LWHS AP Chemistry Summer Assignment 2026-2027

Welcome to AP Chemistry! I am looking forward to having you join our class as a chemist for the 2026-2027 school year. If you have not taken a chemistry course in the past, please be aware that AP Chemistry is equivalent to *one year* of college chemistry and is designed as a second-year course to follow the foundational class, Chemistry in the Earth Systems (CITES). If you are excited about the challenge of taking AP Chemistry as a first-year student, success is possible with dedication and hard work! This will involve extra effort and a strong commitment to mastering the material, but with perseverance and enthusiasm, it can be a rewarding experience. No matter your previous chemistry experience, this class will push you to do your best every day and work with your classmates as a team with the common goal of passing the **AP Chemistry Exam in May 2027!** While the test seems far away right now, it comes up fast. Every day in our class we will be learning new things and practicing new skills, so it is critical that you are prepared. You will begin your preparation by completing this summer assignment.

A good chemist has a reliable system of organization, so some of the tasks in the summer assignment will be to prepare your organization method. **You will need all materials organized and ready for class by the first week of school- you will need to have everything by Friday September 4th, 2026.**

Some of the tasks in the summer assignment are also to review and learn content. Your goal is to master the concepts and skills below prior to the first week of school. Some of the content below will be review from a first-year chemistry course (such as CITES), however there will be some new information, too. This assignment is meant to help you develop research and study skills for novel material so that you have those skills throughout the coming year. Completing this packet will help you review and assess your skills; a key to this packet will be sent out during the last week of summer for you to check your work. **You will need to have completed this packet and be ready for a test covering the material the second week of the school year.** This packet is not designed to teach you the material- you will need to rely on the resources below, or your own preference, such as Khan Academy.

Should you have any questions about links or resources this summer you are welcome to e-mail me at mkoury@lwsd.org – I check my e-mail weekly during the summer and it might be a while before I respond, so thank you for your patience! I am really looking forward to the year ahead and I hope you are, too. See you in the fall! 😊 –Ms. Koury

Summer Assignment Tasks:

1. **Obtain and organize your AP Chemistry binder/dividers**
   - You need a large (3-4 inch) binder and at least 10 dividers for the units of AP Chemistry.
   - *I will have a section of the classroom for you to store this in; you do not need to carry it around everyday.
   - I will be providing notes, practice problems, labs, and assessments that will help you master topics. In order to verify the rigor of the course, colleges will often ask to see these materials in addition to your AP Exam score. Keep everything until you receive credit from your college or university of choice.
   - You are responsible for keeping all materials in your binder for periodic checks throughout the unit. You cannot share this binder with another class. It needs to be dedicated to AP Chemistry.

2. **Obtain a dedicated folder with pockets for AP Chemistry**
   - You will need a folder with two pockets for your current work and reference materials.
   - You cannot share this folder with another class. It needs to be dedicated to AP Chemistry.

3. **Obtain a calculator**
   - You need to have a calculator you can use on classwork, homework, tests, and the AP Exam in May.
   - This calculator needs to come from the CollegeBoard approved list.
   - You may check out a graphing calculator from the LWHS library.
   - You need to be comfortable working whatever calculator you choose. While it isn't required, having a graphing calculator is advantageous for AP chemistry because it can store equations and help with significant figures.

4. **Review the 'Need to Know' AP Chemistry Topics & complete the practice problems in this packet**
   - Read the textbook sections and watch the videos. If you need more examples, search Khan Academy or YouTube for the topic names and watch a few more videos to get yourself acquainted with the material.
   - This is the content-learning portion of the summer assignment. As stated, most of these topics are review from a first-year chemistry course, but some of these topics will be new. **You are responsible for mastering this content and being prepared for a test by the second week of school.**
   - **SHOW YOUR WORK WHEN ASKED.** In AP Chemistry if you do not show work in an organized manner with attention to units and labels to support your answer, you will not receive credit. Practice this now!

5. **Come to school prepared for a test on the material outlined in this document.**
   - Complete the work contained in this packet and check your work with the key posted the last week of August.

Need to Know – AP Chemistry Topics

The following is a list of topics that you need to know and be able to do practice problems with. There are resources listed opposite the topics to help you including sections of a textbook (Chemistry: the Central Science by Brown & LeMay 13th Edition) and YouTube videos. Please note that Khan Academy is also an excellent resource for Topics 2, 3, and 4. I highly suggest using the "High School Chemistry and AP Chemistry" modules to look for these topics if that's your preferred resource. The pages that follow correspond to each topic, providing you will key questions you need to be able to answer, and practice problems that will help you practice the material. **My recommendation is that you wait and work on this in August prior to returning to school so that the information is fresh in your mind.**

Topics You Need to Master!	Resources

Topic 1: Math Skills	• **1.1 Significant Figures**
	  ○ Textbook: Chapter 1.5
	  ○ Significant Figures – Michael Farabaugh
	• **1.2 Scientific Notation**
	  ○ Textbook: Appendix A.1
	  ○ Scientific Notation – Michael Farabaugh
	• **1.3 Metric System**
	  ○ Textbook: Chapter 1.4
	  ○ Metric Conversions – Miss Marusik
	• **1.4 Unit Conversions with Dimensional Analysis**
	  ○ Textbook: Chapter 1.6
	  ○ Unit Conversions – Michael Farabaugh

Topic 2: Atomic Structure	• **2.1 Classification of Matter**
	  ○ Textbook: Chapter 1.2
	  ○ Types of Matter – Professor Dave Explains
	• **2.2 Subatomic Particles**
	  ○ Textbook: Chapter 2.1-2.3
	  ○ Basic Atomic Structure – Tyler DeWitt
	  ○ What are isotopes? – Tyler DeWitt
	• **2.3 Average Atomic Mass**
	  ○ Textbook: Chapter 2.4
	  ○ Average Atomic Mass vs Mass Number – Dr. B
	  ○ Solving for Average Atomic Mass – SciencePost
	  ○ Interpreting Mass Spectrometer Data – Nicholas Johnson
	• **2.4 Ion Formation & Charge**
	  ○ Textbook Chapter 2.7
	  ○ Valence Electrons and the Periodic Table –Tyler DeWitt
	  ○ How are ions formed? – Dr. B

Topic 3: Periodic Trends	• **3.1 Electron Configurations**
	  ○ Textbook Chapter 6.8—6.9
	  ○ Electron Arrangement: Bohr vs Schrodinger—Nicholas Johnson
	  ○ Electron Arrangement: Orbitals—Nicholas Johnson
	  ○ Introduction to Electron Configurations—Nicholas Johnson
	  ○ Electron Configuration –Jeremy Krug
	• **3.2 Coulomb's Law**
	  ○ Textbook Chapter 7.1—7.2
	  ○ Coulomb's Law—Jeremy Krug
	• **3.3 Atomic & Ionic Radius**
	  ○ Textbook Chapter 7.3
	  ○ Periodic Trend in Atomic Radius Part 1—Nicholas Johnson
	  ○ Periodic Trend in Atomic Radius Part 2—Nicholas Johnson
	• **3.4 Ionization Energy**
	  ○ Textbook Chapter 7.4
	  ○ Periodic Trend in Ionization Energy Part 1—Nicholas Johnson
	  ○ Periodic Trend in Ionization Energy Part 2—Nicholas Johnson
	  ○ Periodic Trend in Ionization Energy Part 3—Nicholas Johnson

Topic 4: Reactions & Stoichiometry	• **4.1 Counting Atoms and Compounds**
	  ○ Textbook Chapter 3.1
	  ○ Counting Atoms in a Chemical Equation—MooMoo Math and Science
	• **4.2 Balancing Chemical Reactions**
	  ○ Textbook Chapter 3.1—3.2
	  ○ How to Balance Chemical Equations—Dr. B
	• **4.3 Molar Mass**
	  ○ Textbook Chapter 3.3—3.4
	  ○ How to Calculation Molar Mass—Dr. B
	• **4.4 Mole Ratios**
	  ○ Textbook Chapter 3.6
	  ○ How to Find the Mole Ratio to Solve Stoichiometry Problems—Dr. B
	• **4.5 Stoichiometry**
	  ○ Textbook Chapter 3.6
	  ○ How to Solve Reaction Stoichiometry Problems—Dr. B

Topic 1: Math Skills

1.1: Significant Figures

- Read the textbook section and watch the video on Significant Figures.
- Take notes on what you learn about what significant figures are and why they are used in science.
- Be able to answer the following key questions:
  - What is the precision? What does this look like in terms of measurement in a lab?
  - What is accuracy? What does this look like in terms of measurement in a lab?
  - How do you measure to the limit of precision for a tool, like a ruler or graduated cylinder?
  - What are the rules for identifying if a number in a value is significant?
  - What are the rules for rounding your answer when adding or subtracting values?
  - What are the rules for rounding your answer when multiplying or dividing values?

For the following values, determine how many significant figures are in the value.

1. 584 _____ sig figs
2. 902 _____ sig figs
3. 92020 _____ sig figs
4. 0.902 _____ sig figs
5. 0.888 _____ sig figs
6. 0.880 _____ sig figs
7. 6.548 _____ sig figs
8. 6.5480 _____ sig figs
9. 500 _____ sig figs
10. 500. _____ sig figs
11. 50.0 _____ sig figs
12. 50.00 _____ sig figs
13. 0.50 _____ sig figs
14. 0.050 _____ sig figs

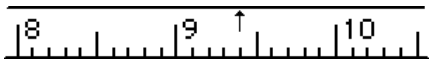
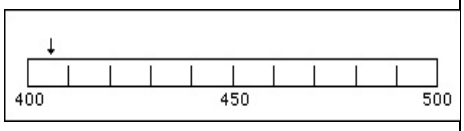
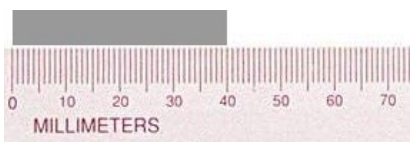
Complete the following calculations, rounding your reported answer to the correct number of significant figures:

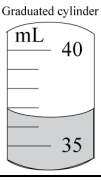
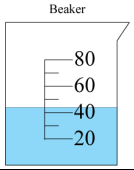
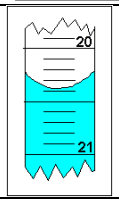
Ex.) 2.6 × 3.78 = 9.828 → 9.8

Ex) 4.51 – 2.1395 = 2.3705 → 2.37

15. 2.31 + 9.1028 =
16. 891.5 – 402 =
17. 3.15 × 2.5 × 4.00 =
18. 40.8 ÷ 5.05 =
19. 312 – 100 =
20. 2451.1- 2450. =
21. 0.0036 × 0.02 =
22. 0.075 ÷ 0.030 =
23. 102 × 0.1924 =
24. 5,600 × 12.31 =

For each of the following devices, Indicate each measurement to the correct limit of precision.

Device	Answer
25. 	_____ cm
27. 	_____ cm
29. 	_____ mm

Device	Answer
26. 	_____ mL
28. 	_____ mL
30. 	_____ mL

1.2: Scientific Notation

- Read the textbook section and watch the video on Scientific Notation (also called Exponential Notation).
- Take notes on what you learn about why this is used.
- Be able to answer the following key questions:
  - What are the rules for converting from standard to scientific notation?
  - What should the exponent be when representing large values? How about small values?
  - How do you incorporate significant figures?

Convert these numbers from scientific notation to standard notation. Then perform the calculations using your calculator. Don't forget to incorporate significant figures into your final answers!

	Scientific	Standard	Calculations Practice
31.	2.3×10^{-3}		39. $(2.31 \times 10^{3}) \times (1.25 \times 10^{5})=$
32.		465,000	40. $(1.64 \times 10^{-6}) \times (8.91 \times 10^{11})=$
33.		0.00187	41. $(3.5 \times 10^{-4}) \times (1.2 \times 10^{-5}) =$
34.	8.20×10^{6}		42. $(7.1 \times 10^{3}) \div (1.2 \times 10^{5})=$
35.		0.000091	43. $(2.10 \times 10^{-4}) \div (1.32 \times 10^{6})=$
36.	3.12×10^{2}		44. $(9.23 \times 10^{8}) \div (3.12 \times 10^{6})=$
37.		1,102,000	
38.	7.120×10^{-7}		

1.3: Metric Conversions

- Read the textbook section and watch the video on Metric System (SI Units).
- Take notes on what types of units are used for different quantities (mass, length, time, temperature).
- Be able to answer the following key questions:
  - What are the SI Base Units for each type of quantity?
  - What are the metric prefixes that you will use and should know? (Check out the reference sheet!)

Solve the following metric conversions:

45. 2000 mg = _____ g
46. 104 km = _____ m
47. 480 cm = _____ m
48. 5.6 kg = _____ g
49. 8.2 mm = _____ cm
50. 5.0 L= _____ ml
51. 198 g = _____ kg
52. 75 mL = _____ L
53. 50.0 cm = _____ m
54. 5.6 m = _____ cm
55. 16 cm = _____ m
56. 2500 m = _____ km
57. 65.0 g = _____ mg
58. 6.3 cm = _____ m
59. 120 mg = _____ g

1.4: Unit Conversions with Dimensional Analysis

- Read the textbook section and watch the video on Dimensional Analysis
- Take notes on this process – make sure to follow along with examples and learn to SHOW YOUR WORK. **Being able to properly document the process is just as important as getting the answer.**
- Be able to answer the following key questions:
    - What is a conversion factor?
    - How do you use two or more conversion factors to convert between units?

Note: Using conversion factors to change between units will be done almost daily in AP Chemistry. This is a critical topic to our course and being able to show your work in this way is exactly what will be necessary on the AP Exam!

Complete the following conversions using the process of dimensional analysis. Show every conversion factor you use, including how the units cancel in the numerator and denominator. Round your final answer to an appropriate amount of significant figures based on the measurement given in the problem.

60. Ms. Koury drives 7.4 miles to work every day. How many kilometers is this? (1 mile = 1609 m)

61. Ms. Koury loves springtime because of all the little snails that she finds with her two kids on nature walks. If a snail moves 4 cm/min, what is the snail's speed in km/hr?

62. Ms. Koury's youngest son, Griffin, has a headache and takes some Tylenol to help. If Griffin is 42 lbs and the recommended dose of Tylenol is 5 mL per 45 kilograms of body weight, how many mL of Tylenol should Griffin take? (2.2 lb = 1 kg)

63. Ms. Koury drinks the standard 8 cups of water a day. If 1 cup of water contains 240 mL and the density of water is 1 g/mL, how many milligrams of water does she consume every day?

64. A red photon of light is emitted from Ms. Koury's laser pointer as she teases Ms. Thody's cat, Sasha. If the wavelength of the red light is 720 nanometers, how many meters is this?

Topic 2: Atomic Structure

2.1: Classification of Matter:

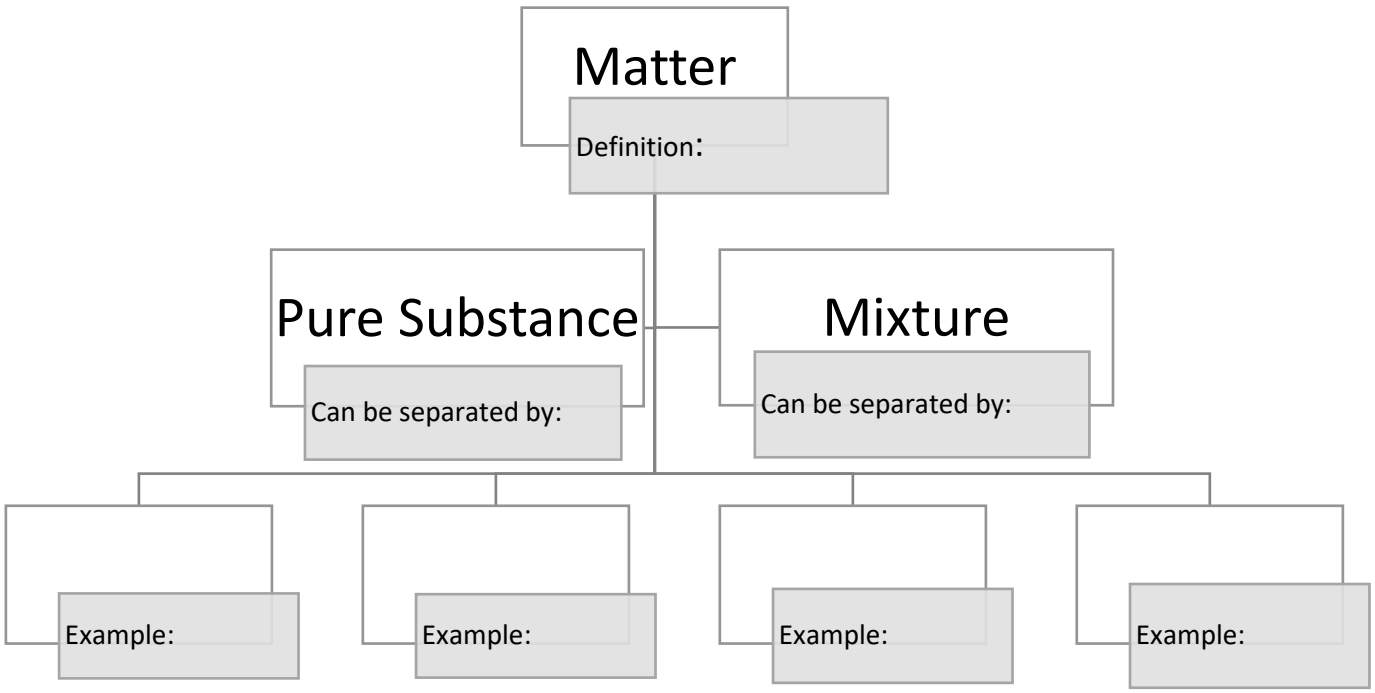
- Read the textbook section and watch the video on the Classification of Matter
- Take notes on how we can describe matter in different ways.
- Be able to answer the following key questions:
  - What are the three states of matter? Describe how they are different from each other.
  - What is the difference between a physical and chemical change?
  - What is the difference between a pure substance and a mixture?

Note: The hardest part about chemistry is that we can't see it. We cannot see individual atoms and molecules moving around, we only see the result of their movement in the macroscopic world. The better you get at imagining and visualizing the molecular world, the better chemist you will become!

65. Draw a particle model for each of the three states of matter. Describe the speed and energy of the particles in each phase.

Solid	Liquid	Gas
Drawing of Particles:	Drawing of Particles:	Drawing of Particles:
Description of Movement:	Description of Movement:	Description of Movement:

66. Complete the following classification chart with how matter is organized. Draw particle models next to each type of matter to show the differences. Then, complete the grey boxes to demonstrate your knowledge of how chemists differentiate between each type.



2.2: Subatomic Particles

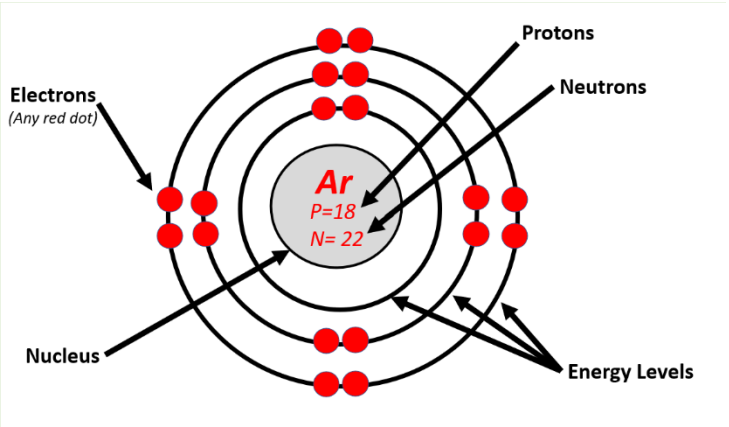
- Read the textbook section and watch the video on the Atomic Theory, focusing on the Modern View of the Atom
- Take notes on Dalton's Atomic Theory and the different subatomic particles.
- Be able to answer the following key questions:
  - What are the three subatomic particles of the atom?
  - What are the charges of each subatomic particle?
  - What is the mass (in amu) of each subatomic particle?
  - What does each subatomic particle do for the atom?

Note: The atom is the smallest unit of an element, which is a pure substance. Our first unit of the year will be all about how to visualize and count atoms. Knowing the structure of the atom is critical because the mass and charge of each subatomic particle will determine how that atom behaves when interacting with other atoms!

67. Complete the following chart using your knowledge of subatomic particles:

Subatomic Particle	Location in Atom	Charge	Mass (amu)	Job (What does this do for the atom?)
Proton				
Neutron				
Electron				

The model of the atom to the right is called a Bohr Model. This is one type of representation of the atom that we will see in AP Chemistry. Using this model, answer the following questions about subatomic particles:



68. Where are the protons and neutrons located in the atom?

69. Where are the electrons located in the atom?

70. How many energy levels does the atom have in the Bohr Model?

71. How many electrons can fit into each of the energy levels?
    a. 1st (closest to the nucleus)
    b. 2nd
    c. 3rd (farthest from the nucleus)

Read the passage in the textbook on Atomic Numbers, Mass Numbers, and Isotopes. Answer the following questions:

72. What makes an atom of one element, like Nitrogen, different from an atom of another element, like Oxygen?

73. The number of protons in any particular element is called that element's _____.

74. The mass number represents the total amount of _____ + _____ in an atoms nucleus.

75. Atoms with identical atomic numbers but different mass numbers (that is, same number of protons but different number of neutrons) are called _____.

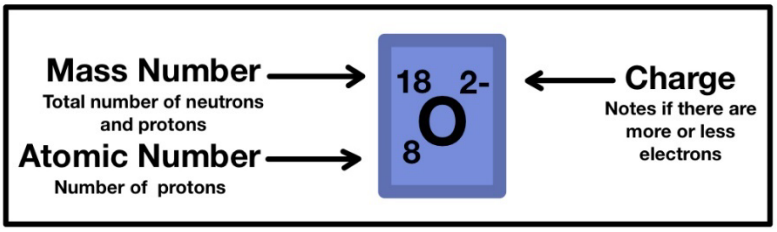
76. An atom has no net electrical charge, so what two subatomic particles must be equal in amounts?

<u>For the remaining sections of the summer assignment, you will need to use a periodic table! Here are some options:</u>

- **Ptable.com**
- **AP Chemistry Periodic Table – You will notice that you DO NOT HAVE NAMES on the periodic table given to you by College Board for this class, this is because the names of an atom will always be included in the work next to the symbol.** 😊

Scientists represent isotopes of different elements using isotope notation. This notation is used because scientists can quickly determine how many protons, neutrons, and electrons are in each atom.

Don't worry about the charge section… yet! For now, atoms are electrically neutral.



77. Using the first row as an example for the isotope, Magnesium-22, complete the table for each isotope. Assume all atoms are electrically neutral.

EXAMPLE:

$^{22}_{12}Mg$

Magnesium- 22

Name	Isotopic Notation	Protons	Neutrons	Electrons	Atomic Number	Mass number
Magnesium-22	$^{22}_{12}Mg$	12	10	12	12	22
	$^{31}_{15}P$					
	$^{27}_{13}Al$					
				38	38	87
		15	16	15		
		10				20

2.3 Average Atomic Mass

- Read the textbook section on Atomic Weights and watch the video on Calculating Average Atomic Mass
- Take notes on how to calculate the average atomic mass of a substance and the purpose of a mass spectrometer.
- Be able to answer the following key questions:
  - What does the average atomic mass of an element represent?
  - Explain why we don't divide by the number of isotopes to find the average mass.
  - What is a mass spectrometer and what type of data does it give chemists?

78. An unknown element has four naturally occurring isotopes. Their actual masses in amu and percent abundances are given in the table below.

Percent Abundance	Atomic Mass (amu)
0.56	83.91
9.86	85.91
7.00	86.91
82.58	87.91

a. Use the data above to calculate the average atomic mass of this unknown element.

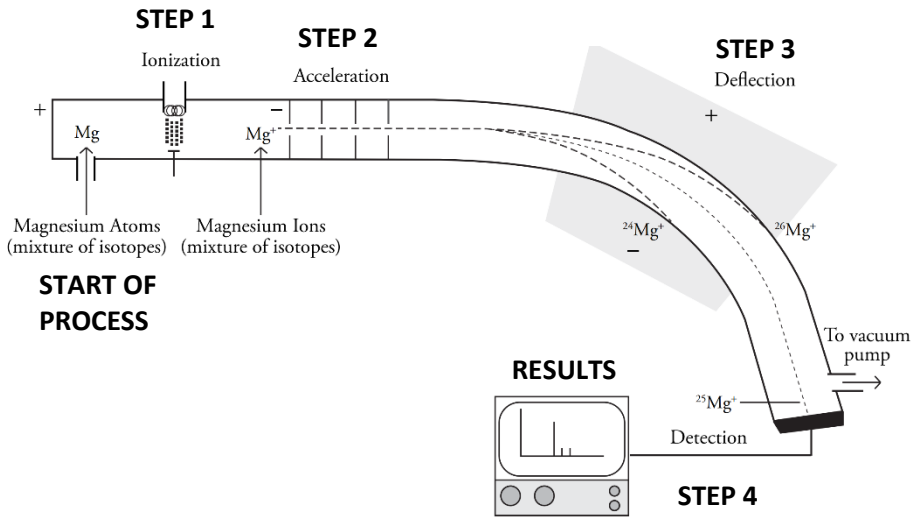
b. Use the periodic table to identify this unknown element.

79. Bromine has two naturally occurring isotope. Bromine-81 and Bromine-79. Bromine has an average atomic mass of 79.90 amu. What can you conclude about the relative abundances of these two isotopes given Bromine's average atomic mass? Is the more of one isotope than the other or do they have relatively similar amounts? Explain how you know.

80. Calculate the atomic mass for Gallium-71 using the data given below:

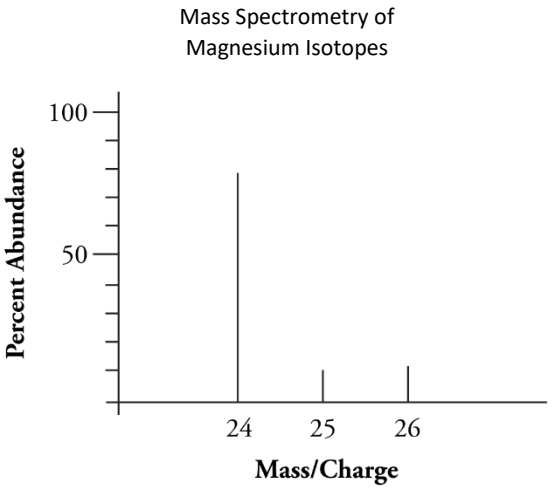
Isotope	Atomic Mass (amu)	Abundance (%)	Relative Contribution
Gallium-69	68.918	60.00108	41.35824
Gallium-71			
		Total Average Atomic Mass:	69.723

Video: Mass Spectroscopy is the elemental analysis method that scientists use in order to identify and differentiate isotopes of different substances. The process starts with a sample containing a mixture of isotopes being introduced into a vacuum tube where no air exists. The sample is then charged through a process called ionization and accelerated through a series of magnetic and electrical fields until a large electromagnet deflects the particles. Deflection allows the particles to be separated based on their mass. The detector determines how many times the particles hit the sides of the tube and generates a graph based on relative abundance.



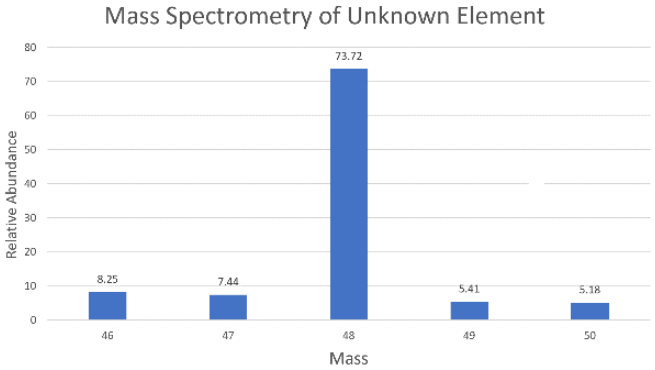
81. Use the mass spectrum of Magnesium to the right to answer the following questions:
    a. What is the mass number of the most common isotope of Magnesium?

    b. What is the percent abundance of the most common isotope of magnesium?



    c. Calculate the average atomic mass of Magnesium based on the Mass Spec data in the graph.

82. The following is a graph generated by a mass spectrometer for an unknown element. Using this data, calculate the average atomic mass and identify the element.

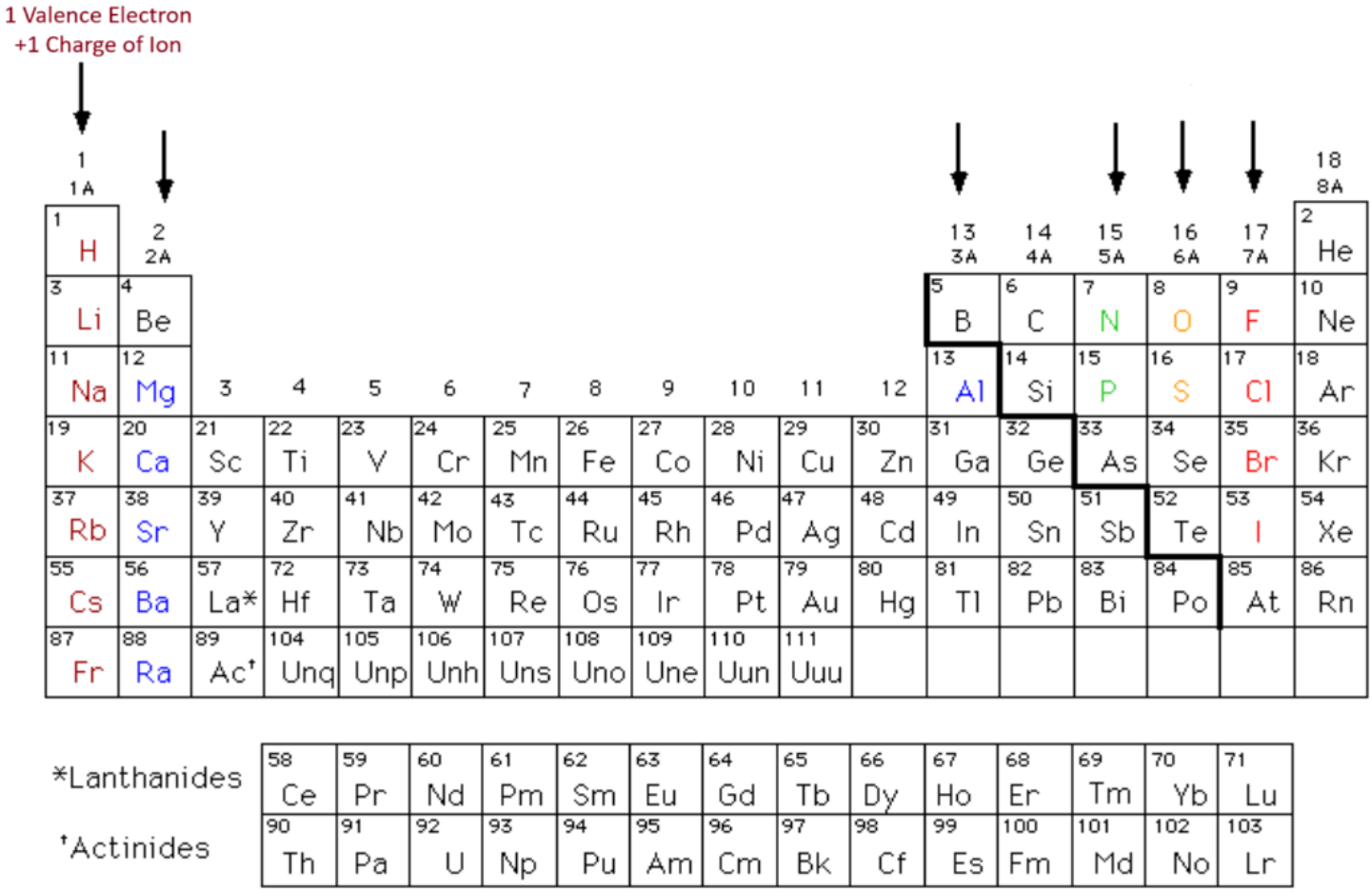


2.4 Ion Formation & Charge

- Read the textbook section and watch the videos on What is an Ion and Valence Electrons & the Periodic Table
- Take notes on how ions form and the patterns that exist across the periodic table.
- Be able to answer the following key questions:
  - What are the two types of ions? What charges do they have?
  - What types of electrons are gained and lost when ions are made?
  - What patterns exist in the number of valence electrons and ion formation in the periodic table?

Note: Atoms can bond to other atoms by gaining or losing electrons from each other. This is a central idea in AP chemistry and one that will be integral to your understanding of chemical reactions. ***MEMORIZE*** *the patterns for valence electrons and cation/anion formation on the periodic table. You need to be able to recite them as fast as your phone number or address!*

83. Using the model of the periodic table below, write in the number of valence electrons and charge formed by the most common ion for each column of main group elements- shown with arrows. Group 1 (Alkali Metals) has been done for you as an example.



84. Complete the following table to determine the most common ion formed for each element listed. The first row has been completed for you as an example.

Element	Number of Valence Electrons	Lose or Gain ____ Electrons	Cation or Anion?	Ion Formed
Lithium (Li)	1	Lose 1 e^-	Cation	Li^{+1}
Magnesium (Mg)				
Nitrogen (N)				
Aluminum (Al)				
Chlorine (Cl)				
Oxygen (O)				
Phosphorous (P)				
Potassium (K)				
Carbon (C)				
Sodium (Na)				
Fluorine (F)				
Silicon (Si)				
Sulfur (S)				
Iodine (I)				

Polyatomic Ions

Note: Polyatomic ions are groups of atoms joined in a molecule, but carrying a net positive or negative charge. There are several common polyatomic ions that will occur over and over in AP Chemistry that you need to ***MEMORIZE – both their name and chemical formula with charge.*** *Again, you should be able to recite these as fast as your phone number or address.*

Please notice that although you are memorizing the -ates, you should also be able to determine the per- -ates, -ites, and hypo- -ites, using the rule outlined to the right below.

85. Polyatomic Ions to Know and ***MEMORIZE***. Make yourself flashcards or a quizlet!

Name	Formula	Name	Formula
Acetate	$C_2H_3O_2^{-1}$	Permanganate	MnO_4^{-1}
Bromate	BrO_3^{-1}	Peroxide	O_2^{-2}
Carbonate	CO_3^{-2}	Phosphate	PO_4^{-3}
Chlorate	ClO_3^{-1}	Sulfate	SO_4^{-2}
Dichromate	$Cr_2O_7^{-2}$	Thiocyanate	SCN^{-1}
Hydroxide	OH^{-1}	Thiosulfate	$S_2O_3^{-2}$
Iodate	IO_3^{-1}	Ammonium	NH_4^{+1}
Nitrate	NO_3^{-1}	Hydronium	H_3O^{+1}

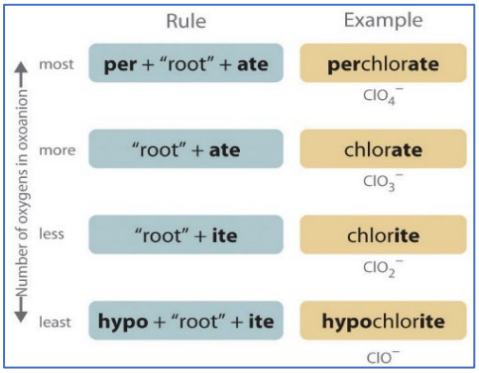


TABLE OF POLYATOMIC IONS					
acetate	CH_3COO^-	dihydrogen phosphate	$H_2PO_4^-$	oxalate	$C_2O_4^{2-}$
arsenate	AsO_4^{3-}	hydrogen carbonate	HCO_3^-	perchlorate	ClO_4^-
arsenite	AsO_3^{3-}	hydrogen oxalate	$HC_2O_4^-$	periodate	IO_4^-
benzoate	$C_6H_5COO^-$	hydrogen sulfate	HSO_4^-	permanganate	MnO_4^-
borate	BO_3^{3-}	hydrogen sulfide	HS^-	peroxide	O_2^{2-}
bromate	BrO_3^-	hydrogen sulfite	HSO_3^-	phosphate	PO_4^{3-}
carbonate	CO_3^{2-}	hydroxide	OH^-	pyrophosphate	$P_2O_7^{4-}$
chlorate	ClO_3^-	hypochlorite	ClO^-	sulfate	SO_4^{2-}
chloride	Cl^-	iodate	IO_3^-	sulfite	SO_3^{2-}
chlorite	ClO_2^-	monohydrogen phosphate	HPO_4^{2-}	thiocyanate	SCN^-
chromate	CrO_4^{2-}	nitrate	NO_3^-	thiosulfate	$S_2O_3^{2-}$
cyanate	CNO^-	nitrite	NO_2^-	POSITIVE POLYATOMIC IONS	
cyanide	CN^-	orthosilicate	SiO_4^{4-}	ammonium	NH_4^+
dichromate	$Cr_2O_7^{2-}$			hydronium	H_3O^+

PERIODIC TABLE OF IONS

KEY: atomic number; symbol; Fe^{3+} ion charge; iron (III) ion name (IUPAC); Fe^{2+} iron (II)

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H^+ hydrogen																1 H^- hydride	2 He helium
3 Li^+ lithium	4 Be^{2+} beryllium											5 B boron	6 C carbon	7 N^{3-} nitride	8 O^{2-} oxide	9 F^- fluoride	10 Ne neon
11 Na^+ sodium	12 Mg^{2+} magnesium											13 Al^{3+} aluminum	14 Si silicon	15 P^{3-} phosphide	16 S^{2-} sulfide	17 Cl^- chloride	18 Ar argon
19 K^+ potassium	20 Ca^{2+} calcium	21 Sc^{3+} scandium	22 Ti^{4+} titanium (IV); Ti^{3+} titanium (III)	23 V^{3+} vanadium(III); V^{5+} vanadium (V)	24 Cr^{3+} chromium (III); Cr^{2+} chromium (II)	25 Mn^{2+} manganese(II); Mn^{4+} manganese(IV)	26 Fe^{3+} iron (III); Fe^{2+} iron (II)	27 Co^{2+} cobalt (II); Co^{3+} cobalt (III)	28 Ni^{2+} nickel (II); Ni^{3+} nickel (III)	29 Cu^{2+} copper (II); Cu^+ copper (I)	30 Zn^{2+} zinc	31 Ga^{3+} gallium	32 Ge^{4+} germanium	33 As^{3-} arsenide	34 Se^{2-} selenide	35 Br^- bromide	36 Kr krypton
37 Rb^+ rubidium	38 Sr^{2+} strontium	39 Y^{3+} yttrium	40 Zr^{4+} zirconium	41 Nb^{5+} niobium (V); Nb^{3+} niobium(III)	42 Mo^{6+} molybdenum	43 Tc^{7+} technetium	44 Ru^{3+} ruthenium(III); Ru^{4+} ruthenium(IV)	45 Rh^{3+} rhodium	46 Pd^{2+} paladium(II); Pd^{4+} paladium(IV)	47 Ag^+ silver	48 Cd^{2+} cadmium	49 In^{3+} indium	50 Sn^{4+} tin (IV); Sn^{2+} tin (II)	51 Sb^{3+} antimony(III); Sb^{5+} antimony(V)	52 Te^{2-} telluride	53 I^- iodide	54 Xe xenon
55 Cs^+ cesium	56 Ba^{2+} barium	57 La^{3+} lanthanum	72 Hf^{4+} hafnium	73 Ta^{5+} tantalum	74 W^{6+} tungsten	75 Re^{7+} rhenium	76 Os^{4+} osmium	77 Ir^{4+} iridium	78 Pt^{4+} platinum(IV); Pt^{2+} platinum(II)	79 Au^{3+} gold (III); Au^+ gold (I)	80 Hg^{2+} mercury (II); Hg^+ mercury (I)	81 Tl^+ thallium (I); Tl^{3+} thallium(III)	82 Pb^{2+} lead (II); Pb^{4+} lead (IV)	83 Bi^{3+} bismuth(III); Bi^{5+} bismuth(V)	84 Po^{2+} polonium(II); Po^{4+} polonium(IV)	85 At^- astatide	86 Rn radon
87 Fr^+ francium	88 Ra^{2+} radium	89 Ac^{3+} actinium															

58 Ce^{3+} cerium	59 Pr^{3+} praseodymium	60 Nd^{3+} neodymium	61 Pm^{3+} promethium	62 Sm^{3+} samarium(III); Sm^{2+} samarium(II)	63 Eu^{3+} europium (III); Eu^{2+} europium (II)	64 Gd^{3+} gadolinium	65 Tb^{3+} terbium	66 Dy^{3+} dysprosium	67 Ho^{3+} holmium	68 Er^{3+} erbium	69 Tm^{3+} thulium	70 Yb^{3+} ytterbium(III); Yb^{3+} ytterbium(II)	71 Lu^{3+} lutetium
90 Th^{4+} thorium	91 Pa^{5+} protactinium(V); Pa^{4+} protactinium(IV)	92 U^{6+} uranium (VI); U^{4+} uranium (IV)	93 Np^{5+} neptunium	94 Pu^{4+} plutonium(IV); Pu^{6+} plutonium(VI)	95 Am^{3+} americium(III); Am^{4+} americium(IV)	96 Cm^{3+} curium	97 Bk^{3+} berkelium(III); Bk^{4+} berkelium(IV)	98 Cf^{3+} californium	99 Es^{3+} einsteinium	100 Fm^{3+} fermium	101 Md^{2+} mendelevium (II); Md^{3+} mendelevium (III)	102 No^{2+} nobelium(II); No^{3+} nobelium(III)	103 Lr^{3+} lawrencium

Topic 3: Periodic Trends

3.1 Electron Configurations

- Read the textbook section and watch the videos on electron configurations
- Take notes on how the Bohr Model relates to the Quantum Mechanical Model of the atom, paying close attention to how to write electron configurations and orbital diagrams for electron placement in this model.
- Be able to answer the following key questions:
  - What is wrong with the Bohr Model? Why do chemists need to use the Quantum Mechanical Model?
  - What is the difference between an energy level, sublevel and orbital?
  - How can you tell the difference between core and valence electrons using electron configurations?

86. Write the electron configuration for the following neutral atoms:

    a. Helium

    b. Carbon

    c. Fluorine

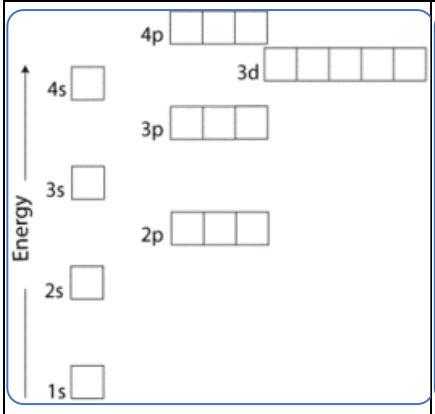
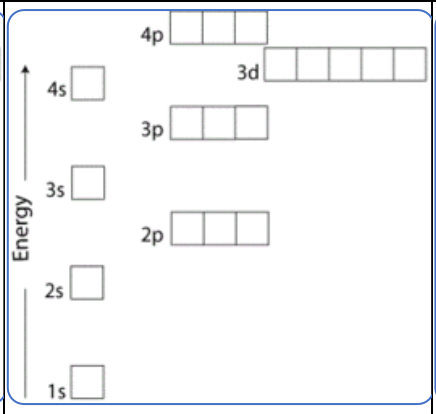
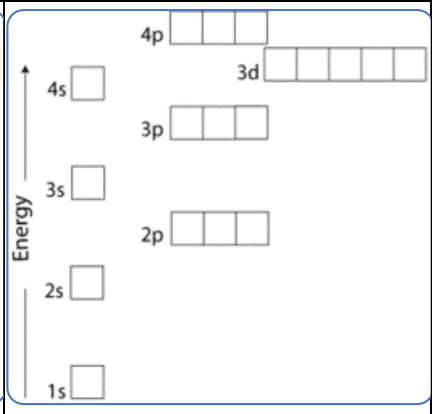
    d. Bromine

    e. Cobalt

    f. Strontium

    g. Iodine

87. Complete the following orbital diagrams for atoms in their ground state (neutral configuration). Then, list how many core electrons and valence electrons there are for each atom (remember a valence electron exists in the outermost energy level!)

Zinc	**Sulfur**	**Potassium**
		
# of Core Electrons: _____ # of Valence Electrons: _____	# of Core Electrons: _____ # of Valence Electrons: _____	# of Core Electrons: _____ # of Valence Electrons: _____

3.2 Coulomb's Law

- Read the textbook section and watch the videos on Coulomb's Law
- Take notes on the mathematical formula and how it relates to the charges within an atom.
- Be able to answer the following key questions:
    - Describe Coulomb's Law and how the strength of attraction changes with the magnitude of charge and the distance between the charges.
    - How does the effective nuclear charge change as you move left to right across a period? Down a group?
    - How does the shielding effect change as you move left to right across a period? Down a group?

Note: You will not have to do calculations with Coulomb's Law – but you need to understand how this equation can be useful when applied to atoms in terms of number of electrons in the nucleus (for q) and number of energy levels.

88. Coulomb's Law describes the force of attraction or repulsion between charged particles. Given this, explain why electrons are held more tightly in atoms with a higher nuclear charge, even when the number of energy levels stays the same.

89. Consider two atoms: Atom A which has 6 protons and Atom B which has 8 protons. Both atoms have electrons in the same valence energy level. Using Coulomb's Law, predict which atom would exert a greater attractive force on its valence electrons and explain why.

90. A neutral atom becomes a cation when it loses one or more electrons. Using Coulomb's Law, explain how the electrostatic force within the atom changes *after* the loss of electrons and how this affects the attraction between the remaining electrons and the nucleus.

3.3 Atomic & Ionic Radius

- Read the textbook section and watch the videos on Atomic & Ionic Radius
- Take notes on the how the atomic radius changes as you move across a period and down a group.
- Be able to answer the following key questions:
    - What is the trend in atomic radius as you move from top to bottom in a group?
    - What is the trend in atomic radius as you move from left to right in a period?
    - What is the trend in atomic versus ionic radius for a cation?
    - What is the trend in atomic versus ionic radius for an anion?

91. Explain why the atomic radius decreases across a period from left to right on the periodic table. In your explanation, reference both electron configuration and Coulomb's Law.

92. Although Sodium (Na) and Chlorine (Cl) have electrons in the third energy level, Chlorine has a significantly smaller atomic radius. Use Coulomb's Law and effective nuclear charge to justify this observation.

93. Compare the atomic radius of Magnesium (Mg) and Calcium (Ca). Both are in Group 2, but in different periods. Use electron configuration and Coulomb's Law to explain which atom is larger.

94. Describe how the atomic radius changes when a neutral atom becomes a cation, such as Na and Na^{+1}. Use Coulomb's Law to explain the difference in size between the ion and its neutral atom.

95. Describe how the atomic radius changes when a neutral atom becomes a cation, such as F and F^{-1}. Use Coulomb's Law to explain the difference in size between the ion and its neutral atom.

3.4 Ionization Energy

- Read the textbook section and watch the videos on Ionization Energy
- Take notes on what ionization energy is and how it changes according to atomic structure.
- Be able to answer the following key questions:
    - What is the trend in ionization energy as you move from top to bottom in a group?
    - What is the trend in ionization energy as you move from left to right in a period?

96. Explain why the ionization energy increases across a period from left to right on the periodic table. Use Coulomb's Law and effective nuclear charge in your explanation.

97. Compare the first ionization energies of Lithium (Li) and Cesium (Cs). Which one is higher, and why? Support your answer using Coulomb's Law.

98. Sodium's first electron is easily removed, however there is a large jump in energy required to move additional electrons. Explain why this occurs using concepts of electron configuration and Coulomb's Law.

99. Rank the following atoms from lowest to highest ionization energy: Li, C, O, Ne

100. Rank the following cations from lowest to highest ionization energy: Na^{+1}, Mg^{+2}, Al^{+3}

Topic 4: Chemical Reactions

4.1 Counting Atoms in Chemical Formulas

- Read the textbook section and watch the videos Counting Atoms in Chemical Formulas
- Take notes on how to determine the amount of atoms in a chemical formula.
- Be able to answer the following key questions:
  - What is the difference between a coefficient and a subscript?
  - How do you distribute a subscript that appears outside of the parenthesis?
  - Why are parenthesis used?

101. Determine the type and number of atoms in each chemical formula below. Follow the example below.

Example: H_2O
H (Hydrogen) = 2 atoms
O (Oxygen) = 1 atom

a. $Ba(OH)_2$

b. $(NH_4)_2SO_4$

c. $CaBr_2$

d. $NaC_2H_3O_2$

e. CH_4

f. P_2O_5

g. CuO

h. $NH_2CH_2CH_2CH_2O$

i. $AuCl_3$

j. $CH_3CH_2CH_2CH_3$

4.2 Balancing Chemical Reactions

- Read the textbook section and watch the videos on Balancing Chemical Reactions
- Take notes on the parts of a chemical reaction and how to balance the reaction with coefficients
- Be able to answer the following key questions:
  - Why can you not change the subscripts of a chemical formula?
  - Why is it necessary to balance chemical equations?

Write in the coefficients to balance the reactions below.

102. ____H_2 + ____ O_2 → ____H_2O

103. ____H_2O_2 → ____ H_2O + ____O_2

104. ____Na + ____O_2 → ____ Na_2O

105. ____N_2 + ____H_2 → ____ NH_3

106. ____CH_4 + ____O_2 → ____CO_2 + ____H_2O

107. ____Al_2O_3 → ____Al + ____O_2

4.3 Calculating Molar Mass (Formula Weight)

- Read the textbook section and watch the videos on Formula Weights/Calculating Molar Mass
- Take notes on how to calculate the molar mass of a chemical formula.
- Be able to answer the following key questions:
  - Why do you multiple the molar mass by the subscripts and not coefficients?
  - What are the units of molar mass?

Find the molar mass of each of the following compounds. Show work (don't just google them!)

108. Nitrogen gas: N_2

109. Calcium Hydroxide: $Ca(OH)_2$

110. Ammonia: NH_3

111. Nickel nitrate: $Ni(NO_3)_2$

112. Copper (II) Chloride: $CuCl_2$

113. Sucrose (table sugar): $C_{12}H_{22}O_{11}$

114. Convert between mass and moles using dimensional analysis.

element or compound	mass	moles	SHOW YOUR WORK – Use dimensional analysis and conversion factors. Round to appropriate significant figures!
Cl_2	1.00 g		
CO_2		2.51 mol	
NaCl	1.50 g		
CH_4		.75 mol	
O_2	1.00 g		
H_2O		2.51 mol	
NaCl		0.045 mol	
$C_6H_{12}O_6$		1.5 mol	

4.4 Mole Ratios

- Read the textbook section and watch the videos on Mole Ratios
- Take notes on how to convert from moles of one substance to moles of another.
- Be able to answer the following key questions:
    - What is a mole ratio?
    - How can you determine a mole ratio using a chemical equation?

Use the balanced reaction and mole ratio to answer the following questions.

115. $N_2 + 3\ H_2 \rightarrow 2\ NH_3$

How many moles of H_2 are needed to completely react with 2 moles of N_2?

116. $2\ KClO_3 \rightarrow 2\ KCl + 3\ O_2$

How many moles of O_2 would be produced if 6 moles of $KClO_3$ were used?

117. $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

How many moles of H_2 are produced when 3 moles of Zn are reacted with an excess of HCl?

118. $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$

How many moles of O_2 are needed to completely react with 4 moles of C_3H_8?

119. $K_3PO_4 + Al(NO_3)_3 \rightarrow 3\ KNO_3 + AlPO_4$

How many moles of KNO_3 are produced when 2 moles of K_3PO_4 are reacted with 2 moles of $Al(NO_3)_3$? Which is the limiting reagent?

4.5 Stoichiometry

- Read the textbook section and watch the videos on Reaction Stoichiometry
- Take notes on how to convert from mass of one substance to mass of another.
- Be able to answer the following key questions:
    - What three conversion factors do you need to convert from mass of one substance to mass of another?
    - How is this different than using the conversion factors from Topic 1 of your summer assignment?

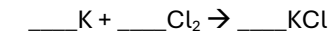
120. For the following chemical reaction, complete all stoichiometric calculations.



____H_2O_2 → ____H_2O + ____O_2

example	work	Answer (sig figs?)
3.0 grams of H_2O_2 decomposes in sunlight. Find the grams of H_2O produced.		
3.0 grams of O_2 was produced in this reaction. Find the grams of H_2O produced.		
3.0 grams of H_2O was produced in this reaction. Find the grams of H_2O_2 consumed.		

121. For the following chemical reaction, complete all stoichiometric calculations.



____K + ____Cl_2 → ____KCl

example	work	answer (sig figs?)
2.5 grams of K was exposed to 2.5 grams of Cl_2. Which reactant was in excess?		
How many grams of potassium chloride (KCl) will form?		

122. Write the balanced equation below then use the mole ratio to complete all calculations.

____Na + ____H_2O → ____H_2 +____NaOH

example	work	answer
1.0 gram of Na was added to 4.0 gram of water. Find the limiting reactant.		
Using the limiting reactant, find the grams of the other reactant consumed.		
Using the limiting reactant, find the grams of NaOH produced.		
Using the grams of NaOH produced, find the grams of H_2 produced.		
Using the limiting reactant, find the grams of H_2 produced.		

123. Write the balanced equation below then use the mole ratio to complete all calculations.

____C_2H_6 + ____O_2 → ____H_2O + ____CO_2

example	work	answer
3.5 g of ethane and 3.5 g of O_2 were present initially. Find the moles of both reactants.		
Find the limiting reactant.		