atomic mass of the element can be a bit more complicated.

Mass Number	Abundance
154	2.18%
155	14.80%
156	20.47%
157	15.65%
158	24.84%
160	22.06%



(a) Based on the information above, estimate the average atomic mass of the element to the nearest whole number. Then use a calculator to determine the average atomic mass.

(b) What is the most likely identity of this element? ________________________________

1.3 Elemental Composition of Pure Substances

Essential knowledge statements from the AP Chemistry CED:

- Some pure substances are composed of individual molecules, while others consist of atoms or ions held together in fixed proportions as described by a formula unit.
- According to the law of definite proportions, the ratio of the masses of the constituent elements in any pure sample of that compound is always the same.
- The chemical formula that lists the lowest whole number ratio of atoms of the elements in a compound is the empirical formula.

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12. Calculate the percent composition by mass of each element in glucose ($C_6H_{12}O_6$).

13. Calculate the percent composition by mass of each element in erythrose ($C_4H_8O_4$).

14. What is the empirical formula of glucose? ____________________

    What is the empirical formula of erythrose? ___________________

Two different compounds with the same empirical formula have the same percent composition by mass.

15. A certain compound has the following percent composition by mass.

    43.64% P   56.36% O

    Determine the empirical formula of this compound.

If you are given mass data for a certain compound, the following procedure will help you to determine the empirical formula of the compound.

- Convert the mass of each element into moles.
- Divide each value of moles by the lowest number.
- At this point, you may already have whole numbers for the moles of each element. If not, then you may need to multiply each value by 2 or by 3 in order to get whole numbers.
- Use the whole number values of moles to write the empirical formula.

16. A certain compound has the following percent composition by mass.

52.14% C   13.13% H   34.73% O

Determine the empirical formula of this compound.

17. A pure sample of tin (Sn) with a mass of 6.18 g is burned in air until the tin is completely converted into tin oxide. The mass of the tin oxide is equal to 7.85 g. Determine the empirical formula of the tin oxide compound.

18. Compound X consists of the elements C, H, and N. A 15.00-g sample of compound X contains 9.81 g C, 1.37 g H, and 3.82 g N.

(a) Determine the empirical formula of compound X.

(b) It is determined that a 25.0-gram sample of compound X contains 9.11×10^{22} molecules. Calculate the molar mass of compound X, in units of g/mol.

18. (c) Based on your answers to parts (a) and (b), determine the molecular formula of compound X.

Another way to determine the empirical formula of a compound is to use data from a combustion experiment. If a compound that contains carbon and hydrogen is burned in the presence of excess oxygen gas, the carbon will be converted into CO_2 and the hydrogen will be converted into H_2O. If the compound contains other elements such as nitrogen or sulfur, other gases may be formed.

Mass of sample that is burned	5.00 g
Mass of CO_2 produced	10.99 g
Mass of H_2O produced	6.00 g

19. A sample of a compound that contains carbon, hydrogen, and oxygen is burned completely in O_2. Data from the combustion experiment is shown in the table above.

    (a) Determine the mass of carbon (C) present in 5.00 g of the compound.

    (b) Determine the mass of hydrogen (H) present in 5.00 g of the compound.

    (c) Determine the mass of oxygen (O) present in 5.00 g of the compound.

    (d) Determine the empirical formula of the compound.

Another type of situation that involves mass and mole ratios involves a substance known as a hydrate. A hydrate is a substance in which water molecules are included in the chemical formula. These substances are often ionic compounds in which water molecules are bonded to the ions in the crystal structure. A hydrate salt can be heated to remove the water through evaporation, forming an anhydrous salt. Two examples of anhydrous salts and hydrates are listed in the table below.

Anhydrous Salt	Hydrate Salt
copper(II) sulfate, $CuSO_4$	copper(II) sulfate pentahydrate, $CuSO_4 \bullet 5H_2O$
calcium chloride, $CaCl_2$	calcium chloride dihydrate, $CaCl_2 \bullet 2H_2O$

20. A sample of $CuSO_4 \bullet 5H_2O$ has a mass of 25.00 g.

(a) Calculate the mass of $CuSO_4$ present in this 25.00-g sample.

(b) Calculate the mass of H_2O present in this 25.00-g sample.

21. Calculate the percentage of H_2O by mass in $CaCl_2 \bullet 2H_2O$.

22. In a certain experiment, a sample of a hydrate of magnesium sulfate, $MgSO_4 \bullet nH_2O$, is heated in order to remove all of the water from the sample. Experimental data is shown in the table below.

mass of empty container	25.356 g
mass of container and hydrate salt, before heating	28.418 g
mass of container and sample after 1st heating	26.931 g
mass of container and sample after 2nd heating	26.853 g
mass of container and sample after 3rd heating	26.852 g

(a) Explain how the data indicates that all of the water has been removed from the hydrate salt in this experiment.

22. (b) Calculate the mass of the hydrate salt used in this experiment.

(c) Calculate the mass of water that was removed from the hydrate sample in this experiment.

(d) Determine the value of n in the formula $MgSO_4 \bullet nH_2O$.

1.4 Composition of Mixtures

Essential knowledge statements from the AP Chemistry CED:

- While pure substances contain molecules or formula units of a single type, mixtures contain molecules or formula units of two or more types, whose relative proportions can vary.
- Elemental analysis can be used to determine the relative numbers of atoms in a substance and to determine its purity.

Mass of NaCl	Mass of $MgCl_2$	Total Mass of Mixture
2.75 g	3.42 g	6.17 g

1. Answer the following questions about the mixture whose composition is listed in the table above.

   (a) Calculate the percentage of NaCl by mass in this mixture.

   (b) Calculate the percentage of Na by mass in this mixture.

   (c) Calculate the percentage of Cl by mass in this mixture.

2. A sample of a solid labeled as $AgNO_3$ may be impure. A student analyzes the sample, and determines that it contains 68% Ag by mass.

   (a) Calculate the percentage of Ag by mass in a pure sample of $AgNO_3$.

2. (b) Which of the following is more likely to represent the solid sample that was analyzed by the student? Justify your answer.

a mixture of $AgNO_3$ and $AgCl$	a mixture of $AgNO_3$ and $AgBr$

3. A student needs to analyze a mixture that contains $BaCl_2$ and NaCl. The student dissolves a 6.75-g sample of this mixture completely into water and adds an excess amount of $Na_2SO_4(aq)$. A white precipitate of $BaSO_4(s)$ is formed, based on the following chemical equation.

$$BaCl_2(aq) + Na_2SO_4(aq) \rightarrow 2\ NaCl(aq) + BaSO_4(s)$$

The solid precipitate is filtered, dried, and weighed, and its mass is recorded as 2.36 g.

(a) Calculate the number of moles of $BaSO_4(s)$ that is recovered in this experiment.

(b) Calculate the percentage of $BaCl_2$ by mass in this mixture.

4. A mixture of $CaCO_3$ and Na_2CO_3 is found to contain 35.00% Na by mass. Calculate the percentage of Na_2CO_3 by mass in this mixture.

1.5 Atomic Structure and Electron Configuration

Essential knowledge statements from the AP Chemistry CED:

- The atom is composed of negatively charged electrons and a positively charged nucleus that is made of protons and neutrons.
- Coulomb's law is used to calculate the force between two charged particles.

$$F_{coulombic} \propto \frac{q_1 q_2}{r^2}$$

- In atoms and ions, the electrons can be thought of as being in "shells (energy levels)" and "subshells (sublevels)," as described by the electron configuration. Inner electrons are called core electrons, and outer electrons are called valence electrons. The electron configuration is explained by quantum mechanics, as delineated in the Aufbau principle and exemplified in the periodic table of the elements.
- The relative energy required to remove an electron from different subshells of an atom or ion or from the same subshell in different atoms or ions (ionization energy) can be estimated through a qualitative application of Coulomb's law. This energy is related to the distance from the nucleus and the effective (shield) charge of the nucleus.

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5. The valence electrons of both Na and Mg are located in the 3rd energy level. Which atom, Na or Mg, experiences a greater attractive force between the nucleus and the valence electrons? Justify your answer in terms of Coulomb's law.

6. The valence electron of Na is located in the 3rd energy level, whereas the valence electron of K is located in the 4th energy level. Which atom, Na or K, experiences a greater attractive force between the nucleus and the valence electron? Justify your answer in terms of Coulomb's law.

7. Ionization energy is normally expressed in units of kilojoules per mole, and is defined as the energy required to remove one mole of electrons from one mole of gaseous atoms (or ions) in their ground states. This process is represented by the equation below.

$$\mathrm{X}(g) \;+\; \text{ionization energy} \;\rightarrow\; \mathrm{X}^{+}(g) \;+\; e^{-}$$

Based on your answers to Questions #5 and #6, arrange the atoms Na, Mg, and K in order of increasing ionization energy value.

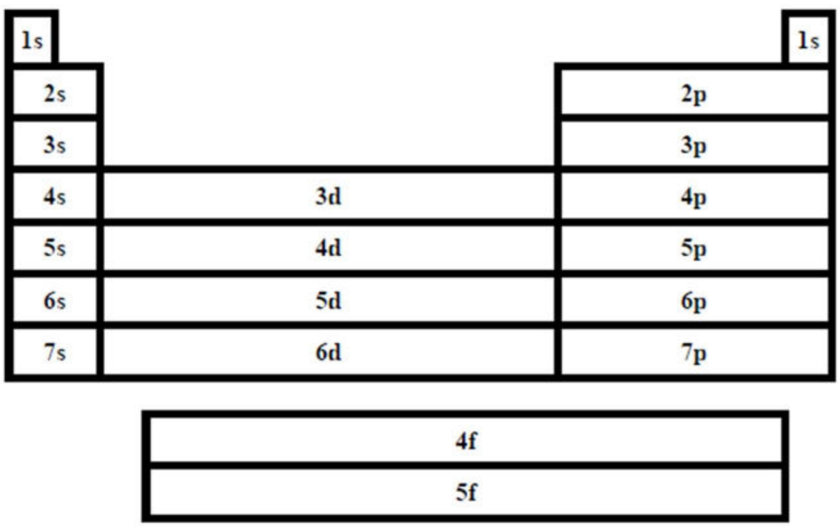
lowest ionization energy value	- - - - - - - ->	highest ionization energy value

The Bohr Model of the Hydrogen Atom (1913)

- Electrons travel in orbits around the nucleus. Only orbits of certain radii, corresponding to certain specific energy values, are permitted for the electron.
- An electron absorbs energy when it moves farther away from the nucleus from a lower energy level to a higher energy level.
- An electron releases energy when it moves closer to the nucleus from a higher energy level to a lower energy level.
- The letter "n" refers to the principal quantum number or the electronic energy level. The lowest energy level (n = 1) for a hydrogen atom is called the ground state. The higher energy levels (n = 2 or higher) are called excited states.

The Bohr model of the hydrogen atom is a primitive, inaccurate model. Today we do not think of electrons as moving in orbits around the nucleus. Instead, we use the term atomic orbital, which is a mathematical function used to indicate the probability of finding an electron. We can visualize atomic orbitals as "electron clouds."

The **electron configuration** is the distribution of the electrons in an atom or an ion among the various orbitals. There are patterns on the periodic table that help you write the electron configuration of an atom or an ion.



8. Fill in the missing information in the table below.

Element Symbol	Atomic Number	Complete Electron Configuration	Noble Gas Abbreviated Electron Configuration
O	8	$1s^2 2s^2 2p^4$	[He] $2s^2 2p^4$
			[Ne] $3s^2 3p^1$
Ca		$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	
	26		
As			
Cd			

An orbital diagram is another way to represent the electron configuration. Each box represents an orbital. Each electron is represented by an arrow. Electrons that have opposite spins are represented by a pair of arrows pointing in opposite directions. Electrons are paired when they occupy the same orbital. An unpaired electron is an electron in an orbital without another electron of opposite spin.

9. Fill in the missing information in the table below.

Element Symbol	Atomic Number	Orbital Diagram for the Electron Configuration
Be	4	$\boxed{\uparrow\downarrow}$ $\boxed{\uparrow\downarrow}$ $\boxed{}\boxed{}\boxed{}$ $\boxed{}$ $1s$ $2s$ $2p$ $3s$
N	7	$\boxed{}$ $\boxed{}$ $\boxed{}\boxed{}\boxed{}$ $\boxed{}$ $1s$ $2s$ $2p$ $3s$
O	8	$\boxed{}$ $\boxed{}$ $\boxed{}\boxed{}\boxed{}$ $\boxed{}$ $1s$ $2s$ $2p$ $3s$
Na	11	$\boxed{}$ $\boxed{}$ $\boxed{}\boxed{}\boxed{}$ $\boxed{}$ $1s$ $2s$ $2p$ $3s$

The **ground state** electron configuration refers to the arrangement of the electrons in the lowest available energy levels. An **excited state** electron configuration refers to a situation in which at least one of the electrons has moved up to a higher energy level.

10. Circle all of the following that represent an excited state electron configuration.

$1s^2 2s^2 2p^6 3s^2 3p^1$ $\qquad$ $1s^2 2s^2 2p^5 3s^1$ $\qquad$ $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

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

Electron Configurations of Ions

- When electrons are removed from an atom to form a cation, they are always removed first from the occupied orbitals having the largest principal quantum number n (energy level).
- When electrons are added to an atom to form an anion, they are added to the empty or partially filled orbital that has the lowest value of n.
- When an atom of a transition metal (e.g., elements #21 – #30 and #39 – #48) loses electrons to become a cation, *the electrons are first removed from the valence s orbitals*. If additional electrons are lost, they are removed from the valence *d* orbitals.

11. Write the ground state electron configuration for each of the following ions.

Ca^{2+} ________________________________ Fe^{2+} ____________________________________

O^{2-} __________________________________ Fe^{3+} ____________________________________

1.6 Photoelectron Spectroscopy

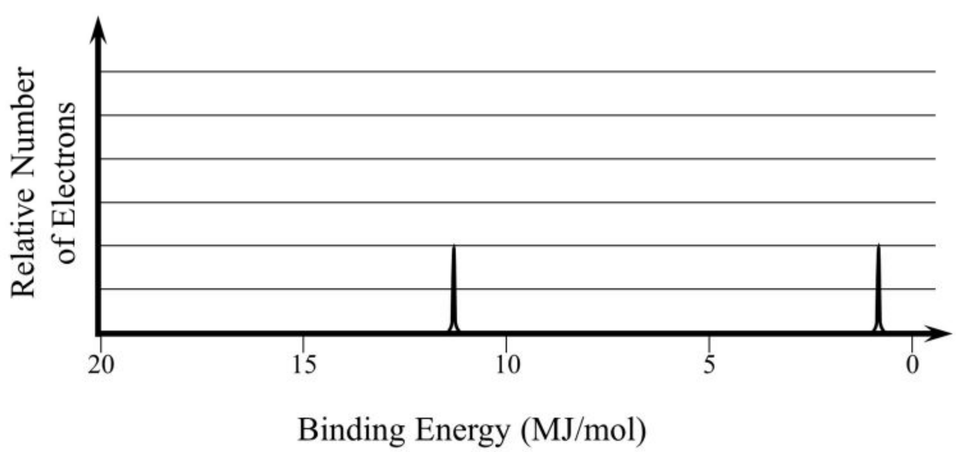
Essential knowledge statements from the AP Chemistry CED:

- The energies of the electrons in a given shell can be measured experimentally with photoelectron spectroscopy (PES). The position of each peak in the PES spectrum is related to the energy required to remove an electron from the corresponding subshell, and the height of each peak is (ideally) proportional to the number of electrons in that subshell.

Photoelectron spectroscopy (PES) is an experimental technique that involves the ionization of a sample by using a high-energy source of light (usually ultraviolet or X-ray). The energy is absorbed by the sample, causing all of the electrons to be removed from the atom. We can use PES to determine the following information.

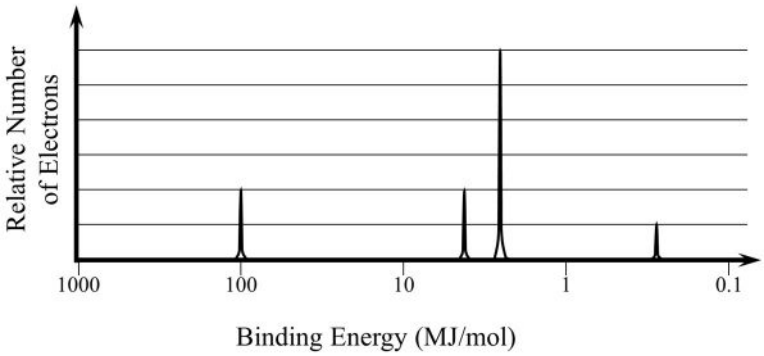
- The binding energy for each subshell
- The number of electrons in each subshell

The relative number of electrons is shown on the y-axis. The binding energy is shown on the x-axis. The appearance of the x-axis for a typical photoelectron spectrum looks a little strange at first. It is sometimes presented as a logarithmic scale. The highest binding energy values are located on the left, and the lowest binding energy values are located on the right. An example of a photoelectron spectrum (PES) for a pure sample of an element is shown below.



12. On the PES diagram above, there are two peaks. Draw a circle around the peak that represents the electrons that are located closer to the atomic nucleus. Justify your answer in terms of Coulomb's law.

The binding energy value in a PES diagram represents the energy required to remove the electrons from a particular subshell. Coulomb's law tells us that the electrons that are located closer to the nucleus should have a stronger attractive force to the nucleus. Therefore the core electrons that are located closer to the nucleus should have a higher binding energy (i.e., require more energy to remove) than the valence electrons in the outermost shell.

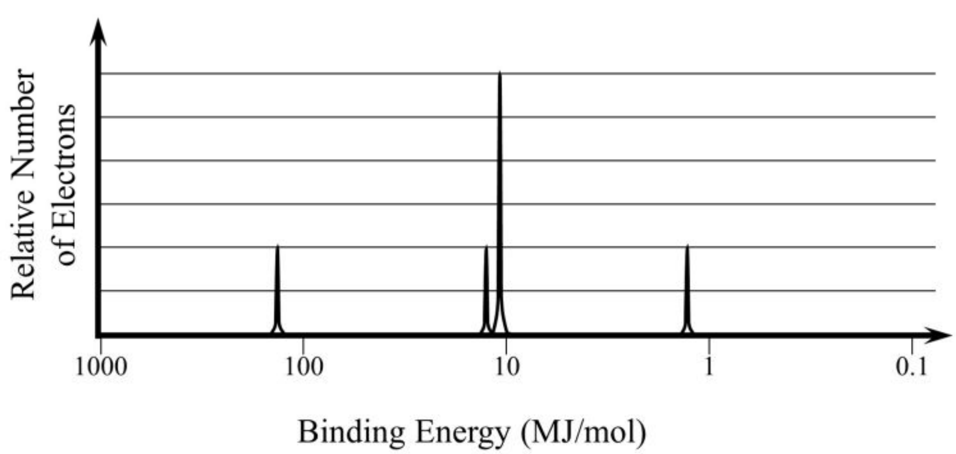


Identity of Element:

13. On the PES diagram above, label each peak as one of the following: 1*s*, 2*s*, 2*p*, or 3*s*. Identify the element that is represented by this PES diagram.

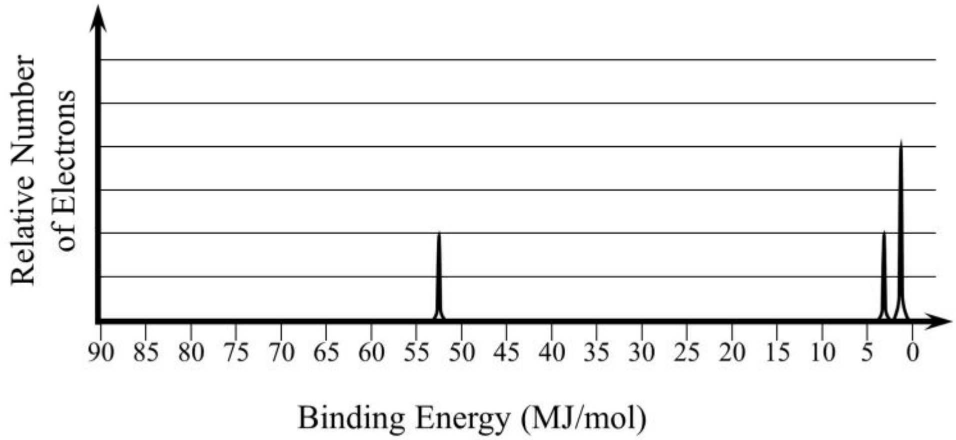
	Binding Energy (MJ/mol)
1*s* electrons in nitrogen (N)	39.6
1*s* electrons in oxygen (O)	52.6

14. The table above shows the binding energy for the 1*s* electrons in a nitrogen atom and the binding energy for the 1*s* electrons in an oxygen atom. Explain the difference in these two values in terms of Coulomb's law and atomic structure.

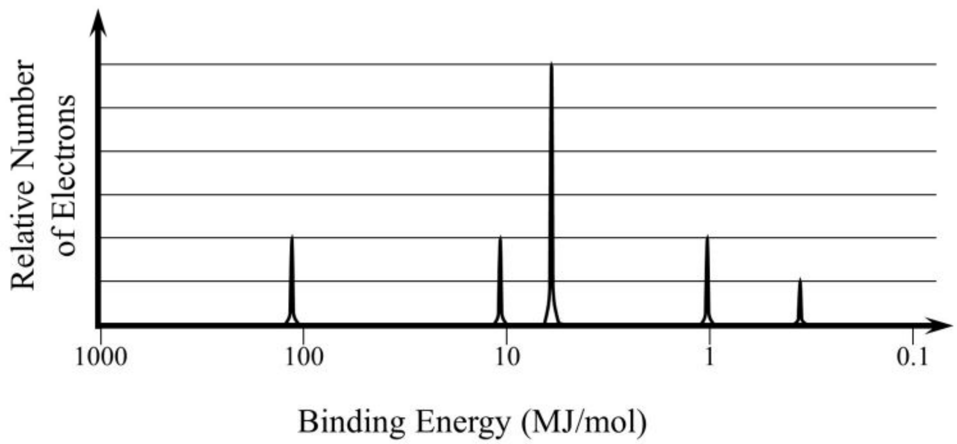


15. A partial photoelectron spectrum of pure phosphorus (P) is shown above. On the spectrum above, draw the missing peak that corresponds to the electrons in the 3*p* sublevel.

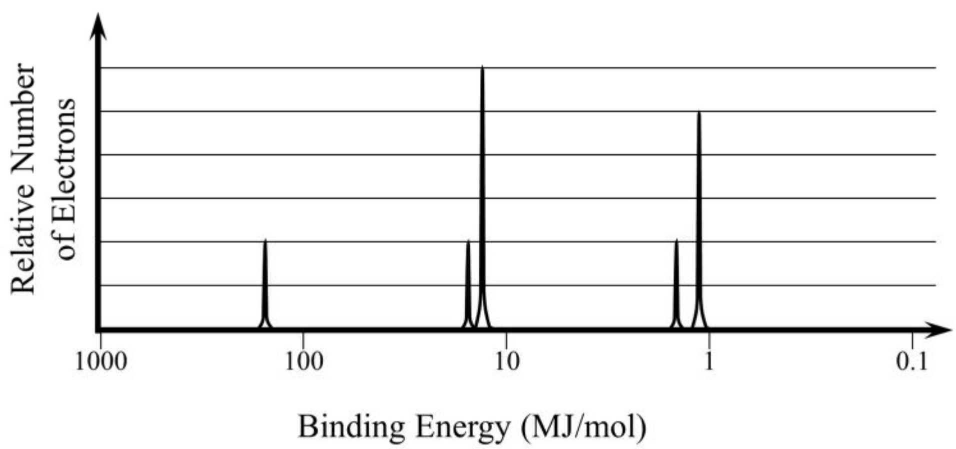
16. The photoelectron spectrum diagrams for three different elements are shown below. Identify the element that is represented by each diagram.



Identity of Element:



Identity of Element:



Identity of Element:

1.7 Periodic Trends

Essential knowledge statements from the AP Chemistry CED:

- The organization of the periodic table is based on the recurring properties of the elements and explained by the pattern of electron configurations and the presence of completely or partially filled shells (and subshells) of electrons in atoms.
- Trends in atomic properties within the periodic table (periodicity) can be qualitatively understood through the position of the element in the periodic table, Coulomb's law, the shell model, and the concept of shielding/effective nuclear charge. These properties include the following.
  - atomic and ionic radii
  - ionization energy
  - electronegativity
  - electron affinity
- Periodicity is useful to predict /estimate values of properties in the absence of data.

Coulomb's law describes the force between two charged particles. This equation is useful when studying periodic trends.

$$F_{coulombic} \propto \frac{q_1 q_2}{r^2}$$

When comparing the atoms of two different elements that are located in the same period,

- The valence electrons of each atom are located in the same energy level.
- The element with more protons has a greater nuclear charge, and there is a stronger attraction between the nucleus and the valence electrons.
- According to Coulomb's law, the greater the magnitude of charge, the stronger the attractive force between oppositely charged particles.

When comparing the atoms of two different elements that are located in the same group,

- The valence electrons of each atom are located in different energy levels.
- Electrons located in a higher energy level are farther away from the nucleus.
- Electrons located in a lower energy level are closer to the nucleus.
- According to Coulomb's law, the smaller the distance between oppositely charged particles, the greater the attractive force between them.

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1. Which element, Li or Be, has a smaller atomic radius? Justify your answer in terms of atomic structure and Coulomb's law.

2. Which element, Li or Na, has a smaller atomic radius? Justify your answer in terms of atomic structure and Coulomb's law.

3. Based on your answers to Questions #1 and #2, arrange the atoms Li, Be, and Na in order of increasing atomic radius.

smallest atomic radius	- - - - - - - ->	largest atomic radius

4. The atomic radius of the Na atom is different than the ionic radius of the Na^+ ion.

(a) Write the complete ground state electron configuration for Na and for Na^+.

Na ________________________________ Na^+ ________________________________

(b) Which particle, Na or Na^+, has a larger radius? Justify your answer in terms of atomic structure.

Ion	Ionic Radius (pm)
Fe^{2+}	92
Fe^{3+}	79

5. The ionic radii of two different ions are shown in the table above.

(a) Write the ground state electron configuration for Fe^{2+} and for Fe^{3+}.

Fe^{2+} ___________________________________ Fe^{3+} ___________________________________

5. (b) In terms of atomic structure, explain why the ionic radius of Fe^{2+} is larger than that of Fe^{3+}.

6. The atomic radius of the F atom is different than the ionic radius of the F^- ion.

(a) Write the complete ground state electron configuration for F and for F^-.

F ________________________________ F^- ________________________________

(b) Which particle, F or F^-, has a larger radius? Justify your answer in terms of atomic structure.

K^+	Ca^{2+}	S^{2-}	Cl^-

7. Each of the ions shown in the table above are members of an isoelectronic series. This means that each ion has the same number of electrons.

(a) Arrange these ions in order of increasing ionic radius.

smallest ionic radius	- - - - - - - ->	- - - - - - - ->	largest ionic radius

(b) Justify your answer.

Ionization energy is normally expressed in units of kilojoules per mole, and is defined as the energy required to remove one mole of electrons from one mole of gaseous atoms (or ions) in their ground states. Removing the outermost electron from a neutral atom is called the *first ionization energy* (IE_1). Removing the outermost electron from a +1 ion is called the *second ionization energy* (IE_2), etc.

$$Mg(g) \rightarrow Mg^{+}(g) + e^{-} \qquad IE_1 = 738 \text{ kJ/mol}$$

$$Mg^{+}(g) \rightarrow Mg^{2+}(g) + e^{-} \qquad IE_2 = 1451 \text{ kJ/mol}$$

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^{-} \qquad IE_3 = 7733 \text{ kJ/mol}$$

8. As you move from left to right across a horizontal row (period) on the periodic table,

   atomic radius values tend to ______________________________ from left to right, and

   first ionization energy values tend to ______________________________ from left to right.

9. As you move from top to bottom down a vertical column (group) on the periodic table,

   atomic radius values tend to ______________________________ from top to bottom, and

   first ionization energy values tend to ______________________________ from top to bottom.

On the AP Exam,

- you will NOT earn credit for simply referring to the relative position of the elements on the periodic table without an explanation.
- you will NOT earn credit for using one trend to explain another trend.

Explain why the first ionization energy value of Mg (738 kJ/mol) is greater than the first ionization energy value of Na (496 kJ/mol).	
Ionization energy increases from left to right across a period. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Unacceptable response because there is no explanation.
Mg has a smaller atomic radius than Na. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Unacceptable response because it uses one trend to explain another trend.
The valence electrons in Na and Mg are located in the same energy level (n = 3). Na has 11 protons, and Mg has 12 protons. Since Mg has a greater nuclear charge than Na, there is a stronger attraction between the nucleus and the valence electrons. Therefore it requires more energy to remove a valence electron from a Mg atom than it does to remove a valence electron from a Na atom.	Acceptable response because it uses principles of atomic structure to explain the data.

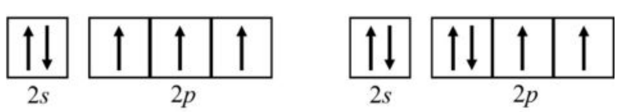
Explain why the first ionization energy value of K (419 kJ/mol) is less than the first ionization energy value of Na (496 kJ/mol).	
Ionization energy decreases from top to bottom down a group. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Unacceptable response because there is no explanation.
K has a larger atomic radius than Na. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Unacceptable response because it uses one trend to explain another trend.
Na has three occupied energy shells, and K has four occupied energy shells. The valence electron in Na is located in a 3s orbital, whereas the valence electron in K is located in a 4s orbital. Since the valence electron in K is farther away from the nucleus than the valence electron in Na, there is a weaker attraction between the nucleus and the valence electron. Therefore it requires less energy to remove a valence electron from a K atom than it does to remove a valence electron from a Na atom.	Acceptable response because it uses principles of atomic structure to explain the data.

Two Anomalies in the Horizontal Trend for First Ionization Energy

Element	Li	Be	B	C
Electron Configuration	$1s^2 2s^1$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^2$
Ionization Energy (kJ/mol)	520	899	801	1086

Although B has one more proton than Be, the ionization energy of B is slightly less than that of Be. This decrease in ionization energy can be explained as follows. The outermost electron for B is located in the 2*p* subshell, whereas the outermost electron for Be is located in the 2*s* subshell. The 2*p* subshell is slightly higher in energy than the 2*s* subshell. It requires slightly less energy to remove an electron from the 2*p* subshell than it does to remove an electron from the 2*s* subshell.

Element	C	N	O	F
Electron Configuration	$1s^2 2s^2 2p^2$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^4$	$1s^2 2s^2 2p^5$
Ionization Energy (kJ/mol)	1086	1402	1314	1681



Nitrogen Oxygen

Although O has one more proton than N, the ionization energy of O is slightly less than that of N. This decrease in ionization energy can be explained as follows. There is slightly more electron-electron repulsion between the paired electrons in the p^4 configuration of O as compared to the p^3 configuration of N. This electron repulsion in the p^4 configuration explains why it requires slightly less energy to remove an electron from an atom of O than it does to remove an electron from an atom of N.

Element	1^{st} IE	2^{nd} IE	3^{rd} IE	4^{th} IE	5^{th} IE	6^{th} IE	7^{th} IE
Na	496	4562	6910	9543	13,354	16,613	20,117
Mg	738	1451	7733	10,543	13,630	18,020	21,711
Al	578	1817	2745	11,577	14,842	18,379	23,326
Si	786	1577	3232	4356	16,091	19.805	23,780
P	1012	1907	2914	4964	6274	21,267	25,431
S	1000	2252	3357	4556	7004	8496	27,107
Cl	1251	2298	3822	5159	6542	9362	11,018

10. Consider the data for successive ionization energy in the table above.

(a) In terms of atomic structure and Coulomb's law, explain why the ionization energy values increase as successive electrons are removed from an atom.

(b) In terms of atomic structure and Coulomb's law, explain why the 2^{nd} IE for Na is much higher than the 2^{nd} IE for Mg.

Element	1^{st} IE	2^{nd} IE	3^{rd} IE	4^{th} IE	5^{th} IE
X	1087	2353	4621	6223	37,831

11. Based on the information in the table above, how many valence electrons does element X have? Justify your answer.

Electronegativity is defined as the tendency of an atom to attract electrons to itself in a chemical bond. The higher the electronegativity value is, the greater the attraction for electrons. Electronegativity values are used when determining if a particular chemical bond is classified as nonpolar covalent, polar covalent, or ionic. The greater the difference in electronegativity between two atoms, the more polar the bond is. Suppose that a polar covalent bond is formed between two atoms X and Y as shown below.

$$\overset{\delta+}{\mathrm{X}}\!-\!\overset{\delta-}{\mathrm{Y}}$$

If atom X is less electronegative than atom Y, there is a partial positive charge (δ+) on atom X and a partial negative charge (δ–) on atom Y. The arrow above the polar covalent bond represents the dipole, which is generated whenever two electrical charges of opposite sign are separated by a distance. The arrow always points toward the atom that has the higher electronegativity value. The measure of the magnitude of the dipole is called the dipole moment. In general, the greater the difference in electronegativity, the greater the magnitude of the dipole moment.

Electronegativity Values

H 2.1						
Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
K 0.8	Ca 1.0	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8
Rb 0.8	Sr 1.0	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5

Notice that the noble gases (He, Ne, Ar, etc.) are not included in the data table above. This is because the atoms of the noble gases ordinarily do not form chemical bonds or share electrons with other atoms.

12. As you move from left to right across a horizontal row (period) on the periodic table,

    electronegativity values tend to ______________________________ from left to right.

    As you move from top to bottom down a vertical column (group) on the periodic table,

    electronegativity values tend to ______________________________ from top to bottom.

13. The smaller the atomic radius is, the __________________________ the electronegativity value is.

    The larger the atomic radius is, the ___________________________ the electronegativity value is.

    The most electronegative element on the periodic table is _______________________.

Electron affinity is a periodic trend is a bit confusing to understand. Electron affinity is defined as the energy change that occurs when an electron is added to a gaseous atom to form a negatively charged anion. Consider the following examples.

$$F(g) + e^- \rightarrow F^-(g) \qquad \Delta E = -328 \text{ kJ/mol}$$

$$Li(g) + e^- \rightarrow Li^-(g) \qquad \Delta E = -60 \text{ kJ/mol}$$

If ΔE is negative, energy is released. If ΔE is positive, energy is absorbed. The greater the attraction is between an atom and an added electron, the more negative the value of ΔE is. The more negative the value of ΔE is, the greater the electron affinity is. As you can see in the table below, the trends in electron affinity are not necessarily clear and predictable.

In general, more energy is released when a nonmetal atom gains an electron than when a metal atom gains an electron. For the noble gases, the electron affinity has a positive value. This indicates that the $X^-(g)$ ion is less stable than the $X(g)$ atom.

Electron Affinity (kJ/mol)

H –73							He +48
Li –60	Be +48	B –27	C –122	N +7	O –141	F –328	Ne +116
Na –53	Mg +40	Al –42	Si –134	P –72	S –200	Cl –349	Ar +96
K –48	Ca –2	Ga –29	Ge –119	As –78	Se –195	Br –325	Kr +96

1.8 Valence Electrons and Ionic Compounds

Essential knowledge statements from the AP Chemistry CED:

- The likelihood that two elements will form a chemical bond is determined by the interactions between the valence electrons and nuclei of elements.
- Elements in the same column of the periodic table tend to form analogous compounds.
- Typical charges of atoms in ionic compounds are governed by their location on the periodic table and the number of valence electrons.

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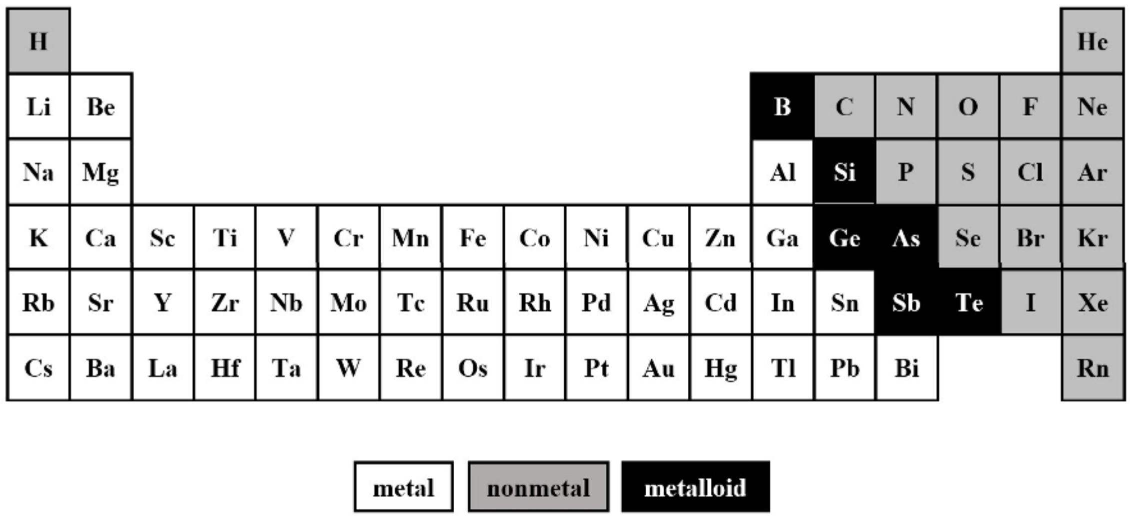
14. Write the correct number of valence electrons for each of the following elements.

Element	Li	Be	B	C	N	O	F	Ne
Valence Electrons								

15. Write the correct charge (e.g., 1+, 2+, 1–, 2–, etc.) that each of the following elements has when it forms a stable monoatomic ion.

Element	Li	Be	B	C	N	O	F	Ne
Charge				N/A				



Binary ionic compounds (e.g., NaCl) normally consist of a metal and a nonmetal. The chemical formula of a binary ionic compound can be determined be examining the charges on each ion. The formula is written as an empirical formula, and should have an overall charge of zero.

16. Write the correct chemical formula for the binary ionic compound that is formed from the combination of each of the following pairs of elements.

Elements	Chemical Formula of the Binary Ionic Compound
Li and F	
Na and S	
Mg and Cl	
Al and O	
Ca and P	

Elements in the same group (column) of the periodic table have the same number of valence electrons. This explains why elements in the same group tend to form analogous compounds.