A Collection of Worksheets and Laboratory Experiments
2019 Edition | Bellevue College Chemistry Department Faculty

CHEM& 140 Workbook

VAN-GRINER LEARNING

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A Collection of Worksheets and Laboratory Experiments
Bellevue College Chemistry Department Faculty
2019 Edition

These experiments are designed to be used in college or university level laboratory courses, and should not be conducted unless there is an appropriate level of supervision, safety training, personal protective equipment and other safety facilities available for users. The publisher and authors believe that the lab experiments described in this manual, when conducted in conformity with the safety precautions described herein and according to appropriate laboratory safety procedures, are reasonably safe for students for whom this manual is directed. Nonetheless, many of the experiments are accompanied by some degree of risk, including human error, the failure or misuse of laboratory or electrical equipment, mis-measurement, spills of chemicals, and exposure to sharp objects, heat, blood, body fluids or other liquids. The publisher and authors disclaim any liability arising from such risks in connection with any of the experiments in the manual. Any users of this manual assume responsibility and risk associated with conducting any of the experiments set forth herein. If students have questions or problems with materials, procedures, or instructions on any experiment, they should always ask their instructor for immediate help before proceeding.

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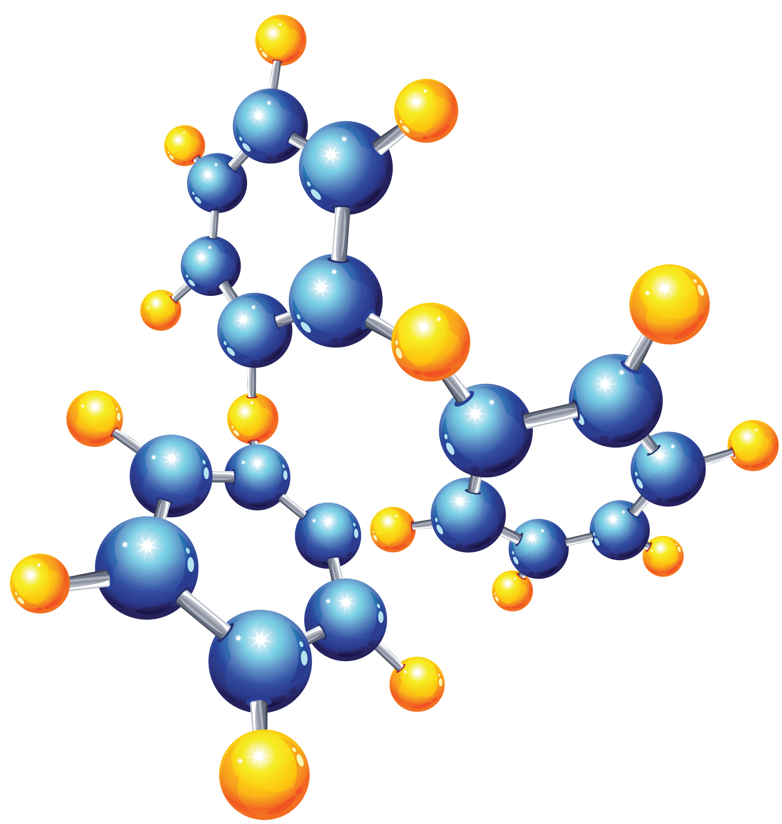


CHEM& 140 Student Learning Objectives

SECTIONS FROM CHEMISTRY LIBRETEXTS, A FREE ONLINE RESOURCE (CHEMISTRY TEXTBOOK)

https://chem.libretexts.org/Textbook_Maps/Introductory_Chemistry/map%3A_Introductory_Chemistry_(Tro)

Unit 1: The Science of Chemistry (Pages 6–18)	
1. Define chemistry as a science and explain how it is useful and relevant to everyday life.	1.1
2. Define and distinguish the major components of the "scientific approach to knowledge": observations, hypothesis, experiments, scientific law, and theory.	1.3–1.4
3. Distinguish between what is considered science and what is a belief.	Workbook
4. State and apply the law of conservation of mass.	3.7
5. **Memorize** chemical symbols for elements with atomic numbers 1 through 56 and 78 through 84 and locate them on the periodic table.	
6. Define matter, atom, molecule, pure substance, element, compound, mixture, heterogeneous mixture, and homogeneous mixture, and classify a given substance as an element/compound/heterogeneous or homogeneous mixture based on its chemical composition.	3.2, 3.4
7. Interpret and draw atomic/molecular scale drawings.	Workbook
8. Explain what is meant by physical state of matter and describe the differences among a solid, liquid, and gas in terms of particle position and strength of attraction.	3.3
9. Determine whether changes to matter are chemical or physical in nature, and which properties of matter are chemical versus physical.	3.5–3.6
10. Write the chemical formula for a compound given the number of atoms, and given the formula of a compound, state the number of atoms (including the use of parentheses in chemical formulas).	5.3
Unit 2: Measurements and Units (Pages 19–47)	
11. Use scientific notation for a given number and apply this in multiplication, division, addition, subtraction, powers and roots. Review exponent rules ($10^a \times 10^b = 10^{a+b}$, etc).	2.2
12. Distinguish between exact and measured numbers; count the number of significant figures in a measured number, record measurements using proper number of significant figures, andreport correct number of significant figures in single and multi-step calculations, including temperature conversions—do not memorize °C/°F conversion.	2.1, 2.3–2.4
13. Describe the difference between accuracy and precision and calculate percent error. (Workbook, Lab)	

14. State the base units of measurements in SI and metric system. **Memorize** the metric prefixes: **tera, giga, mega, kilo, deci, centi, milli, micro, nano, and pico,** and their relationship to base unit and apply them in metric conversions.	2.5
15. Apply dimensional analysis by identifying and using conversion factors to solve single-step and multi-step calculation problems, including density.	2.6–2.9
16. Perform square and cubic unit conversions (units raised to a power). Interconvert area/volumes involving metric (memorized prefixes) and English system (provided). **Memorize 1 mL = 1 cm^3 = 1 cc.**	2.8; Workbook;Lab
Unit 3: Atoms and Elements (Pages 48–61)	
17. Identify the main points of Dalton's atomic theory (early 1800s) and some of the ideas and natural laws leading up to it. Explain what is wrong with Dalton's atomic theory according to our current knowledge of atomic structure.	4.2
18. Explain the difference between Thomson's plum pudding model of the atom and Rutherford's nuclear model. Describe the observations which led to Rutherford's conclusions about atomic structure.	4.3
19. State the relative masses and charges of the subatomic particles: protons, neutrons, and electrons. Explain where most of the atomic mass is located in an atom.	4.4
20. Use the atomic number, mass number, and charge given the ${}^{A}_{Z}X$ notation of isotopes to determine the number of protons, neutrons, and electrons in neutral atoms.	4.5–4.6, 4.8
21. For ions, explain the difference between atoms versus ions, and cations versus anions, in reference to the subatomic particles.	4.7
22. Identify regions in periodic table as metals, nonmetals, semimetals (metalloids), main group (representative) elements, transition metals (elements), inner transition metals or elements (lanthanides and actinides), and also by families/groups in periodic table—alkali metals, alkaline earth metals, halogens, and noble gases.	4.6
23. List the common properties of metals, nonmetals, and metalloids.	4.6
24. Calculate average atomic mass for an element from the given natural abundance and mass of each isotope of that element. For elements with two or more naturally occurring isotopes, calculate the natural abundance of each isotope or isotope masses from provided data.	4.9
Unit 4: Molecules and Compounds (Pages 62–79)	
25. Explain this statement by providing specific examples: When elements combine to form compounds, an entirely new substance is created with properties different than the constituent elements.	5.1
26. Use the Bohr model to predict the charge of ions for ionic compounds. Contrast with covalent compounds which share electrons rather than form ions.	Workbook
27. Identify the basic unit of an element or compound as atoms, molecules, formula units, or molecules. Given a chemical formula, determine if ionic bonding or covalent bonding is present.	5.4

28. Predict the correct formula for an ionic compound. Write the name of binary ionic compound from a given formula that has fixed charge metals and variable charge metals (using Roman numerals), and vice versa, and for ionic compounds containing polyatomic ions. Memorize seven required polyatomic ions (sulfate, nitrate, carbonate, chlorate, phosphate, hydroxide and ammonium) and derive others based on these.	5.5–5.7
29. Write the binary molecular compound name using Greek prefixes (**memorize 1 to 10**) if given the formula, and vice versa.	5.8
30. Given a chemical formula, classify it as an ionic or a covalent (molecular) compound and name it (or given the name, provide its chemical formula).	5.10
Unit 5: The Mole and Chemical Composition (Pages 80–95)	
31. Explain the concept of the mole in chemistry and why it is useful.	6.1–6.3
32. Distinguish between formula masses (in amu) and molar masses (in g/mol).	
33. Convert grams to moles (and vice versa) and moles to number of particles such as atoms, molecules, formula units (and vice versa), using molar mass and Avogadro's number as conversion factors.	6.3–6.4
34. Given a chemical formula, identify possible mole ratios and use the ratio as a conversion factor.	6.5
35. Given a chemical formula, calculate the mass percent of each element in the compound. Use mass percent as a conversion factor to convert between mass of an element and mass of a compound.	6.6–6.7
36. Define the empirical formula and molecular formula, and determine the empirical formula if the molecular formula is given. Determine the empirical formula of an unknown compound from mass or mass percent data.	6.8
Unit 6: Chemical Reactions (Pages 96–100)	
37. Describe or identify observations that indicate a chemical change (reaction) has occurred.	7.2
38. Distinguish between reactants and products, and the meaning of the chemical arrow. Write and balance chemical equations. Translate sentences describing a chemical reaction into a balanced chemical equation. Apply the abbreviations for the four physical states in the balanced equations: solid, liquid, gas, aqueous.	7.3–7.4
39. Classify reactions as synthesis (combination), decomposition, single/double replacement, and combustion.	7.10
40. Define combustion reaction, and write the chemical equation for the combustion of a fuel (given its formula).	7.9, end of section only
Unit 7: Stoichiometry (Pages 101–114)	
41. Explain what coefficients in chemical equations represent on a molecular level and macroscopic (molar) scale. Use balanced chemical equations to write out mole ratios, and use them to determine moles of a substance reacted or formed when moles of another substance are known, *i.e.,* moles A $\longrightarrow$ moles B.	8.2–8.3
42. Apply the concept of mole ratios in a balanced chemical equation and use molar mass conversions to determine the **theoretical yield** of a product when the mass of any one reacant is known, assuming the other reactant(s) is in excess. g A $\longrightarrow$ moles A $\longrightarrow$ moles B $\longrightarrow$ g B	8.4

43. Define limiting reactant and determine the limiting reactant and theoretical yield of a product for a chemical reaction given the moles or mass of two or more reactants.	8.5
44. Define actual yield and determine the percent yield for a chemical reaction given the moles or mass of one or more reactants.	8.6
Unit 8: Solution Concentrations (Pages 115–125)	
45. Define the following terms and identify specific examples of each in every day solutions: solution, solvent, solute, aqueous solution, concentration, dilute solution, and concentrated solution.	13.2–13.3
46. Define % (*m*/*m*) and molarity as conversion factors. Calculate the % (*m*/*m*) or molarity of a solution or use % (*m*/*m*) or molarity as a conversion factor to solve problems involving mass, moles, and volumes.	13.6
47. Distinguish concentration from density and identify when to use each concept, molarity or density.	
48. Solve solution stoichiometry problems to determine the volume or molarity of one reactant needed to react with a given volume or molarity of a second reactant. Given V and M of solution A $\longrightarrow$ moles A $\longrightarrow$ moles B $\longrightarrow$ M or V of solution B. Apply the concept of limiting reactant to reactions in aqueous solutions when necessary.	13.8
49. Define equivalence point and explain why it is useful. Define indicator and end point and explain how these are used in titrations.	Titration lab; 14.6
Unit 9: Electrons in Atoms (Pages 126–142)	
50. Describe the wave and particle nature of electrons and light. Identify the properties of a wave (frequency, wavelength, amplitude), recognize the Greek letters used to represent these properties, and determine which of these are directly proportional and inversely proportional using the relationship $c = \lambda\nu$. Discuss photons in relation to energy, frequency and wavelength using the relationship $E = h\nu$.	9.2
51. Name the seven major regions of electromagnetic radiation in order from lowest to highest energy (mnemonic: RMIVUXG). State the range of wavelengths of visible light (in nm) and name the seven majors colors of visible light (mnemonic: ROY G BIV).	9.3
52. Describe the Bohr model of the atom and how this model explains atomic emission spectra. Distinguish between continuous and atomic line spectra, and how it relates to quantized energy. Explain how the atomic emission spectrum of an element can be used to identify an element.	Spectroscopy lab; Workbook; 9.4
53. Describe the quantum mechanical model of the atom and why it is an improvement on the Bohr model. Define probability density and use it to describe where an electron exists in an orbital. Define principal level (or principal shell) and sublevel (or subshell).	9.5
54. Write the full and Noble gas electron configuration and orbital box diagram for atoms, cations, and anions formed from main group elements.	9.6
55. Use the electron configuration to determine the valence and core electrons for main group elements and relate this to general shape of periodic table: s-block, p-block, d-block, and f-block elements.	9.7

Unit 10: Chemical Bonding (Pages 143–153)	
56. Draw Lewis dot structures for an atom, and use electron configurations to review valence electrons. Describe the concept of the octet (or duet) rule to determine the Lewis dot structures for monatomic ions. Show how ionic compounds are formed by a transfer of electrons	10.2–10.3
57. Draw the Lewis dot structures to represent covalent bonds in molecular compounds and polyatomic ions. Show lone pairs (unshared or non-bonded pairs) and single, double, and triple bonds. Distinguish between covalent bonding and ionic bonding	10.4–10.5
58. Describe how valence shell electron pair repulsion (VSEPR) theory is used to determine molecular shape and count the number of "electron groups" around a central atom (one electron group = a single bond or a double or a triple bond or a lone pair on a central atom).	10.7
59. Given a molecule, predict its electron geometry (linear, trigonal planar, and tetrahedral) and the bond angle around the central atom. Determine molecular geometry (linear, bent, trigonal planar, trigonal pyramidal, tetrahedral) for single center molecules.	10.7

Scientific Method: Observation, Hypothesis, Law, or Theory?

- Observation: Qualitative descriptions or quantitative measurements of what happened in an experiment.
- Hypothesis: A tentative statement such as "if A happens then B must happen" that can be tested by direct experiment or observation and is falsifiable (can be proven false). A hypothesis supported by evidence can be expressed as a law or a theory. A disproven hypothesis can sometimes be re-tested and found correct as measurements improve.
- Law: A logical relationship (generalized rule) between two or more things that is based on a series of observations and supported hypotheses. It is often a mathematical statement of how two or more quantities relate to each other.
- Theory: A well-substantiated and reliable explanation for some aspect of nature. Explains ***why*** (the underlying reasons) certain laws exist and predicts natural behaviors beyond the observations and or laws that it is based on. May be modified based on further experiments. Theories are not said to be true; instead, they are said to be supported by evidence.
- Belief: A statement that is not scientifically provable in the same way as facts, laws, hypotheses, or theories. Scientifically disproven beliefs can still be held (by a person) to be true.

Law 1. For every action, there is an equal and opposite reaction.

Law 2. $F = ma$

Observation, fact 3. Water freezes at 32 °F.

Observation, fact 4. The Earth is a sphere.

Theory 5. Matter is composed of small particles called atoms that obey the law of conservation of mass.

Belief 6. Humans were created separately from all other life on Earth.

Theory 7. Humans and gorillas evolved from a common ancestor species.

Theory 8. Light is an electromagnetic phenomenon of energy described by Maxwell's Laws.

Hypothesis 9. The sun will likely die in 7.5 billion years.

Observation 10. Sunspots are colder than the surface of the Sun.

11. There are such things as ghosts. **Belief**

12. Matter can be converted into energy, and they are related by Einstein's equation, $E = mc^2$. **Law**

13. The positions of the planets can cause humans to act in specific ways. **Belief**

14. Momentum is the product of a body's mass and its velocity. **Law**

15. We will never know how life started on Earth. **Belief**

16. The Milky Way is a spiral-type galaxy. **Observation**

17. The sun will rise tomorrow morning. **Hypothesis**

18. The Earth is older than 10,000 years. **Observation**

19. Some numbers are luckier than others. **Belief**

20. Planetary orbits are ellipses and are governed by the inverse-square law for gravity and Newton's laws of motion. **Theory**

Adapted from http://image.gsfc.nasa.gov/poetry

Element Symbols

Memorize the names and element symbols for elements #1–56 and 78–84. You do not need to memorize the element numbers or order on the periodic table, just the names and element symbols!

1. Write the symbol for the following elements. Remember that the first letter is capitalized and the second letter is not capitalized.

   a. oxygen **O**
   b. hydrogen **H**
   c. chlorine **Cl**
   d. mercury **Hg**
   e. fluorine **F**
   f. barium **Ba**
   g. helium **He**
   h. uranium **U**
   i. radon **Rn**
   j. sulfur **S**
   k. plutonium **Pu**
   l. americium **Am**
   m. radium **Ra**
   n. germanium **Ge**
   o. zinc **Zn**
   p. arsenic **As**
   q. lead **Pb**
   r. iron **Fe**
   s. calcium **Ca**
   t. cobalt **Co**

2. Write the name of the element that corresponds to each of the following symbols.

   a. Kr **Krypton**
   b. K **Potassium**
   c. C **Carbon**
   d. Ne **Neon**
   e. Si **Silicon**
   f. Zr **Zirconium**
   g. Sn **Tin**
   h. Pt **Platinum**
   i. Na **Sodium**
   j. Al **Aluminum**
   k. Cu **Copper**
   l. Ag **Silver**
   m. P **Phosphorus**
   n. Mn **Manganese**
   o. I **Iodine**
   p. Au **Gold**
   q. Mg **Magnesium**
   r. Ni **Nickle**
   s. Br **Bromine**
   t. Hg **Mercury**

Elements with Unique Symbols

The symbols for some elements do not match the English name for the element. This is because the symbols for some elements are based on their ancient names, coming from languages other than English, like Latin, Greek, or even German.

Write the element symbol for each of these elements.

a. lead **Pb**
b. mercury **Hg**
c. tin **Sn**
d. copper **Cu**
e. potassium **K**
f. tungsten **W**
g. gold **Au**
h. silver **Ag**
i. iron **Fe**
j. sodium **Na**

Chemistry Matters!

Welcome to chemistry! What is chemistry? **Chemistry** is the study of matter and how it changes. **Matter** is anything that has mass and takes up space. All matter is composed of very small particles called atoms.

PART 1: ELEMENT, COMPOUND, MOLECULE?

An **atom** is the smallest particle that has the characteristics of the element. We will get into more details about atoms (and what they're made of) in a future chem lesson.

An **element** is made up of only one type of atom. There are approximately 118 elements known at this time. These are combined in various ways to make up ALL the matter in the universe. How can you tell if something is an element? **All of the known elements are listed on the periodic table.**

Notice that all the elements are usually abbreviated by their one- or two-letter chemical symbol. The first letter is always capitalized. Here is a list of the elements, by name and symbol: http://en.wikipedia.org/wiki/List_of_elements_by_name (of course, it's also in your textbook!).

The first 36 elements (plus a few others) in the periodic table are fairly common. You should recognize whether something is an element or not. For example, you should learn that Cu is copper and is on the periodic table which tells us it is an element. CuS is a combination of copper and sulfur (more than one element), so it is a compound. You will always be able to use your periodic table so feel free to take your time and look up the element if you're not sure!

✅ **Checkpoint 1:** Name the following elements based on the chemical symbol given. You may use your book or the internet:

a. O **Oxygen**
b. Ca **Calcium**
c. Na **Sodium**
d. Ag **Silver**

✅ **Checkpoint 2**: Give the chemical symbol for the following elements. You may use your book or the internet:

a. neon **Ne**
b. chlorine **Cl**
c. potassium **K**
d. iron **Fe**

Compounds are formed when atoms of different elements undergo a **CHEMICAL** change.

A very important idea in chemistry: Compounds are chemically and physically different than the elements that make them up. For example, water is made up of hydrogen, a flammable gas, and oxygen, a gas that we breathe. Water, however, is very different. Water is a liquid at room temperature, it is denser than air, not flammable, and we cannot breathe it.

Atoms can bond with each other and form molecules. Molecules can be considered elements or compounds. See if you can figure out when a molecule is an element and when it is a compound:

- **Checkpoint 3:** A molecule of oxygen gas is composed of two oxygen atoms bonded together. It is considered a(n) (element or compound)?

- **Checkpoint 4:** A molecule of carbon dioxide gas is composed of one carbon atom and two oxygen atoms bonded together. It is considered a(n) (element or compound)?

- **Checkpoint 5:** What is the chemical difference between Co and CO?

Co is the element cobalt.

CO is the compound carbon monoxide.

PART 2: CHEMICAL FORMULAS

Chemists use special notation called "chemical formulas" to describe elements and compounds. A molecule of oxygen described in Question 3 above is written as O_2. The **subscript** 2 indicates the number of O atoms bonded together.

NOTE: Are you wondering why there are two ways oxygen atoms can make molecules—as O_2 (oxygen gas) or O_3 (ozone). Which is considered an element? They both are! Oxygen gas and ozone are called "allotropes" of oxygen (different forms).

Try answering these questions:

- **Checkpoint 6:** What is the chemical formula of carbon dioxide? **CO_2**

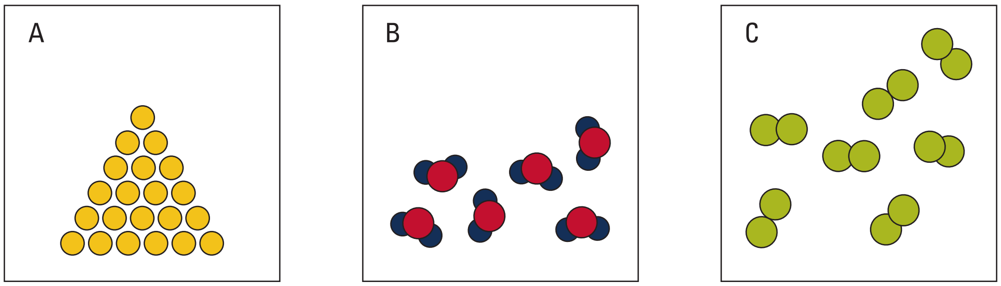
- **Checkpoint 7:** How many atoms in an ozone molecule, O_3? **3**

- **Checkpoint 8:** How many atoms are in $CuSO_4$? **1 atom Cu, 1 atom S, 4 atoms O = 6 atoms total**

- **Checkpoint 9:** How many atoms are in $Ba(OH)_2$? **HINT:** The $(\ )_2$ means that everything inside the ( ) is multiplied by two. **1 atom Ba, 2 atoms O, 2 atoms H = 5 atoms total**

PART 3: MOLECULAR-LEVEL DIAGRAMS

Are you a visual learner? Sometimes it helps to draw pictures of what the atoms and molecules might look like. These are called molecular-level diagrams. In these diagrams, each circle is one atom. The shading of the atom (empty, solid, etc. …) indicates a particular element. Bonded atoms are shown as touching each other. Here are some examples of molecular-level diagrams for terms you just learned above:



Checkpoint 10: Which boxes show molecules? List all that apply.

B and C

Checkpoint 11: Which boxes show elements? List all that apply.

A and C

PART 4: MIXTURES

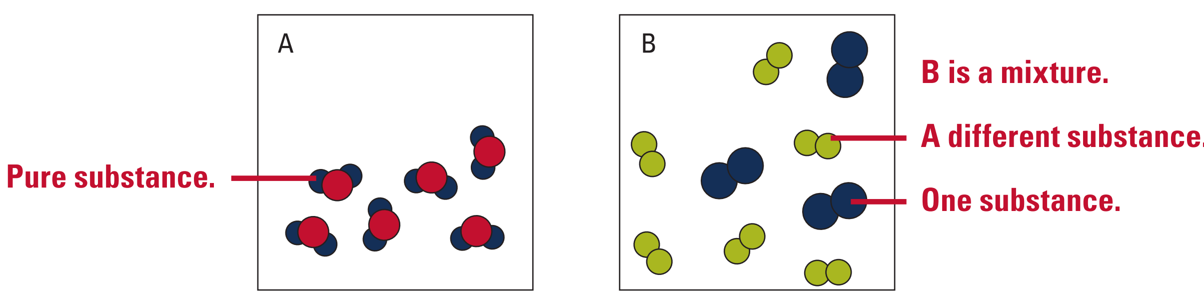
Mixtures contain two or more separate components. Each component retains its properties.

EXAMPLES

Salt water. This mixture is made up of two compounds, water and salt. The water and salt do not undergo a chemical change. How do you know? Because you can separate the water and salt by evaporating off the water.

Sand is a mixture. It contains the components shell, rock, glass, etc.

Checkpoint 12: Which ***one*** of these is considered a mixture, A or B? Explain.



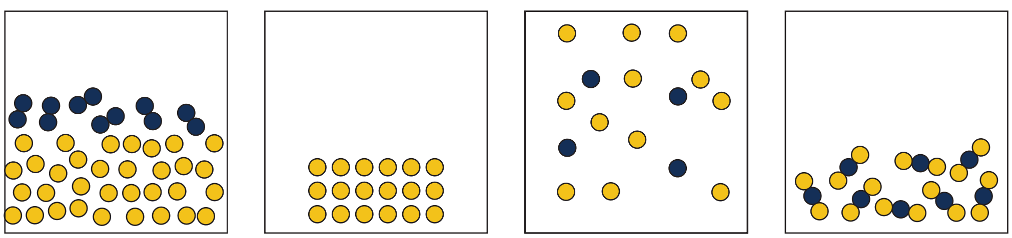
There are two types of mixtures, homogeneous and heterogeneous.[1]

Homogeneous mixtures: The components of the mixture are uniformly distributed throughout the mixture. These are sometimes called solutions. Salt water is an example of a homogeneous solution. Air is also an example of a homogeneous mixture of various gases (oxygen, nitrogen, etc.).

Heterogeneous mixtures: A difference can be seen in the different parts of the mixture. Some examples of heterogeneous mixtures are sand, a chocolate chip cookie, and carbonated soda (you can see the difference between the liquid and the bubbles).

How do you tell the difference between a compound, element, and a mixture? Mixtures can be separated into their components by exploiting differences in physical properties. For example, you can take advantage of a difference in phase by using filtration. You probably do this every morning when you make coffee. The coffee filter separates the solid grounds from the liquid coffee.

- ✓ **Checkpoint 13:** Which of the following is a pure element (no mixtures)? **B**
- ✓ **Checkpoint 14:** Which of the following is a pure compound (no mixtures)? **D**
- ✓ **Checkpoint 15:** Which of the following are mixtures? **A and C**
- ✓ **Checkpoint 16:** Which of the following is a homogeneous mixture? **C (uniformly distributed)**
- ✓ **Checkpoint 17:** Which of the following is a heterogeneous mixture? **A (more than one phase)**



A B C D

IMPORTANT!

Notice in Checkpoint 12 that the **elements** hydrogen and oxygen are drawn in Box B as **diatomic molecules**. There are seven elements that occur as diatomic molecules under standard condition (you should remember this):

Br_2, I_2, N_2, Cl_2, H_2, O_2, and F_2
(Mnemonic: Brinclhof or Hofbrincl)

Another note: H atom is an element; H_2 molecule is also called hydrogen and is an element, not a compound. (Sometimes it's called "dihydrogen.")

1 **SPECIAL NOTE:** It's also possible to have a colloid—a homogenous mixture in which larger particles (> 1 nm) are dispersed in different phases. One example of a colloid is fog, in which water droplets are suspended in air. Colloids display the Tyndall Effect: http://silver-lightning.com/tyndall/

PART 5: PHYSICAL VS. CHEMICAL

Physical properties describe characteristics of matter. Examples are color, odor, melting and boiling points, density, and phase (solid, liquid, gas). In a **physical change** the identity of the chemical remains the same. For example, when ice melts, the chemical is still water. Just the phase has changed. ***NO NEW SUBSTANCES ARE FORMED IN A PHYSICAL CHANGE.***

Chemical properties describe the ability of matter to react or ***BECOME ANOTHER SUBSTANCE.*** An example is flammability (the ability to react with oxygen and burn). Most **chemical changes** involve color changes (as one substance with one color changes into a substance with another color), bubbling as a gas is formed, cloudiness as a precipitate is formed, or the exchange of heat (the release of heat is called an **exothermic reaction,** and the absorption of heat is called an **endothermic reaction**).

Checkpoint 18: Open YouTube and search for the two videos below. For each one, determine if you are observing a physical change or a chemical change.

a. In YouTube, search for **"sublimation of iodine"** video (length is 0:52). Or here's the URL: http://www.youtube.com/watch?v=0_LWBgeQrvk

   Physical change

b. In YouTube, search for **"ammonium dichromate volcano"** video (length is 0:59). Or here's the URL: http://www.youtube.com/watch?v=UIa2NWi3Q34

   Physical change

Checkpoint 19: Are each of the following statements *chemical* or *physical* descriptions?

a. alcohol evaporates easily **Physical**
b. iron rusts over time **Chemical**
c. natural gas is burned for heat **Chemical**
d. salt is crushed into a fine powder **Physical**

Checkpoint 20: Lightning converts O_2 (oxygen we breathe) to O_3 (ozone). Is this a physical or chemical change?

EXERCISES

1. Define each of these terms and give an example from everyday life, using either words or chemical symbols/formulas:

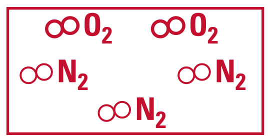
   a. element **Cu**

   b. compound **CO_2**

   c. pure substance **CO_2, Cu**

   d. mixture **Coffee, milk, blood, air**

2. Air is composed of roughly 20% oxygen and 80% nitrogen. Make a molecular-level drawing to represent this statement using 10 atoms or molecules to show the relative proportions. Don't forget the diatomic elements (BrINClHOF)!



3. Is one water molecule considered a pure substance or a mixture? Explain.

Only H_2O is present.

4. Can a glass of water be a pure substance, a mixture, or both? Explain.

Can be distilled or purified H_2O or tapwater—a mixture.

5. If you have a container with hydrogen gas and oxygen gas in it, do you have water? Explain why or why not.

No, it will be a mixture of H_2 and O_2. In order to have water, H_2 and O_2 must combine in a chemical bond to make H_2O.

6. Categorize each of the following as a pure substance, a homogeneous mixture, or a heterogeneous mixture.

   a. white grape juice **Homogeneous mixture**

   b. open bottle of beer **Heterogeneous mixture (bubbles, foam in a liquid)**

   c. soda with ice cubes **Heterogeneous mixture (ice (solid) in liquid)**

   d. fresh garden salad **Heterogeneous mixture**

7. Suppose salt water (consisting of NaCl and water) is found to contain several grains of pure, quartz sand (sand is composed of silicon dioxide, SiO_2). The mixture is filtered and the solid sand placed in a vial labeled **A.** Half the remaining liquid is placed in vial **B.** The other half of the liquid is heated until it boils. The gas that escapes is collected and cooled. The liquid that reforms is placed in vial **C.** The solid that remains after the boiling is placed in vial **D.**

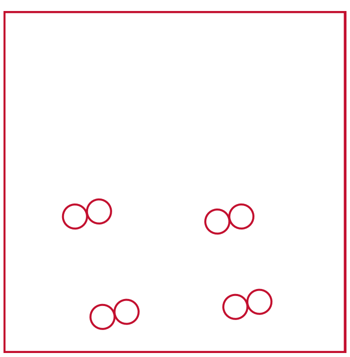
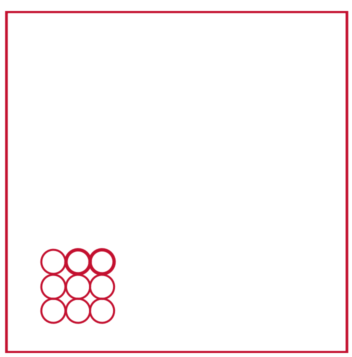
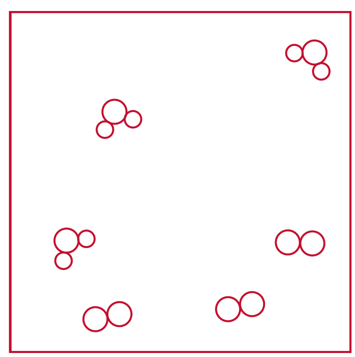
   Use the following four choices to categorize each sample below:

   a. pure element
   b. pure compound
   c. homogeneous mixture
   d. heterogeneous mixture

   i. The original sample **d** **(sand and sea water)**
   ii. The sample in vial **A** **b** **(pure SiO_2, compound)**
   iii. The sample in vial **B** **c** **(sea water)**
   iv. The sample in vial **C** **b** **(pure water)**
   v. The sample in vial **D** **b** **(salt, NaCl)**

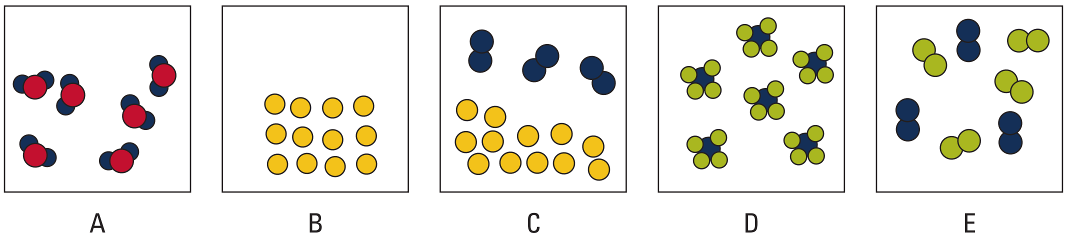
8. In each of the three boxes below, draw the following. Use different shading to show different types of atoms.

   a. Pure substance of molecules (that are compounds) in the liquid phase
   b. Homogeneous mixture of atoms in the solid phase
   c. Heterogeneous mixture of molecules in the gas phase

Molecular Level Diagrams

Refer to these drawings for the questions that follow.



1. Which of these drawings contains elements only (no compounds)? List all. **B, C, E**
2. Which of these drawings represent pure substances? List all. **A, B, D**
3. Which of these drawings contains molecules? List all. **A, C, D, E**
4. Which drawings contain at least two chemical substances? List both. **C, E**
5. Drawing C shows a (circle one):

   pure substance / homogeneous mixture / heterogeneous mixture
6. Which drawing includes two different diatomic molecular elements? One answer: **E**
7. Which drawings contain substances in their gaseous state? List all. **A, C, D, E**
8. Which drawing has only a single monatomic element? **B**
9. Which drawings include a compound? Circle all that apply:

   A / B / C / D / E
10. Which drawing(s) show a mixture of two substances? **C, E**
11. Drawing E shows (circle all that apply):

    two diatomic elements / two molecular elements / a compound
12. Which drawing appears to contain substances in two different physical states? **C**
13. The outer layer of a penny is composed of copper. Which drawing best represents the penny's outer layer? **B**
14. Which drawing best represents a sample of liquid water? **A (H_2O molecules)**
15. Air is composed of gaseous O_2 and N_2 molecules. Which diagram best represents air? **E**

Elements, Compounds & Mixtures

PART 1

Read the following information on elements, compounds and mixtures. Fill in the blanks where provided.

ELEMENTS

- A pure substance contains only one **atom**.
- An element is always uniform all the way through (homogeneous).
- An element **can not** be separated into simpler materials (except during nuclear reactions).
- Over 100 existing elements are listed and classified on the **Periodic table**.

COMPOUNDS

- A compound contains two or more **elements**.
- The atoms are **chemically** combined in some way. Often times (but not always) they come together to form groups of atoms called molecules.
- A compound is always homogeneous (uniform).
- Compounds **can not** be separated by physical means. Separating a compound requires a chemical reaction.
- The properties of a compound are usually **different than** the properties of the elements it contains.

MIXTURES

- In a mixture two or more **elements** or **compounds** are NOT chemically combined.
- There is no reaction between substances.
- Mixtures can be uniform (called **homogeneous** mixtures) and are known as solutions.
- Mixtures can also be non-uniform (called **heterogeneous** mixtures).
- Mixtures **can** be separated into their components by chemical or physical means.
- The properties of a mixture are **similar to** the properties of its components.

PART 2

Classify each of the following as elements (E), compounds (C) or Mixtures (M). Write the letter X if it is none of these.

E Diamond (C)

M Air

E Krypton (K)

C Water (H_2O)

C Ammonia (NH_3)

M Wood

C Dry Ice (CO_2)

C Sugar ($C_6H_{12}O_6$)

C Sulfuric Acid (H_2SO_4)

E Bismuth (Bi)

C Alcohol (CH_3OH)

C Salt (NaCl)

M Bronze

C Baking Soda ($NaHCO_3$)

M Milk

M Gasoline

E Uranium (U)

M Pail of Garbage

X Energy

M Ink

E Titanium (Ti)

E Iron (Fe)

X Electricity

M Popcorn

M A dog

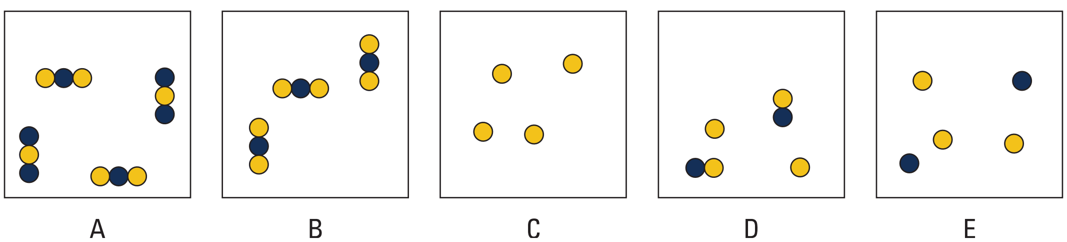
E Gold (Au)

M Pizza

M Concrete

PART 3

Match each diagram with its correct description. Diagrams will be used once.



1. Pure element—only one type of atom present **C**
2. Mixture of two elements—two types of uncombined atoms present **E**
3. Pure compound—only one type of compound present **B**
4. Mixture of two compounds—two types of compounds present **A**
5. Mixture of a compound and an element **D**

PART 4

Identify whether the substance is an element (E), a compound (C), a heterogeneous mixture (HM), or a homogenous mixture (hm).

1. Chocolate chip cookie **HM**
2. Steam **C**
3. Toothpaste **hm**
4. Chicken soup **HM**
5. Tap water **hm**
6. Salt water **hm**
7. Pencil lead (graphite) **hm**
8. A burrito **HM**
9. Lemonade **HM (if particles floating, hm if clear)**
10. Distilled water **C**
11. Dirt **HM**
12. Soda/cola **HM (gas bubbles out)**
13. Italian salad dressing **HM (if differnt particles there)**
14. Pure gold **E**
15. Smoothie **hm**

Scientific Notation: Powers of 10

In the sciences, sizes of atoms and molecules are extremely small compared to the sizes of cells and small organisms, and even smaller than the distance to planets and galaxies. Many zeros in a number can lead to error, so a compact way of indicating size is to use exponents/powers of 10. **Guideline:** Use scientific notation generally when numbers are larger than 10^3 (1000) or smaller than 10^{-3} (0.001).

Here are some examples of how scientific notation is used:

Decimal Notation	Scientific Notation
0.000134	1.34×10^{-4} (coefficient is 1.34)
895000	8.95×10^{5} (coefficient is 8.95)
2100.7	2.1007×10^{3} (coefficient is 2.1007)

1. For numbers smaller than 1, do you move the decimal to the left or the right to convert the number into scientific notation?
2. For numbers larger than 1, do you move the decimal to the left or the right to convert the number into scientific notation?
3. True or False? The coefficient in scientific notation is usually between 1 and 10. **From 1–9**
4. Express the two numbers below in decimal form. Verify that they are the same number.
   a. 450×10^{-9} **0.000,000,450.**
   b. 4.5×10^{-7} **0.000,000,45**
5. Practice converting these numbers into scientific notation:
   a. 0.33 **3.3×10^{-1}**
   b. 45,000 **4.5×10^{4}**
   c. 3,450,000 **3.45×10^{6}**
   d. 0.009708 **9.708×10^{-3}**
6. Practice converting these numbers into decimal form:
   a. 4×10^{2} **400**
   b. 5.3×10^{-3} **0.0053**
   c. 1.01×10^{-5} **0.0000101**
   d. 8.9345×10^{6} **8,934,500**
7. Practice converting these to "standard" scientific notation:
   a. 39×10^{-1} **0.39**
   b. 3859×10^{7} **3.859×10^{10}**
   c. 0.00075×10^{8} **7.5×10^{4}**
   d. 53.498×10^{-34} **5.3498×10^{-33}**

Exponents

How to perform simple mathematical operations on exponents without a calculator:

$10^a \times 10^b = 10^{(a+b)}$

$(10^a)^b = 10^{(a*b)}$

$10^a \div 10^b = 10^{(a-b)}$

$10^0 = 1$

Practice doing these calculations without a calculator. Put your answer in scientific notation when appropriate (smaller than 10^{-4} or larger than 10^4):

a. $(1 \times 10^9) \times (1 \times 10^{-4}) = \mathbf{1 \times 10^5}$

b. $(1 \times 10^4)/(1 \times 10^{-6}) = \mathbf{1 \times 10^{10}}$

c. $(3.50 \times 10^7)^2 = \mathbf{1.225 \times 10^{15}}$

d. $0.45 \times 10^0 = \mathbf{0.45}$

USING A SCIENTIFIC CALCULATOR

Try this calculation on your calculator: $60 \div 2.5 \times 10^4 = \mathbf{0.0024 = 2.4 \times 10^{-3}}$

The answer should be 2.4×10^{-3}. If you are not careful, your calculator will tell you it's 240,000! This type of error tends to occur when you divide by powers of 10.

There are several ways to avoid the wrong answer:

- You can use parentheses: $60 \div (2.5 \times 10\text{^}4)$
- You can use the EE or EXP function on your calculator as a shortcut for "10^." Try 60 ÷ 2.5E4. (Ask your instructor if you have trouble locating these buttons on your calculator!)

Practice on these calculations:

1. $(3.5 \times 10^4) \div (2 \times 10^9) = \mathbf{1.75 \times 10^{-5}}$ (Answer: 1.75×10^{-5} or 1.75E-5)
2. $(4.3 \times 10^2) \times (5 \times 10^{-3}) = \mathbf{2.15}$ (Answer: 2.15)
3. $(3.2 \times 10^4) - (2.0 \times 10^3) = \mathbf{2.15}$ (Answer: 3.0×10^4)
4. $(5.9 \times 10^2) + (1.6 \times 10^3) = \mathbf{2.15}$ (Answer: 2.2×10^3)
5. $(6.4 \times 10^2) \div (3.0 \times 10^{-4}) = \mathbf{2.15}$ (Answer: 2.1×10^6)

Significant Figures

WHAT ARE SIGNIFICANT FIGURES?

In a recorded measurement, the significant figures (sig. figs. or s.f.) include all of the certain digits, plus one last uncertain digit (estimated or guessed digit).

# of sig figs = # certain digits + 1 last uncertain digit

WHY DO WE USE SIGNIFICANT FIGURES?

Significant figures are used as a way of communicating the amount of precision that a recorded measurement has. When you record a measurement to a specific decimal place, you are communicating the precision of the measuring device that was used. If you write too few or too many digits, you are misrepresenting the precision of the measurement, suggesting that the measurement was more approximately made or more finely made than it actually was. Below are examples of volume measurements in which the units are milliliters (mL).

EXAMPLE

3.0 mL The 3 is certain and the 0 is uncertain. There is a total of 2 significant figures (2 s.f.).

Try these:

1) 3.1 mL **2** 2) 3.05 mL **3** 3) 3.00 mL **3** 4) 3.000 mL **4**

ANOTHER EXAMPLE

0.31 mL The 3 is certain and the 1 is uncertain. There is a total of 2 s.f. The 0 is neither, it is a place holder to indicate magnitude.

Now, try these:

5) 0.30 mL **2** 6) 0.300 mL **3** 7) 0.00300 mL **3**

A METHOD FOR COUNTING SIGNIFICANT FIGURES

Reading left to right, start counting at the first non-zero digit.

- If the number has a decimal point, count all the way to the end.
- If there is no decimal point included, stop counting at the last non-zero digit.

Try a few more:

8) 3 mL **1**
9) 31 mL **2**
10) 301 mL **3**
11) 310 mL **2**
12) 30000. mL **5**
13) 30000 mL **1**
14) 30000.00 mL **3**

Measurements and Significant Figures 1

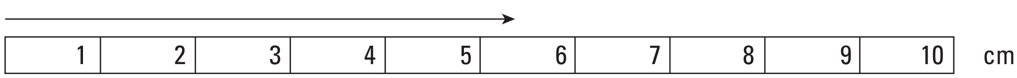
1. Accuracy refers to how close a measurement is to an accepted or true value. Suppose you are a technician measuring a 1 gram weight on four different balances and obtain the following four measurements. Which measurement is from the most accurate balance?

   a. 0.99 g b. 1.20 g c. 1.05 g d. 0.89 g

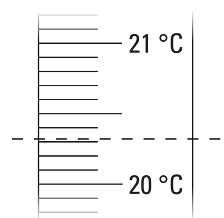
2. Precision refers to the fineness of a measurement and is often indicated by how many digits are given in a measurement. Suppose you are a nurse using three different thermometers. Which measurement is from the highest precision instrument?

   a. 99 °F b. 98.6 °F c. 98.584 °F **Finest measurement (most decimal places)**

3. Measure the arrow using the ruler provided below it using the number of digits allowed by this ruler. **5.2 or 5.3 cm**

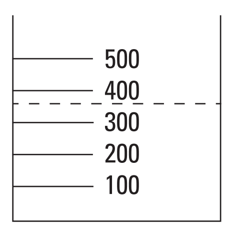


4. Provide a measurement for the temperature displayed on this Celsius thermometer (a portion of the thermometer is shown, the dashed line indicates the temperature). **20.33 °C (increment = 0.1 °C, record to 0.01 °C)**



5. Beaker

   a. Report a measurement for the volume in this beaker.
   360 or 370 mL (increment = 100, record to 10 mL)



   b. Would this beaker be a good choice for measuring a volume of 5.0 mL? **No**

Measurements and Significant Figures 2

In the lab you will practice using various instruments to measure length, volume, or mass. Here are some common measuring tools. Determine the appropriate number of decimal places to record for each measuring tool based on the level of graduation.

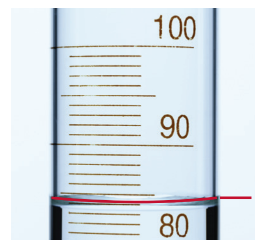
Decimal places:

hundreds	tens	ones	**decimal point**	tenths	hundredths	thousandths
100	10	1	.	1/10	1/100	1/100

0.50 cm

5.0 mm

1. Ruler
   a. In centimeters, record to **2**
   b. In millimeters, record to **1**



84.0 mL

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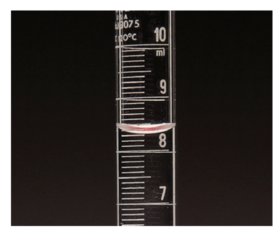
2. 100-mL graduated cylinder
   a. Record to **1**

© Credit: Martyn F. Chillmaid/Science Source

3. Buret
   a. Record to **2**
   b. This reading is about **20.90** mL **or 20.88 mL**

   (**NOTE:** A buret is marked in reverse!)



© Credit: Charles D. Winters/Science Source

4. 10-mL graduated cylinder
   a. Record to **2**
   b. This reading is about **3.68** mL

5. 100-mL beaker
   a. Record to **0**

   No decimal place here.

30
35 mL

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Significant Figures and Scientific Notation

1. How many significant figures are in:
   a. 1.204×10^{-2} g **4**
   b. 3.160×10^{8} Å **4**
   c. 0.00281 g **3**
   d. 810 mL **2**
   e. 12.82 Liters **4**
   f. 3.19×10^{15} atoms **3**
   g. 4.300×10^{-6} cm **4**
   h. 0.00641 g **3**
   i. 8.2354×10^{-19} m **5**
   j. 0.0559 g **3**
   k. 2.92×10^{2} g **2**
   l. 4.1 Liters **2**
   m. 0.0002 cm **1**
   n. 0.045 g **2**

2. Convert the standard (*decimal*) notation numbers above into standard *scientific* notation. Remember to carry all significant figures into the coefficient.

3. Assign ± error readings to the following measurements (**HINT:** Which digit has uncertainty?) Assume units of 1 (0.1, 0.01, etc.)

   a. 3.412 g **± 0.001 g**
   b. 45 mL **± 1 mL**
   c. 0.00498 g **± 0.00001 g**
   d. 8.2 cm **± 0.1 cm**
   e. 559 L **± 1 L**
   f. 1.00×10^{2} m **$\pm 0.01 \times 10^{2}$ m or 100 m ± 1 m**

Significant Figures in Calculations

ROUNDING

If the digit to the right of the final digit you want to keep is

- less than 5, round down.
- 5 or more, round up.

EXAMPLES

Round 0.067556 to 2 sig figs.   Answer: 0.068

Round 3.145×10^3 to 3 sig figs.   Answer: 3.15×10^3

MULTIPLICATION/DIVISION

Round your calculated answer to have the same number of significant figures as the least precise measurement (fewest number of s.f.).

EXAMPLES

$5.02 \times 89.665 \times 0.10 = 45.00681 = 45$ (3 s.f. × 5 s.f. × 2 s.f. ⟶ round to 2 s.f.)

5.892/6.10 = **0.965901639 = 0.966**

ADDITION/SUBTRACTION

Round your calculated answer to have the same number of decimal places as the least precise measurement.

EXAMPLES

5.74 + 0.823 + 2.651 = 9.214 = 9.21 (rounded to 1/100s place, results in 3 s.f.)

0.987 + 125.1 – 1.22 = **124.9**

MULTI-STEP/COMBINATION CALCULATIONS

Keep track at each step based on multiply/divide or add/subtract, but only round at the very end!

EXAMPLE

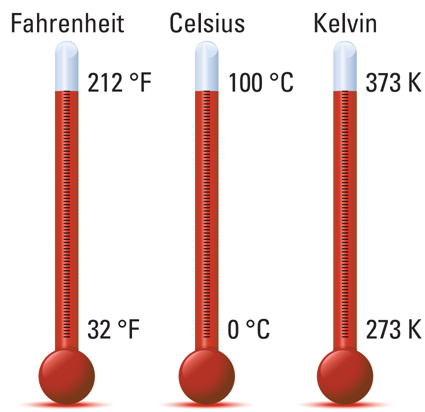
$6.78 \times 5.903 \times (5.489 - 5.01) = 19.17070086 = 19$

(1/100 place for subtraction step results in 2 s.f, then in multiply step 2 s.f. is fewest → round to 2 s.f.)

$19.667 - (5.4 \times 0.916) =$ **$14.\overline{7}206 = 14.7$**
$19.66\overline{7} - 4.\overline{9}464$

PRACTICE

1. Use the equation for converting 545.0 kelvins to °C (degree Celsius) and round using significant figures. °C = K + 273.15 = **545.0 + 273.15 = 818.2 K**



2. Use the equation for °C/°F to convert body temperature, 37.0 °C to units of °F.

   (°C * 1.8) + 32 = °F [**NOTE:** The 1.8 and 32 are considered exact numbers.]

   (37.0 °C × 1.8) + 32 = 66.6 + 32 = 98.6 °F

3. Convert 114°F to °C. (Answer: 45.5°C)

4. Convert –10.0°C to °F and Kelvin. (Answer: 14°F, 263 K)

List of Metric Prefixes

Prefix	Symbol	Mathematical Meaning	In Words ...
Yotta-	Y	10^{24}	
Zetta-	Z	10^{21}	
Exa-	E	10^{18}	
Peta-	P	10^{15}	
Tera-	**T**	$\mathbf{10^{12}}$	**trillion**
Giga-	**G**	$\mathbf{10^{9}}$	**billion**
Mega-	**M**	$\mathbf{10^{6}}$	**million**
Kilo-	**k**	$\mathbf{10^{3}}$	**thousand**
Hecto-	h	10^{2}	hundred
Deca-	da	10^{1}	ten
		$1 = 10^{0}$	one
Deci-	**d**	10^{-1}	tenth
Centi-	**c**	$\mathbf{10^{-2}}$	**hundredth**
Milli-	**m**	$\mathbf{10^{-3}}$	**thousandth**
Micro-	**μ**	$\mathbf{10^{-6}}$	**millionth**
Nano-	**n**	$\mathbf{10^{-9}}$	**billionth**
Pico-	**p**	$\mathbf{10^{-12}}$	**trillionth**
Femto-	f	10^{-15}	
Atto-	a	10^{-18}	
Zepto-	z	10^{-21}	
Yocto-	y	10^{-24}	

Memorize the prefixes, symbols, and mathematical meanings in **bold** (tera, giga, mega, kilo, deci, centi, milli, micro, nano, pico).

COMMONLY USED BASE UNITS IN THE METRIC SYSTEM

Dimension	Unit
Mass	grams (g)
Length	meters (m)
Volume	liters (L)
Time	seconds (s)

Metric System
Units & Definitions

1. Fill in the following blanks for units of length.

a.	1 terameter (Tm) =	10^{12} m		or	1 m =	10^{-12} Tm	
b.	1 gigameter (Gm) =	**10^{12}**	m	or	1 m =	**10^{-12}**	Gm
c.	1 megameter (Mm) =	**10^{6}**	m	or	1 m =	**10^{-6}**	Mm
d.	1 kilometer (km) =	**10^{3}**	m	or	1 m =	**10^{-3}**	km
e.	1 decameter (dam) =	**10^{1}**	m	or	1 m =	**10^{-1}**	dam
f.	1 centimeter (cm) =	**10^{-2}**	m	or	1 m =	**10^{2}**	cm
g.	1 decimeter (dm) =	**10^{-1}**	m	or	1 m =	**10^{1}**	dm
h.	1 millimeter (mm) =	**10^{-3}**	m	or	1 m =	**10^{3}**	mm
i.	1 micrometer (μm) =	**10^{-6}**	m	or	1 m =	**10^{6}**	μm
j.	1 nanometer (nm) =	**10^{-9}**	m	or	1 m =	**10^{9}**	nm
k.	1 picometer (pm) =	**10^{-12}**	m	or	1 m=	**10^{12}**	pm

2. Give the full names of each of the following units:

a. mL **milliliter**　　c. mg **milligram**

b. Ds ______________________　　d. cL **centiliter**

3. Write symbols for each of the following units:

a. liter **L**　　c. kilosecond **ks**　　e. microliter **μL**

b. decimeter **dm**　　d. megameter **Mm**　　f. milligram **mg**

4. Express each of the following in terms of their base units (e.g., 1 mL = 0.001 L)

a. 1 cm **0.01 m**　　b. 1 mg **0.001 g**　　c. 1 km **1,000 m**　　d. 1 μL **0.000001 m**

5. Write the correct prefixes for the following measurements:

a. 1000 g = **1 kg**　　c. 1,000,000 s = **1 Ms**　　e. 0.0046 g = **4.6 mg**

b. 0.001 m = **1 mm**　　d. 0.01 g = **1 cg**　　f. 3300 s = **3.3 ks**

6. Which is the larger unit?

a. centimeter or decimeter　　c. milligram or megagram

b. milliliter or microliter　　d. deciliter or decaliter

7. Complete the following blanks:

a. 1 L = $\mathbf{10^{3}}$ mL

b. 1 km = $\mathbf{10^{3}}$ m

c. 1 mg = $\mathbf{10^{-3}}$ g

d. 1 μL = $\mathbf{10^{-6}}$ L

e. 1 cm = $\mathbf{10^{-2}}$ m

f. 1 g = $\mathbf{10^{12}}$ μg

g. 1 dL = $\mathbf{10^{-1}}$ L

h. 1 μm = $\mathbf{10^{-6}}$ m

i. 1 Tg = $\mathbf{10^{12}}$ g

j. 1 m = $\mathbf{10^{2}}$ cm

k. 1 mL = $\mathbf{10^{-3}}$ L

l. 1 L = $\mathbf{10^{6}}$ μL

m. 1 mm = $\mathbf{10^{-3}}$ m

n. 1 ML = $\mathbf{10^{6}}$ L

o. 1 L = $\mathbf{10^{-12}}$ μL

p. 1 m = $\mathbf{10^{-1}}$ dam

q. 1 mL = $\mathbf{10^{-3}}$ L

r. 1 L = $\mathbf{10^{-9}}$ GL

Two Methods for Metric Conversions

There are several ways to perform metric conversions. Choose the method that works best for you. For example, let's convert 0.000500 meters to micrometers.

Write the relationship between the base unit and the prefixed unit.

The micro (μ) prefix mathematically means 10^{-6}. The equality between the base unit meter and micrometer is 1 μm = 1×10^{-6} m, since μ means 10^{-6} (you are replacing the μ symbol with 10^{-6}). Or, you can write this as 1 m = 10^{6} μm if you think in terms of 1 of the larger unit equals a lot of the smaller unit.

$$\mathbf{1\ \mu m = 1 \times 10^{-6}\ m \text{ or } 1\ m = 10^{6}\ \mu m}$$

METHOD #1: USING DIMENSIONAL ANALYSIS

POSSIBLE CONVERSION FACTORS

A conversion factor is a relationship between two units written as a fraction. The relationship between micrometers and meters can be written two different ways:

$$\frac{1\ \mu m}{10^{-6}\ m} \text{ or } \frac{10^{-6}\ m}{1\ \mu m}$$

SETTING UP THE DIMENSIONAL ANALYSIS CALCULATION

Start with what you are trying to convert, the given information, and write that down. Choose which version of the conversion factor to use so that the units you are trying to convert cancel out.

$$0.0005000\ m \times \frac{1\ \mu m}{10^{-6}\ m} = 0.0005000 \times 10^{6} = 5.00 \times 10^{2}\ \mu m$$

METHOD #2: MOVING THE DECIMAL POINT

For this method, it is helpful to use a drawing of the metric prefix scale.

M			k			Base Unit	d	c	m			μ	
10^{6}	|	|	10^{3}	|	|	|	10^{-1}	10^{-2}	10^{-3}	|	|	10^{-6}	|

Find the base unit (marked as –) then count spaces over to the μ symbol. There are six spaces, so there is a difference of 10^{6} in scale. Since micrometers are smaller than meters, converting from meters to micrometers results in a more of those smaller units, so the decimal point moves to the right six times. (0.000500 becomes 500.)

$$\mathbf{0.000500\ m = 500\ \mu m = 5.00 \times 10^{2}\ \mu m}$$

Metric Conversions: Dimensional Analysis

1. Fill in the number which represents the meaning of each prefix. Use scientific notation (or exponents) when appropriate, usually for numbers smaller than 0.001 and larger than 1000.

   a. m 10^{-3} d. c 10^{-2} g. d 10^{-1}

   b. k 10^{3} e. µ 10^{-6} h. da 10^{1} or 10

   c. M 10^{6} f. n 10^{-9} i. G 10^{9}

2. Complete each conversion as a fraction with the correct number.

   a. $\dfrac{1\ \text{cm}}{10^{-2}\ \text{m}}$ e. $\dfrac{1\ \text{mL}}{10^{-3}\ \text{L}}$

   b. $\dfrac{1\ \text{m}}{10^{2}\ \text{cm}}$ f. $\dfrac{1\ \text{L}}{1000\ \text{mL}}$

   c. $\dfrac{1\ \text{kg}}{10^{3}\ \text{g}}$ g. $\dfrac{1\ \text{ns}}{10^{-9}\ \text{s}}$

   d. $\dfrac{10^{3}\ \text{g}}{1\ \text{kg}}$ h. $\dfrac{10^{9}\ \text{ns}}{1\text{s}}$

3. Can you flip each conversion? Write these next to the conversions above. Check your answers with others.

4. Using any of the conversions above, fill in the blanks below and solve:

   a. $30\ \text{cm} \times \left(\dfrac{1\ \text{m}}{10^{2}\ \text{cm}}\right) = 0.3\ \text{m}$

   c. $150\ \text{g} \times \left(\dfrac{1\ \text{kg}}{10^{3}\ \text{g}}\right) = 0.150\ \text{kg}$

   b. $45\ \text{m} \times \left(\dfrac{1\ \text{cm}}{10^{-2}\ \text{m}}\right) = 4500\ \text{cm}$

   d. $10\ \text{mL} \times \left(\dfrac{1\ \text{L}}{1000\ \text{mL}}\right) = 0.01\ \text{L}$

   e. $0.005\ \text{ks} \times \left(\dfrac{10^{3}\ \text{s}}{1\ \text{ks}}\right) \times \left(\dfrac{1\ \text{ns}}{10^{-9}\ \text{s}}\right) = 5 \times 10^{9}\ \text{ns}$

5. Now try these conversions from scratch. Use the set up provided above and cancel out your units.

   a. 900 g to kg

   $\dfrac{900\ \text{g}}{} \Big| \dfrac{1\ \text{kg}}{10^{3}\ \text{g}} = 0.9\ \text{kg}$

   c. 800 mm to m

   $\dfrac{800\ \text{mm}}{} \Big| \dfrac{\times 10^{-3}\ \text{m}}{1\ \text{mm}} = 0.8\ \text{m}$

   e. 31 mm to nm

   $\dfrac{31\ \text{mm}}{} \Big| \dfrac{10^{-3}\ \text{m}}{1\ \text{mm}} \Big| \dfrac{1\ \text{nm}}{10^{-9}\ \text{m}}$

   $= 31 \times 10^{6}\ \text{nm}$ or $3.1 \times 10^{7}\ \text{nm}$

   b. 3.4 cm to m

   $\dfrac{3.4\ \text{cm}}{} \Big| \dfrac{\times 10^{-2}\ \text{m}}{1\ \text{cm}} = 0.034\ \text{m}$

   d. 0.0450 kg to cg

   $\dfrac{0.0450\ \text{km}}{} \Big| \dfrac{10^{3}\ \text{m}}{1\ \text{km}} \Big| \dfrac{1\ \text{cm}}{10^{-2}\ \text{m}}$

   $= 0.0450 \times 10^{5}\ \text{cm}$

   or $4.50 \times 10^{3}\ \text{cm}$

   f. 2.87 kL to µL

   $\dfrac{2.87\ \text{kL}}{} \Big| \dfrac{10^{3}\ \text{L}}{1\ \text{kL}} \Big| \dfrac{1\ \mu\text{L}}{10^{-6}\ \text{L}}$

   $= 2.87 \times 10^{9}\ \mu\text{L}$

One-Step Metric Conversions Using Dimensional Analysis

Perform the following metric system conversions that involve converting between a base unit and a prefixed unit. Put all answers in both standard and scientific notation. Use significant figures and use dimensional analysis to show your work.

1. 2.14 kg to grams

$$? \text{ g} = \begin{array}{c|c} 2.14 \text{ kg} & 10^3 \text{ g} \\ \hline & 1 \text{ kg} \end{array} = 2.14 \times 10^3 \text{ g}$$

2. 6172 mm to meters

$$? \text{ m} = \begin{array}{c|c} 6172 \text{ mm} & 10^{-3} \text{ m} \\ \hline & 1 \text{ mm} \end{array} = 6.172 \text{ m}$$

3. 0.0256 L to milliliters

$$? \text{ mL} = \begin{array}{c|c} 0.0256 \text{ L} & 1000 \text{ mL} \\ \hline & 1 \text{ L} \end{array} = 25.6 \text{ mL}$$

4. 3.14×10^{-5} g to micrograms

$$? \text{ µg} = \begin{array}{c|c} 3.14 \times 10^{-5} \text{ g} & 1 \text{ µg} \\ \hline & 10^{-6} \text{ g} \end{array} = 31.4 \text{ µg}$$

5. 1.01×10^{-8} s to deciseconds

$$? \text{ ns} = \begin{array}{c|c} 1.01 \times 10^{-8} \text{ s} & 1 \text{ ns} \\ \hline & 10^{-9} \text{ s} \end{array} = 10.1 \text{ ns}$$

6. How many milliliters are in 5 liters?

$$? \text{ mL} = \begin{array}{c|c} 5 \text{ L} & 1000 \text{ mL} \\ \hline & 1 \text{ L} \end{array} = 5000 \text{ mL}$$

7. How many kilograms are in 598 grams?

$$? \text{ kg} = \begin{array}{c|c} 598 \text{ g} & 1 \text{ kg} \\ \hline & 10^3 \text{ g} \end{array} = 0.598 \text{ kg}$$

8. How many centiseconds are in 12 seconds?

$$? \text{ ps} = \begin{array}{c|c} 12 \text{ s} & 1 \text{ ps} \\ \hline & 10^{-12} \text{ s} \end{array} = 12 \times 10^{12} \text{ ps} \quad \text{or} \quad 1.2 \times 10^{13} \text{ ps}$$

Two-Step Metric Conversions Using Dimensional Analysis

Perform the following metric system conversions that involve converting between two prefixed units.

Use the relationships between each prefixed unit and the base unit to generate two conversion factors that will allow you to convert in a two-step process from the given prefixed unit to the base unit and then to the desired prefixed unit.

Put all answers in both standard and scientific notation. Use significant figures and use dimensional analysis to show your work.

EXAMPLE

How many centimeters are in 0.5 kilometers?

$$0.5\ \text{km} \times \frac{10^3\ \text{m}}{1\ \text{km}} \times \frac{1\ \text{cm}}{10^{-2}\ \text{m}} = 0.5 \times 10^5\ \text{cm} = 5 \times 10^4\ \text{cm (or 50,000 cm)}$$

1. How many millimeters are in 29 micrometers?

$$?\ \text{mm} = \frac{29\ \mu\text{m}}{} \left| \frac{10^{-6}\ \text{m}}{1\ \mu\text{m}} \right| \frac{1\ \text{mm}}{10^{-3}\ \text{m}} = 2.9 \times 10^{-8}\ \text{mm}$$

2. How many kilograms are in 100 gigagrams?

$$?\ \text{kg} = \frac{100\ \text{Gg}}{} \left| \frac{10^9\ \text{g}}{1\ \text{Gg}} \right| \frac{1\ \text{kg}}{10^3\ \text{g}} = 100 \times 10^6\ \text{kg} = 1.00 \times 10^8\ \text{kg}$$

3. Convert 1316 mg to kilograms

$$?\ \text{kg} = \frac{1316\ \text{mg}}{} \left| \frac{10^{-3}\ \text{g}}{1\ \text{mg}} \right| \frac{1\ \text{kg}}{10^3\ \text{g}} = 1316 \times 10^{-6}\ \text{kg or } 1.316 \times 10^{-3}\ \text{kg}$$

4. Convert 5.88 mL to deciliters

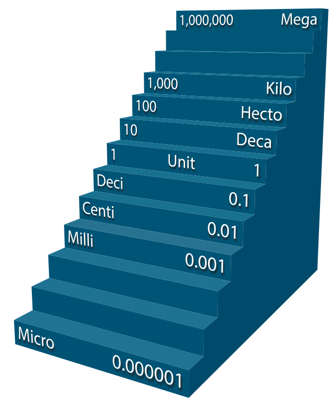
$$?\ \text{mL} = \frac{5.88\ \text{mL}}{} \left| \frac{10^{-3}\ \text{L}}{1\ \text{mL}} \right| \frac{1\ \text{nL}}{10^{-9}\ \text{L}} = 5.88 \times 10^6\ \text{nL}$$

5. Convert 2.19 cm to micrometers

$$?\ \mu\text{m} = \frac{2.19\ \text{pm}}{} \left| \frac{10^{-12}\ \text{m}}{1\ \text{pm}} \right| \frac{1\ \mu\text{m}}{10^{-6}\ \text{m}} = 2.19 \times 10^{-6}\ \mu\text{m}$$

Metric System: The Ladder Method

Count the number of “jumps” from start to end. Move the decimal that number of jumps in the correct direction.



EXAMPLES

35 milligrams (mg) = **0.035** grams (g)

5 kilometers (km) = **5000** meters (m)

1. Write the correct abbreviation for each metric unit. Identify the base unit in each.

		Abbreviation	Base Unit			Abbreviation	Base Unit
a.	kilogram	**kg**	**g**	f.	centimeter	**cm**	**m**
b.	milliliter	**mL**	**L**	g.	gram	**g**	**g**
c.	decimeter	**dm**	**m**	h.	liter	**L**	**L**
d.	meter	**m**	**m**	i.	decagram	**dg**	**g**
e.	megameter	**Mm**	**m**	j.	microliter	**µl**	**L**

2. Try these conversions using the ladder method. Use scientific notation for numbers bigger than 1000 and smaller than 0.001.

   a. 2000 mg = **2** g
   b. 5 L = **5000** mL
   c. 16 cm = **160** mm
   d. 104 km = **104,000** m
   e. 2500 s = **2,500 × 10^9 or 2.5 × 10^{12}** ns
   f. 480 cm = **0.48** dm
   g. 8 mm = **0.8** cm
   h. 65 daL = **65,000** mL
   i. 198 g = **0.198** kg
   j. 120 mg = **0.120** g

3. Complete these equalities or inequalities with the symbols =, > (greater than), or < (less than).

   a. 63 cm **<** 6 m
   b. 5 g **>** 508 mg
   c. 536 cm **>** 53.4 dm
   d. 1500 mL **=** 1.5 L
   e. 500 mg **<** 5 g
   f. 3.6 ns **>** 360 ps
   g. 4×10^{-7} m **=** 400 nm
   h. 3.5×10^{-3} mm **=** 3.5 µm
   i. 100 pL **<** 10^{-6} mL
   j. 500 Gb **<** 5 Tb

Metric Conversions

Use significant figures. Practice with dimensional analysis or the ladder method.

Use scientific notation for numbers larger than 10^4 and smaller than 10^{-4}.

1) 1 mm = __________ m

2) 1 km = __________ m

3) 1 m = __________ cm

4) 1 m = __________ km

5) 1 cm = __________ µm

6) 5.0 mm = __________ m

7) 15 km = __________ cm

8) 6.00×10^2 m = __________ cm

9) 0.45 km = __________ m

10) 104 km = __________ µm

11) 1×10^9 Mm = __________ cm

12) 6×10^{-4} cm = __________ nm

13) 1 µm = __________ km

14) 4.9×10^{-10} mm = __________ km

15) 8×10^{-100} mm = __________ nm

Answers: (1) 0.001 m (2) 1000 m (3) 100 cm (4) 0.001 km (5) 1×10^4 µm (6) 5.0×10^{-3} m (7) 1.5×10^6 cm (8) 6.00×10^4 cm (9) 450 m (10) 1.04×10^{11} µm (11) 1×10^{17} cm (12) 6×10^3 nm (13) 1×10^{-9} km (14) 4.9×10^{-16} km (15) 8×10^{-94} nm

Multi-Step Conversions

1. For the following multi-step conversion problem, use the provided unit relationships to fill in the conversion factors in the unit analysis set up.

   A gas can holds 2.0 gallons of gasoline. What is this quantity in teaspoons?

   1.057 quarts = 1 liter

   4 quarts = 1 gallon

   1 teaspoon = 4.93 milliliters

   1,000 milliliter = 1 liter (you should know this!)

$$\textbf{20 gallons} \times \frac{\text{4 qts}}{\text{1 gal}} \times \frac{\text{1 liter}}{\text{1.057 qts}} \times \frac{\text{1000 mL}}{\text{1 L}} \times \frac{\text{1 tsp}}{\text{4.93 mL}} = \frac{\text{1535 tsp}}{\text{1500 tsp (2 s.f.)}}$$

2. A swimming pool is 50 meters long. How many feet is this? (1 in = 2.54 cm, exactly) (12 in = 1 ft) (Answer: 200 ft)

$$?\text{ ft} = \frac{\text{50 m}}{} \cdot \frac{\text{100 cm}}{\text{1 m}} \cdot \frac{\text{1 in}}{\text{2.54 cm}} \cdot \frac{\text{1 ft}}{\text{12 in}} = \text{164 ft} = \text{200 ft (1 s.f.)}$$

3. It has been reported that the distance from Los Angeles to New York City is 177,000,000 inches. What is this distance in miles? (Check unit conversions listed on the inside front cover.) (Answer: 2.79×10^3 m)

$$?\text{ mi} = \frac{1.77 \times 10^8\text{ in}}{} \cdot \frac{\text{2.54 cm}}{\text{1 in}} \cdot \frac{\text{1 m}}{\text{100 cm}} \cdot \frac{\text{1 km}}{10^3\text{ m}} \cdot \frac{\text{1 mile}}{\text{1.609 km}} = \text{2794. miles (3 s.f.)} = 2.79 \times 10^3\text{ mi}$$

4. A unit of length called the "furlong" is used in horse racing. The units of length called the "chain" and the "link" are used in surveying. What is the length of 1 link (in inches) if: There are 8 furlongs in 1 mile (mi), 10 chains in 1 furlong, and 100 links in 1 chain. Report your answer to three significant figures. (Answer: 7.92 in)

$$?\text{ in} = \frac{\text{1 link}}{} \cdot \frac{\text{1 chain}}{\text{100 links}} \cdot \frac{\text{1 furlong}}{\text{10 chains}} \cdot \frac{\text{1 mile}}{\text{8 furlong}} \cdot \frac{\text{5280 ft}}{\text{1 mile}} \cdot \frac{\text{12 in}}{\text{1 ft}} = \text{7.92 in (3 s.f.)}$$

5. Use these relationships for the problems that follow.

1 nautical mile = 6076.11549 feet	1 degree = 69.047 miles
1 inch = 2.54 cm	1 mile = 5280 feet
3 feet = 1 yard	1 hand = 4 inches
1 league = 5280 yards	1 foot = 12 inches
1 cable = 120 fathoms	1.609 km = 1 mile
1 fathom = 6 feet	

a. Your cruise ship is leaving for a 610-league adventure. How many nautical miles is this? (Answer: 1600 na mi)

$$\text{? na. miles} = \frac{610\text{ league}}{} \left|\frac{5280\text{ yds}}{1\text{ league}}\right| \frac{3\text{ ft}}{1\text{ yd}} \left|\frac{1\text{ na. mile}}{6076.11549\text{ ft}}\right. = 1590 \approx 1600\text{ na. miles } (2\text{ s.f.})$$

b. Later the ship is discovered at 38 fathoms deep under water. Convert this to meters. (Answer: 69 m)

$$\text{? meters} = \frac{38\text{ fathoms}}{} \left|\frac{6\text{ ft}}{1\text{ fathom}}\right| \frac{12\text{ in}}{1\text{ ft}} \left|\frac{2.54\text{ cm}}{1\text{ in}}\right| \frac{1\text{ m}}{100\text{ cm}} = 69.49 \approx 69\text{ m } (2\text{ s.f.})$$

c. Fortunately, you survived! You are stranded on a deserted island that is located 12.5 degrees north of the equator. How many kilometers is this? (Answer: 1.39×10^3 km)

$$\text{? km} = \frac{12.5\text{ degree}}{} \left|\frac{69.047\text{ mi}}{1\text{ degree}}\right| \frac{1.609\text{ km}}{1\text{ mi}} = 1388\text{ km} = 1.3\text{g} \times 10^3\text{ km}$$

d. To reach the top of a palm tree for a coconut, you will have to climb 7.4 meters. How many hands is this? (Answer: 73 hands)

$$\text{? hands} = \frac{7.4\text{ m}}{} \left|\frac{100\text{ cm}}{1\text{ m}}\right| \frac{1\text{ in}}{2.54\text{ cm}} \left|\frac{1\text{ hand}}{4\text{ in}}\right. = 72.8 \approx 73\text{ hands } (2\text{ s.f.})$$

Volume Conversions

LIQUID-BASED VOLUME UNITS RELATED TO CUBED LENGTH VOLUME UNITS

How many cubic millimeters (cm^3) are in a 2-L bottle of soda? At first glance, it may seem like the units are not related because the unit "cm^3" does not contain "L" and instead contains "cm" which is a length, not a volume. But, the unit of cm^3 is indeed a volume and does relate to a volume in liters.

The relationship between a liter-based volume and a cubed-length-based volume is a defined equality.

$$1\ \text{cm}^3 = 1\ \text{mL} \quad \text{or} \quad 1000\ \text{cm}^3 = 1\ \text{L} \quad \text{or} \quad 1\ \text{m}^3 = 10^6\ \text{mL}$$

The most commonly memorized of these is

$$1\ \text{cm}^3 = 1\ \text{mL}$$

PRACTICE

1. How many milliliters are in 853 cm^3?

$$?\ \text{mL} = \begin{array}{c|c} 853\ \text{cm}^3 & 1\ \text{mL} \\ \hline & 1\ \text{cm}^3 \end{array} = 853\ \text{mL}$$

2. How many cubic centimeters are in 3597 mL?

$$?\ \text{cm}^3 = \begin{array}{c|c} 3597\ \text{mL} & 1\ \text{cm}^3 \\ \hline & 1\ \text{mL} \end{array} = 3597\ \text{cm}^3$$

3. Convert 0.020 L to cm^3.

$$?\ \text{cm}^3 = \begin{array}{c|c|c} 0.020\ \text{L} & 1000\ \text{mL} & 1\ \text{cm}^3 \\ \hline & 1\ \text{L} & 1\ \text{mL} \end{array} = 20\ \text{cm}^3$$

4. How many cubic centimeters are in 8500 gallons? (Answer: $3.2 \times 10^7\ \text{cm}^3$)

$$?\ \text{cm}^3 = \begin{array}{c|c|c|c} 8500\ \text{gal} & 4\ \text{qt} & 0.946\ \text{L} & 1000\ \text{cm}^3 \\ \hline & 1\ \text{gal} & 1\ \text{qt} & 1\ \text{L} \end{array} = 3.2 \times 10^7\ \text{cm}^3$$

5. How many cubic meters are in 8500 gallons? (Answer: $32\ \text{m}^3$)

$$?\ \text{m}^3 = \begin{array}{c|c} 3.2 \times 10^7\ \text{cm}^3 & \left(\dfrac{1\ \text{m}}{10^2\ \text{cm}}\right)^3 \\ \hline & \end{array} = 32\ \text{m}^3$$

Ratios as Conversion Factors

When you see a ratio given with different units, it can be written as a conversion factor.

For example, the density of ethanol is **0.789 grams per milliliter (same as 0.789 g/mL).** If you are using it to convert grams to mL (or vice versa), then write it as a fraction:

$$\frac{0.789\text{ g}}{1\text{ mL}} \text{ or } \frac{1\text{ mL}}{0.789\text{ g}}$$

Think about the chain of conversions you'll need to do, linking the units together: If you have 4.00 kg of ethanol, and you want to find out how many liters of ethanol you have, provided the density is 0.789 grams/mL, what conversions do you need?

Unit Map: kg g mL L

Set it up here: $\frac{4.00\text{ kg}}{} \left| \frac{1000\text{ g}}{1\text{ kg}} \right| \frac{1\text{ mL}}{0.789\text{ g}} \left| \frac{1\text{ L}}{1000\text{ mL}} \right. = 5.07\text{ L (3 s.f.)}$

Use dimensional analysis to solve all of these problems. Use significant figures and unit abbreviations.

1. The density of gasoline is 0.70 kg/L.

   a. Write this as a fraction. $\frac{0.70\text{ kg}}{1\text{ L}}$ or $\frac{1\text{ L}}{0.70\text{ kg}}$

   b. What is the mass in kilograms of 15.6 liters of gasoline?

   $$?\text{ kg} = \frac{15.6\text{ L}}{} \left| \frac{0.70\text{ kg}}{1\text{ L}} \right. = 10.9\text{ kg}$$

2. When one gram of liquid ethanol, C_2H_5OH, is burned, 29.7 kJ of heat are released.

   a. Write this as a fraction. $\frac{29.7\text{ kJ}}{1\text{ g}}$ or $\frac{1\text{ g}}{29.7\text{ kJ}}$

   b. How much heat in kilojoules is released when 4.274 grams of liquid ethanol are burned?

   $$?\text{ kJ} = \frac{4.274\text{ g}}{} \left| \frac{29.7\text{ kJ}}{1\text{ g}} \right. = 126.9 \approx 127\text{ kJ (3 s.f.)}$$

3. A sample of iron has a mass of 242.6 grams. What is the volume in cubic centimeters of this iron? Iron has a density of 7.86 g/cm^3.

   a. Write out a unit map to solve this problem. $\frac{7.86\text{ g}}{1\text{ cm}^3}$ or $\frac{1\text{ cm}^3}{7.86\text{ g}}$

   b. Solve the problem.

   $$?\text{ V cm}^3 = \frac{242.6\text{ g}}{} \left| \frac{1\text{ cm}^3}{7.86\text{ g}} \right. = 30.9\text{ cm}^3\text{ (3 s.f.)}$$

4. Suppose an adult male has about 11 kg of fat. Each gram of fat can provide the body with about 38 kJ of energy. If this person requires 8.0×10^3 kJ of energy per day to survive, how many days could he survive on his fat alone? **kg ⟶ g ⟶ kJ ⟶ days** (Answer: 52 days)

$$\textbf{? days} = \begin{array}{c|c|c|c} 11\text{ kg} & 1000\text{ g} & 38\text{ kJ} & 1\text{ day} \\ \hline & 1\text{ kg} & 1\text{ g} & 8.0 \times 10^3\text{ kJ} \end{array} = 52\text{ days } (2\text{ s.f.})$$

5. A doctor orders Medrol to be given 1.5 mg/kg of body weight. Medrol is an anti-inflammatory administered as an intramuscular injection. If a child weighs 72.6 pounds and the available stock of Medrol is 20. mg/mL, how many milliliters does the doctor prescribe? **(1 kg = 2.2 lbs)** (Answer: 2.5 mL)

   a. Write the two ratios in this problem as fractions.

$$\frac{1.5\text{ mg}}{1\text{ kg}} \text{ or } \frac{1\text{ kg}}{1.5\text{ mg}} \qquad \frac{20.\text{ mg}}{1\text{ mL}} \text{ or } \frac{1\text{ mL}}{20.\text{ mg}}$$

   b. Solve the problem. **72.6 lbs ⟶ kg ⟶ mg ⟶ mL**

$$\textbf{? mL} = \begin{array}{c|c|c|c} 72.6\text{ lbs} & 1\text{ kg} & 1.5\text{ mg} & 1\text{ mL} \\ \hline & 2.2\text{ lbs} & 1\text{ kg} & 20.\text{ mg} \end{array} = 2.475 \approx 2.5\text{ mL } (2\text{ s.f.})$$

6. A 5.00 mL block of aluminum contains how many aluminum atoms? The density of aluminum is 2.71 grams per mL, the molar mass of aluminum is 26.98 grams per mole, and one mole of aluminum is 6.022×10^{23} atoms. **mL ⟶ g ⟶ mole ⟶ atoms** (Answer: 3.01×10^{23} atoms)

$$\textbf{? Al atoms} = \begin{array}{c|c|c|c} 5.00\text{ mL Al} & 2.71\text{ gal} & 1\text{ mol Al} & 6.022 \times 10^{23}\text{ atoms} \\ \hline & 1\text{ mL Al} & 26.98\text{ g} & 1\text{ mol Al} \end{array} = 3.01 \times 10^{23}\text{ atom Al}$$

7. Given that there are four quarts in a gallon, 453.59 g in exactly 1 pound, and 946.3 mL in exactly 1 quart, ***calculate the mass of one gallon of mercury,*** hence verifying the statement that a gallon of mercury weighs more than 100 pounds. (The density of mercury is 13.55 g/mL.) (Answer: 113.1 lbs)

$$\textbf{? lbs} = \begin{array}{c|c|c|c|c} 1\text{ gal} & 4\text{ qt} & 946.3\text{ mL} & 13.55\text{ g} & 1\text{ lb} \\ \hline & 1\text{ gal} & 1\text{ qt} & 1\text{ mL} & 453.59\text{ g} \end{array} = 113.1\text{ lbs } (4\text{ s.f.})$$

This is an exact number. **4 s.f.** **4 s.f.** **5s.f.**

The Jabberwocky and Googs per Mulm

Use dimensional analysis to solve these nonsense conversions.

1. Nonsense words taken from the poem *Jabberwocky* (from Lewis Carroll's *Through the Looking Glass)*

   - There are 20 $\overset{\text{TT}}{\underline{\text{tumtum trees}}}$ in the $\overset{\text{TW}}{\underline{\text{tulgey wood.}}}$
   - In each $\underline{\text{tulgey wood}}$ is one $\overset{\text{FB}}{\underline{\text{frumious Bandersnatch.}}}$
   - There are 5 $\overset{\text{ST}}{\underline{\text{slithy toves}}}$ in 2 $\overset{\text{BG}}{\underline{\text{borogoves.}}}$
   - There are 2 $\overset{\text{MR}}{\underline{\text{mome raths}}}$ per Jabberwock.
   - There are 2 $\overset{\text{JB}}{\underline{\text{Jubjub birds}}}$ in 200 tumtum trees.
   - There are 200 mome raths in each borogove.
   - There are 5 Jubjub birds per slithy tove.

   If there are 5 $\underline{\text{frumious Bandersnatches,}}$ how many $\overset{\text{JW}}{\underline{\text{Jabberwocks}}}$ should there be? (Answer: 8 JW)

   HINT: *What is your "given" info (other than conversion factors)? Use this to start the set up. See "A Curious Exercise in Backchaining" for help.*

$$? \text{ JW} = \frac{\text{given 5 FB}}{} \left| \frac{1 \text{ TW}}{1 \text{ FB}} \right| \frac{20 \text{ TT}}{1 \text{ TW}} \left| \frac{2 \text{ JB}}{200 \text{ TT}} \right| \frac{1 \text{ ST}}{5 \text{ JB}} \left| \frac{2 \text{ BG}}{5 \text{ ST}} \right| \frac{200 \text{ MR}}{1 \text{ BG}} \left| \frac{1 \text{ JW}}{2 \text{ MR}} \right. = \frac{80{,}000}{10{,}000} = 8 \text{ JW}$$

2. A spaceship from another planet travels at a speed of 4.27 googs per mulm. There are 256 googs in a plotz and 12.3 plotz in a wraslm. If 3.4 tpocks equal one mulm, what is the ship's speed in wraslm per tpock? Show your work. In this case, assume numbers with decimal places are measured but whole numbers are exact (we'll pretend that "googs" are creatures!). (Answer: 4.0×10^{-4} wraslm/tpock)

$$\frac{? \text{ wraslm}}{\text{tpock}} = \frac{\text{given 4.27 googs}}{1 \text{ mulm}} \left| \frac{1 \text{ plotz}}{256 \text{ googs}} \right| \frac{1 \text{ wraslm}}{12.3 \text{ plotz}} \left| \frac{1 \text{ mulm}}{3.4 \text{ tpocks}} \right. = 3.998 \times 10^{-4} \frac{\text{wraslm}}{\text{tpocks}}$$

$$\text{or } 4.0 \times 10^{-4} \frac{\text{wraslm}}{\text{tpocks}}$$

(2s.f.)

Converting Both Numerator & Denominator

EXAMPLE

The speed of light is 2.998×10^8 m/s. How many feet per minute is this?

$$\frac{2.998 \times 10^8 \text{ m}}{\text{s}} \times \frac{10^2 \text{ cm}}{\text{m}} \times \frac{1 \text{ in}}{2.54 \text{ cm}} \times \frac{1 \text{ ft}}{12 \text{ in}} \times \frac{60 \text{ s}}{1 \text{ min}} = 5.902 \times 10^{10} \text{ ft/min}$$

1. Convert 2.7×10^{-8} J/g to cal/oz. (Answer: 1.8×10^{-7} cal/oz)

$$? \frac{\text{cal}}{\text{oz}} = \frac{2.7 \times 10^{-8} \text{ J}}{\text{g}} \left| \frac{1 \text{ cal}}{4.184 \text{ J}} \right| \frac{28.35 \text{ g}}{1 \text{ oz}} = 1.8 \times 10^{-7} \frac{\text{cal}}{\text{oz}} \text{ (2 s.f.)}$$

2. Convert 14 m/s^2 to ft/hr^2 to 2 significant figures. (Conversion: 1 hr^2 = 1.296×10^7 s^2) (Answer: 6.0×10^8 ft/hr^2)

$$? \frac{\text{ft}}{\text{hr}^2} = \frac{14 \text{ m}}{\text{s}^2} \left| \frac{10 \text{ cm}^2}{1 \text{ m}} \right| \frac{1 \text{ in}}{2.54 \text{ cm}} \left| \frac{1 \text{ ft}}{12 \text{ in}} \right| \frac{1.296 \times 10^7 \text{ s}^2}{1 \text{ hr}^2} = 595{,}275{,}590.6$$

$$= 6.0 \times 10^8 \frac{\text{ft}}{\text{hr}^2} \text{ (2 s.f.)}$$

3. Average gasoline mileage for a particular sports utility vehicle is 13 miles per gallon of gasoline. Express this miles/gal value in km/L. (Answer: 5.5 km/L)

$$? \frac{\text{km}}{\text{L}} = \frac{13 \text{ mi}}{\text{gal}} \left| \frac{1 \text{ gal}}{3.785 \text{ L}} \right| \frac{1.609 \text{ km}}{1 \text{ mi}} = \frac{5.526 \text{ km}}{\text{L}} = \frac{5.5 \text{ km}}{\text{L}} \text{ (2 s.f.)}$$

4. The kidneys of a normal adult male filter 125 milliliters of blood per minute. How many gallons of blood are filtered per day? (Answer: 47.6 gal/day)

$$? \frac{\text{gal}}{\text{day}} = \frac{125 \text{ mL}}{\text{min}} \left| \frac{1 \text{ L}}{1000 \text{ mL}} \right| \frac{1 \text{ gal}}{3.785 \text{ L}} \left| \frac{60 \text{ min}}{1 \text{ hr}} \right| \frac{24 \text{ hr}}{\text{day}} = \frac{47.6 \text{ gal}}{\text{day}} \text{ (3 s.f.)}$$

5. A patient in the hospital is given an intravenous fluid that must deliver 1.0×10^3 cc of a dextrose (sugar) solution over 8.0 hours. The intravenous fluid tubing delivers 15 drops/cc. What is the drop rate (in units of drops/min) that must be administered to the patient? (Answer: 31 drops/min)

$$? \frac{\text{drops}}{\text{min}} = \frac{15 \text{ drops}}{\text{cc}} \left| \frac{1000 \text{ cc}}{8 \text{ hrs}} \right| \frac{1 \text{ hr}}{60 \text{ min}} = \frac{31 \text{ drops}}{\text{min}}$$

Squared or Cubed Unit Conversions

SQUARED UNITS

Consider a piece of paper that is 8.5 inches by 11 inches. The area of the paper is 8.5 in × 11 in = 93 in^2. If you wanted to report this area in units of square centimeters (cm^2), you will need to convert the units.

$$8.5 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 21.59 \text{ cm} \qquad 11 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 27.94 \text{ cm}$$

$$21.59 \text{ cm} \times 27.94 \text{ cm} = 603.2246 \text{ cm}^2 = 6.0 \times 10^2 \text{ cm}^2$$

This is the same as

$$8.5 \text{ in} \times 11 \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 603.2246 \text{ cm}^2 = 6.0 \times 10^2 \text{ cm}^2$$

And, that is the same as

$$93 \text{ in}^2 \times \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^2 = 603.2246 \text{ cm}^2 = 6.0 \times 10^2 \text{ cm}^2$$

Often you may not be given the measurements for the length or width but only the total area, so converting the squared units of an area by squaring the length-based conversion factor is a skill you need to master.

When starting with a squared unit and trying to convert it to another square unit, you will need to square the conversion factor.

original squared unit × (conversion factor)^2 = another squared unit

EXAMPLE

Convert 87 ft^2 to yd^2. Given: 3 ft = 1 yd.

$$87 \text{ ft}^2 \times \left(\frac{1 \text{ yd}}{3 \text{ ft}}\right)^2 = 9.6667 \text{ yd}^2 = 9.7 \text{ yd}^2$$

CUBED UNITS

Unless a volume is a liquid-based unit like liters or gallons, volume units are based on lengths in three dimensions, such as cubic meters (m^3) or cubic yards (yd^3). The concept for performing cubed unit volume conversions is very similar to the squared unit area conversions shown above, except that instead of working in two dimensions and squaring the length-based conversion factor, volumes involve three dimensions and cubing the length-based conversion factor.

EXAMPLE

Convert 246 m^3 to mm^3.

$$\mathbf{246\ m^3 \times \left(\frac{10^3\ mm}{1\ m}\right)^3 = 246 \times (10^3)^3\ mm^3 = 246 \times 10^9\ mm^3 = 2.46 \times 10^2 \times 10^9\ mm^3}$$

$$\mathbf{= 2.46 \times 10^{11}\ mm^3}$$

Units Raised to a Power

(Units raised to a power are also called squared or cubed unit conversions.)

EXAMPLE

My house is 1044 ft^2. How many in^2 is this? (12 in = 1 ft, exactly)

Just because 12 in = 1 ft DOESN'T mean 12 in^2 = 1 ft^2!

Try these problems for more practice on units raised to a power. Show your work!

1. Convert 1018 ft^2 to in^2 (Answer: 1.466×10^5 in^2)

$$? \text{ in}^2 = \frac{1018 \text{ ft}^2}{} \left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^2 = 146{,}592 \text{ in}^2 = 146{,}600 \text{ in}^2 \text{ (4 s.f.)}$$

2. There are 3 ft in a yard. How many cubic feet are in one cubic yard?

$$? \text{ ft}^3 = \frac{1 \text{ yd}^3}{} \left(\frac{3 \text{ ft}}{1 \text{ yd}}\right)^3 = 27 \text{ ft}^2$$

3. Convert 1 m^3 to cm^3.

$$? \text{ cm}^3 = \frac{1 \text{ m}^3}{} \left(\frac{10^2 \text{ cm}}{1 \text{ m}}\right)^3 = 1 \times 10^6 \text{ cm}^3$$

4. Convert 1 m^3 to mm^3.

$$\text{mm}^3 = \frac{1 \text{ m}^3}{} \left(\frac{10^3 \text{ mm}}{1 \text{ m}}\right)^3 = 1 \times 10^9 \text{ mm}^3$$

5. Convert 1 m^3 to μm^3.

$$\frac{1 \text{ m}^3}{} \left(\frac{10^6 \ \mu\text{m}}{1 \text{ m}}\right)^3 = 1 \times 10^{18} \ \mu\text{m}^3$$

6. The density of platinum is 21.4 grams per cm^3. What is the density of platinum in units of pounds (lb) per ft^3? (Conversions: 2.205 lb = 1 kg, 1 in = 2.540 cm, 12 in = 1 ft).

(Answer: 1340 lb/ft^3)

$$? \frac{\text{lb}}{\text{ft}^3} = \frac{21.4 \text{ g}}{\text{cm}^3} \left|\frac{1 \text{ kg}}{10^3 \text{ g}}\right| \frac{2.205 \text{ lb}}{1 \text{ kg}} \left(\frac{2.54 \text{ cm}}{1 \text{ in}}\right)^3 \left(\frac{12 \text{ in}}{1 \text{ ft}}\right)^3 = \frac{1336 \text{ lb}}{\text{ft}^3} = \frac{1340 \text{ lb}}{\text{ft}^3}$$

Density

- **Density** is the **ratio** of the **mass** of the substance to the **volume** of the substance at a given temperature.
- Density has units of **g/cm^3** or **g/cc** or **g/mL** for liquids and solids, and **g/L** for gases.
- Density is an **intensive** property. It does not depend on the amount of substance.
- Density **varies** with change in **temperature.**

1. A gold-colored ring has a mass of 18.9 grams and a volume of 1.12 mL. Is the ring pure gold? (The density of gold is 19.3 g/mL.)

$$d = \frac{m}{v} = \frac{18.9\text{ g}}{1.12\text{ mL}} = \frac{16.9\text{ g}}{\text{mL}} < \text{gold } d \ \frac{19.3\text{ g}}{\text{mL}}$$

It is not a pure gold ring.

2. What volume would a 0.871 gram sample of air occupy if the density of air is 1.29 g/L?

$$d = \frac{m}{v} \text{ or } v = \frac{m}{d} \qquad v = \frac{0.871\text{ g}}{1.29\text{ g/L}} \text{ or } \frac{0.871\text{ g}}{} \left| \frac{1\text{ L}}{1.29\text{ g}} \right. = 0.675\text{ L} \quad (3\text{ s.f.})$$

3. Pumice is volcanic rock that contains many trapped air bubbles. A 225 gram sample occupied 236.6 mL.
   a. What is the density of pumice? (Answer is 0.951 g/mL.)

$$d = \frac{225\text{ g}}{236.6\text{ mL}} = 0.951\text{ g/mL} \ (3\text{ s.f.})$$

   b. Will pumice float on water? (The density of water is 1.0 g/mL.)

Yes, because density of pumice $\frac{0.951\text{ g}}{\text{mL}}$ is lower than water.

4. Which has the greater mass, 1 liter of water or 1 liter of gasoline? The density of water is 1.00 g/mL and that of gasoline is approximately 0.68 g/mL.

$$\frac{1\text{ L } H_2O}{} \left| \frac{1000\text{ mL}}{1\text{ L}} \right| \frac{1.00\text{ g}}{1\text{ mL}} = 1000\text{ g} \qquad \frac{1\text{ L gas}}{} \left| \frac{1000\text{ mL}}{1\text{ L}} \right| \frac{0.68\text{ g}}{\text{mL}} \quad \text{gas} = 680\text{ g}$$

Water has more mass as its density is higher.

5. From their density values, decide whether each of the following substances will sink or float when placed in sea water, which has a density of 1.025 g/mL.

   a. Gasoline 0.66 g/mL **Float**

   b. Mercury 13.6 g/mL **Sink**

   c. Asphalt 1.2 g/mL **Sink**

   d. Cork 0.26 g/mL **Float**

6. Mercury is a liquid metal having a density of 13.6 g/mL. What is the volume of 1.00 lb of mercury metal? (33.4 mL)

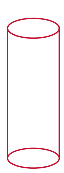
$$v = \begin{array}{c|c|c} 1.00\text{ lb} & 453.6\text{ g} & 1\text{ mL} \\ \hline & 1\text{ lb} & 13.6\text{ g} \end{array} = 33.4\text{ mL}$$

7. A sample of lead is found to have a mass of 32.6 g. A graduated cylinder contains 2.8 mL of water. After the lead sample is added to the cylinder the water level reads 5.7 mL. Calculate the density of the lead sample. (11g/mL)

v of lead = 5.7 mL – 2.8 mL = 2.9 mL

$$d = \frac{32.6\text{ g}}{2.9\text{ mL}} = \frac{11\text{ g}}{\text{mL}}$$

8. A piece of magnesium is in the shape of a cylinder with a height of 5.62 cm and a diameter of 1.34 cm. If the magnesium sample has a mass of 14.1 g, what is the density of the sample? (1.78 g/mL)



$h = 5.62\text{ cm}$

$d = 1.34\text{ cm}$

$$r = \frac{d}{2} = \frac{1.34\text{ cm}}{2}$$

$r = 0.67\text{ cm}$

$V\text{ cylinder} = A \times H = \pi r^2 h$

$$V = \frac{22}{7} \times (0.67\text{ cm})^2 \times 5.62\text{ cm}$$

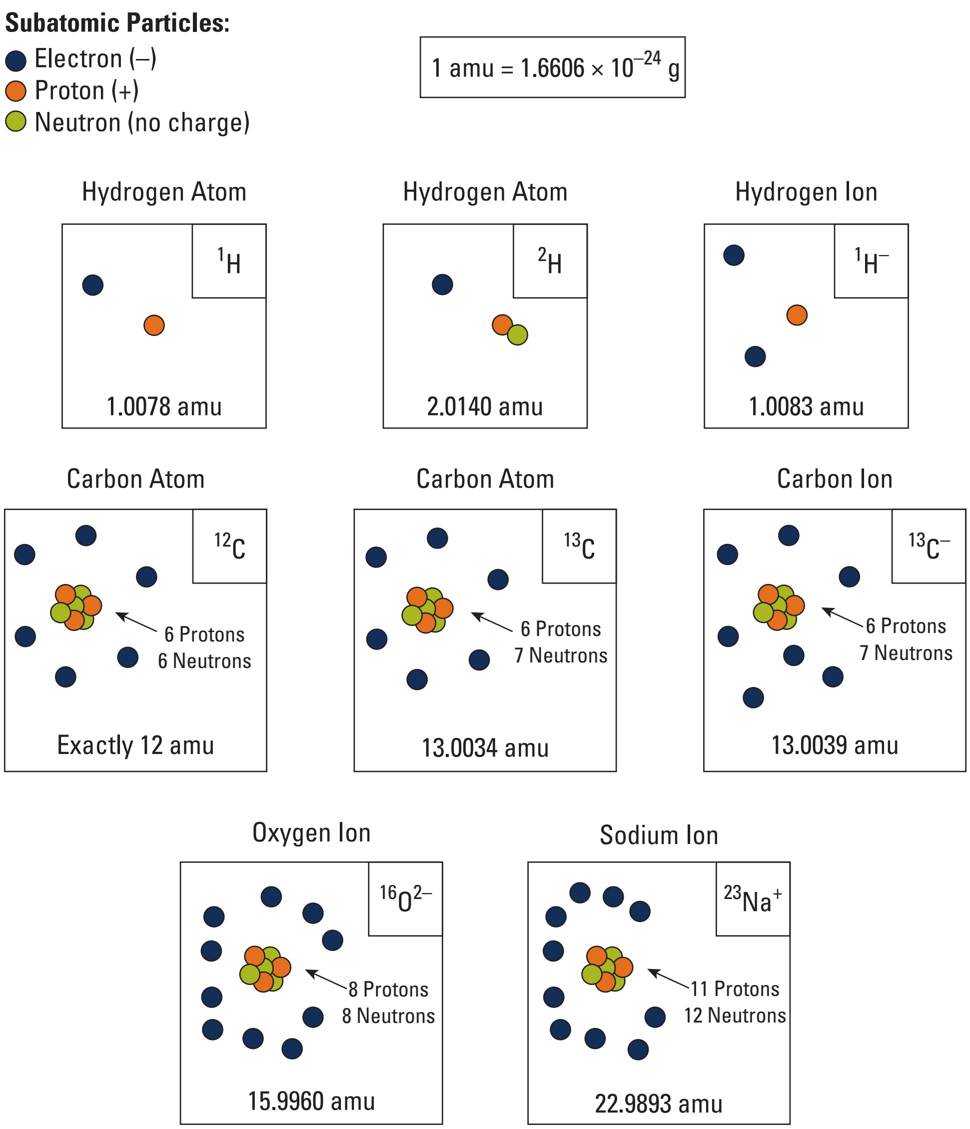
$V = 7.92\text{ cm}^3$ or mL

$$d = \frac{m}{v} = \frac{14.1\text{ g}}{7.92\text{ cm}^3} = \frac{1.78\text{ g}}{\text{cm}^3}$$

Activity: The Nuclear Atom

This activity is modified from *Chemistry: A Guided Inquiry* (3/e) by R.S. Moog and J.J. Farrell, Wiley, 2006.

Model: Schematic diagrams for various atoms and ions.



^{1}H and ^{2}H are ***isotopes*** of hydrogen.

^{12}C and ^{13}C are ***isotopes*** of carbon.

The ***nucleus*** of an atom contains the protons and neutrons.

You will need a periodic table for some of these questions.

1. How many ***protons*** are found in
   a. ^{12}C? **6** b. ^{13}C? **6** c. $^{13}C^{-}$? **6**

2. How many ***neutrons*** are found in
   a. ^{12}C? **6** b. ^{13}C? **7** c. $^{13}C^{-}$? **7**

3. How many ***electrons*** are found in
   a. ^{12}C? **6** b. ^{13}C? **6** c. $^{13}C^{-}$? **7**

4. Based on the model,
   a. What do all carbon atoms and ions have in common?

      6 protons

   b. What do all hydrogen atoms and ions have in common?

      1 proton

5. Look at the periodic table. What does the ***atomic number*** (above each atomic symbol) represent in terms of an atom's subatomic particle?

   There are two possible answers; choose the BEST one!

   Number of protons (p)
   Number of electrons (e)

   Atomic Number → 1
   H
   1.0079

6. Based on your answer to Question 6, what do all nickel (Ni) atoms have in common? Be as specific as possible. Use your periodic table. **HINT:** Nickel is element 28.
   All Ni atoms have 28 protons and 28 electrons.

7. ^{12}C and ^{13}C are isotopes of carbon. ^{1}H and ^{2}H are isotopes of hydrogen. What structural feature is different in ***isotopes*** of a particular element?
   Neutrons (n)

8. In the model, the ***mass number*** is shown as the left-hand superscript next to the atomic symbol. How can you determine the mass number from the structure of an atom (its subatomic particles)? In this case, explain how the number 2 is determined for ^{2}H isotope based on its subatomic particles. Does that work for all the other isotopes as well?
   Mass number = n + p. ^{2}H has 1 proton and 1 neutron, their sum is equal to mass number. Mass number for any isotope is equal to p + n.

   Mass Number → $^{2}\mathbf{H}$

9. Summary (fill-in-the-blanks): The subatomic particles in the nucleus of an atom are the **protons** and **neutrons**. The **atomic** (atomic or mass) number represents the number of protons in an atoms nucleus. The **mass** (atomic or mass) number includes protons and neutrons. Isotopes of an element are the same in the number of **protons** but differ in the number of **neutrons**.

Now let's look at atoms (neutral particles) versus ions (charged particles).

10. Select either the hydrogen or the carbon series and find an example of a symbol which contains a charge (in the upper right of the symbol, usually a – or a +). Compare this to an isotope of the same element which doesn't have a charge.

    a. Write them down here (include the element symbol, mass number, and charge):
    ${}^{13}_{6}C$ ${}^{13}_{6}C^{1-}$

    b. Using the drawings in Model 1, which kind of subatomic particle (proton, neutron, or electron) is responsible for making atoms into ions (charged particles, shown by – or + sign in its symbol)? **Compared to number of protons.**

    c. In terms of subatomic particles, explain how ${}^{13}C^{-}$ got a negative charge compared to ${}^{13}C$ (neutral, no charge).
    ${}^{13}_{6}C + \overset{\text{gain}}{1e^{-}} \rightarrow {}^{13}_{6}C^{1-}$ It has 6 protons, 7 electrons. Net charge –1, anion

    d. In terms of subatomic particles, explain why ${}^{23}Na^{+}$ has a positive charge.
    ${}^{23}_{11}Na \rightarrow {}^{23}_{11}Na^{+} + 1e^{-}$ lost It has 11 protons, 10 electrons. Net charge +1, ion, cation

    e. How many protons, neutrons, and electrons would be in ${}^{1}H^{+}$? (**NOTE:** ${}^{1}H^{+}$ is not in the model! See if you can predict what it would be …)
    ${}^{1}H^{+}$ will have 1 proton, 0 neutrons, and 0 electrons.

    f. In what way are ${}^{16}O^{2-}$ and ${}^{23}Na^{+}$ similar in atomic structure?
    Both have 10 electrons; same number of electrons.

11. An atomic mass unit (amu) is the unit of mass used for small particles like atoms and ions. The ***average atomic masses*** are given in the model for various atoms and ions below the element symbol (e.g., hydrogen's average atomic mass is 1.0079 amu).

    It's an average mass because it takes into account all the isotopes of hydrogen and their relative amounts in nature.

1
H
1.0079

    (average) Atomic Mass

    a. Use the ${}^{1}H$ and ${}^{2}H$ and ${}^{1}H^{-}$ (first row) in the model to determine the approximate mass of each individual subatomic particle.

       A neutron weighs about **1.0087** amu   A proton weighs about **1.0073** amu

       An electron weighs about **0.0005** amu

    b. Is most of the mass of the atom inside or outside of the nucleus? Explain.
    Inside the nucleus. Both protons and neutrons weigh about 1 amu. An electron weighs 0.0005 amu, close to zero.

12. Bonus question: There are 3 isotopes of hydrogen, ${}^{1}H$, ${}^{2}H$ (deuterium), and ${}^{3}H$ (tritium). If the average mass of all hydrogen isotopes is 1.0079, which isotope is in greatest abundance (amounts) in nature?
    ${}^{1}_{1}H$ is most abundant since average atomic mass is close to 1.

EXERCISES (START THESE WHEN YOU FINISH THE ACTIVITY)

1. Complete the following table:

Symbol	Atomic Number	Mass Number	Number of Protons	Number of Neutrons	Number of Electrons
^{40}K	**19**	**40**	**19**	**21**	**19**
$^{32}P^{3-}$	**15**	**32**	15	17	18
Zn^{2+}	**30**	**68**	**30**	38	**28**
$^{81}Br^{1-}$	35	81	**35**	**46**	36

2. Indicate whether the following statement is true or false and explain your reasoning.

An ^{18}O atom contains the same number of protons, neutrons, and electrons.

False. ^{18}O has $8p^+$, $8e^-$, and 10n. ^{16}O contains 8p, $8e^-$, and 8n.

3. **e^-** is the symbol for electrons. Electrons are gained if present on the left-side of the equation and electrons are lost if present on the right side of the equation. See if you can complete the following equations by filling in the correct ion or the correct number of electrons involved:

a. $Mg \longrightarrow$ **Mg^{2+}** + 2**e^-** **lost**
(**HINT:** Mg^{2+} or Mg^{2-}?)

b. $F + e^- \longrightarrow$ **F^{1-}**
gained

c. $Al \longrightarrow Al^{3+}$ + **$3e^-$ lost**
(**HINT:** e^- or $2e^-$ or $3e^-$?)

d. Ca^{2+} + **$2e^-$ gained** $\longrightarrow$ Ca **neutral atom**

4. How many protons, neutrons, and electrons are found in each of the following?

a. ^{24}Mg
12p, 12e, 12n

b. $^{23}Na^+$
11p, 10e, 12n

c. ^{35}Cl
17p, 17e, 18n

d. $^{35}Cl^-$
17p, 18e, 18n

e. $^{56}Fe^{3+}$
26p, 23e, 30n

f. ^{15}N
7p, 7e, 8n

g. $^{16}O^{2-}$
8p, 10e, 8n

h. $^{27}Al^{3+}$
13p, 10e, 14n

5. Using grammatically correct sentences, describe what the isotopes of an element have in common and how they are different.

Isotopes of an element have the same number of protons (same atomic number), but different number of neutrons (mass number).

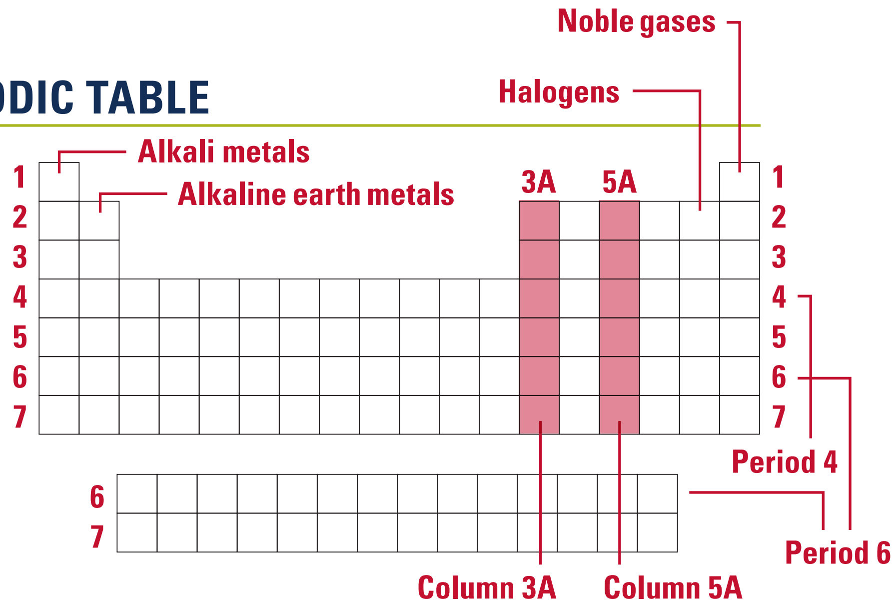
6. Define the terms: atomic number, mass number, average atomic mass, isotope, and ion. Learn them (be able to use them in a sentence).

Atomic number = number of protons
Mass number = number of protons + number of neutrons

Periodic Table Search

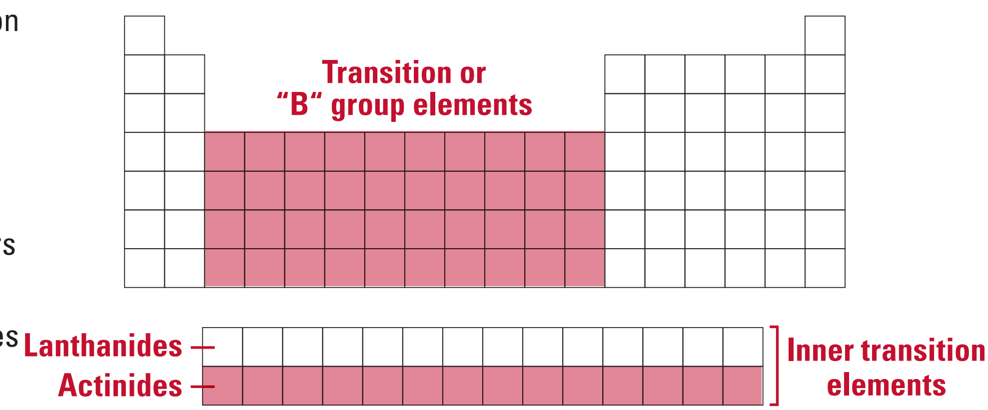
On the periodic tables provided, locate the following and label them using colored pencils, cross-hatching, or arrows.

ON THE FIRST PERIODIC TABLE



1. Identify the elements of the periodic table that are
   a. metals
   b. nonmetals
   c. metalloids (semi-metals)
2. Identify the following families (groups):
   a. alkali metals (**NOTE:** Hydrogen is NOT an alkali metal)
   b. alkaline earth metals
   c. noble gases
   d. halogens
3. Identify the location of all the members of Group 3A.
4. Identify the location of all the members of the 4th period.
5. Identify the location of all the members of Group 5A.
6. Identify the location of all the members of the 6th period.

ON THE SECOND PERIODIC TABLE



7. Identify the transition elements (B group elements).
8. Locate the inner transition elements.
9. Identify the members of the
   a. Lanthanide series
   b. Actinide series
10. Find the representative elements (main group elements, A group elements).

Atomic Theory, Periodic Table, and Ions

Some of these questions may not be covered in lecture. Use the internet to search for some of the answers.

1. a. List at least 3 observations or laws that allowed John Dalton to formulate his atomic theory.
   1. **Law of Conservation of Mass**
   2. **Law of Constant or Definite Compostion**
   3. **Law of Multiple Proportions**

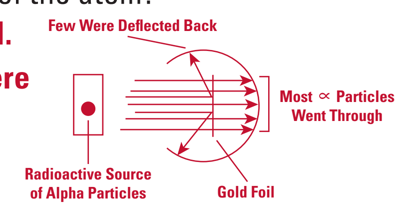
   b. What are at least 3 postulates of Dalton's atomic theory?
   1. **All matter is made up of tiny indivisible particles, called atoms.**
   2. **Atoms of one element differ from atoms of another element. All atoms of one element have the same mass.**
   3. **Atoms combine in chemical reactions by rearrangements, and can not be created or destroyed.**

2. How did Mendeleev construct the first periodic table without knowing how many protons were in each element?

   Mendeleev arranged elements in the periodic table by their relative masses and their chemical properties.

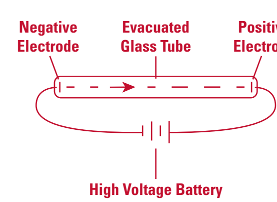
   (Later on, scientist found that masses of elements depend on the number of protons and neutrons.)

3. How did Ernest Rutherford's experiments lead to the nuclear theory of the atom?
   1. **Alpha particles were known to be dense and positively charged.**
   2. **Since $\propto$ particles were deflected and bounced back as they were repelled by dense $\oplus$ charged center called nucleus.**
   3. **Most of $\propto$ particles passed through atom, meant atom had smaller $\oplus$ center and the rest was empty space.**



4. How did J.J. Thomson's cathode ray experiments illustrate the existence of the electron?

   A beam of cathode rays (negative charged particles called electrons) travel from cathode to anode.



5. Which element in row 1A has the fewest number of protons?

   Hydrogen

6. The element with 15 protons is in what row?

   Row 3 or period 3, it is phosphorus.

7. Most element symbols are derived from their names in English. For example, Argon is Ar, Lithium is Li, Phosphorus is P and so on. There are 6 common elements with symbols derived from their Latin name. Write the symbol of each element next to the names below

   a. Iron **Fe**
   b. Silver **Ag**
   c. Tin **Sn**
   d. Copper **Cu**
   e. Mercury **Hg**
   f. Lead **Pb**

8. Where are the noble gases on the periodic table? Why are they called *noble* gases?

   Group 8A → He, Ne, Ar, Kr, Xe, Rn
   They are called noble gases because they are not chemically reactive.
   They are inert or unreactive.

9. Where are the alkali metals on the periodic table? Why are they called *alkali* metals?

   Group IA → Li, Na, K, Rb, Cs, Fr
   They react with water to make alkaline solutions (pH > 7.0).

10. Where on the periodic table are the elements that exist as gases at standard temperature and pressure?

				He
H_2	**N_2**	**O_2**	**F_2**	**Ne**
			Cl_2	**Ar**
				Kr
				Xe
				Rn

11. How are metals different from metalloids? How do we exploit the properties of metalloids for our computing needs?

Metals

1. **Are hard and shiny.**
2. **Excellent conductors of heat and electricity.**
3. **Can be hammered into thin sheets (malleable).**
4. **Can be drawn into wires (ductile).**

Metalloids

1. **Are dull and brittle.**
2. **Are semiconductors, used in electronic devices. (Silicon, Arsenic, Germanium)**

12. Determine the charge of each atom described below:

a. A gallium ion with 28 electrons and 38 neutrons
z = 31, 31 p, 28 e, charge is 3+, Ga^{3+}

b. An iron ion with 23 electrons
z = 26 p, 23 e^-, Fe^{3+}

c. An iron ion with 24 electrons and 30 neutrons
Fe^{2+}

d. A sulfur ion with 18 electrons and 16 neutrons
z = 16, 16 p, 18 e, means S^{2-}

13. An atom with 15 protons, 16 neutrons, and 15 electrons will have what charge?

No charge. The number of protons equals the number of electrons.

14. How many protons, neutrons and electrons will be in each atom below:

a. $^{138}Ba^{2+}$ **56 p, 54 e, 82 n**

b. ^{138}Ba **56 p, 56 e, 82 n**

c. $^{25}Mg^{2+}$ **12 p, 10 e, 13 n**

d. $^{37}Cl^{1-}$ **17 p, 18 e, 20 n**

15. How can we predict the charge of ions formed from the main group elements (1A-8A)?

Groups 1A, 2A, and 3A are metal, charge on cation = the group number. (E.g., 1A has 1+, 2A has 2+, and 3A has 3+ charge.)

Groups 5A, 6A, and 7A are nonmetals, charge on anions = (group number –8). (E.g., N, (5 – 8) = –3. N makes N^{3-} ion.

16. What is the common charge for ions of the group 1A elements?

1+, Li^+, Na^+, K^+, Rb^+, Cs^+

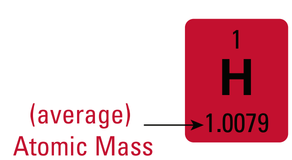
17. What is the common charge for ions of the group 6A elements?

2–, O^{2-}, S^{2-}, Se^{2-}, Te^{2-}

18. Where on the periodic table is the atomic mass listed? What does the atomic mass represent?

Average atomic mass of an element on the periodic table is listed below the symbol of the element.

It represents the weighted average of all isotopic masses and takes into account their percent of natural abundance.



Atomic Structure: Isotopes and Ions

- **An atom** is made up of protons and neutrons (both found in the nucleus) and electrons (in the surrounding electron cloud).
- **Chemical symbol:** a one or two letter abbreviation for the element.
- **The atomic number (Z)** is equal to the number of protons in the nucleus of an atom.
- **In a neutral atom,** the number of protons equals the number of electrons.
- **The mass number (A)** is equal to the number of protons plus neutrons.
- **The (average) atomic mass** of an element is the weighted average mass of an element considering its isotopes and their natural abundances.
- **The charge** on an ion indicates an imbalance between protons and electrons. Too many electrons produces a negative charge, too few, a positive charge.
- **Isotopes** of the same element have the same number of protons but different numbers of neutrons.

Isotope Symbol Notation: ${}^{A}_{Z}X$

Isotope Name: element-A

1. The atomic number of Zn is **30** and represents the number of **protons** in a Zn atom.
2. The mass number of bromine-79 is **79** and it contains **35** protons and **35** electrons.
3. It has **2** (how many?) fewer neutrons than bromine-81.
4. The average atomic mass of bromine is **79.904** amu.
5. ${}^{39}K$ is an atom, but when it loses an electron, it will have a charge of **1+**.
6. In an atom of ${}^{24}Mg$, the mass number is **24**, the atomic number is **12** and the charge is **0**.

Complete the following chart:

Element/Ion	Atomic Number	Mass Number	Charge	Protons	Electrons	Neutrons
${}^{14}C$	6	14	0	6	6	8
${}^{39}K$	19	39	0	19	19	20
${}^{26}Mg^{2+}$	12	26	2+	12	10	14
${}^{75}As^{3-}$	33	75	3–	33	36	42
${}^{109}Ag^{1+}$	47	109	1+	47	46	62
${}^{235}U$	92	235	0	92	92	143

Isotope Calculations

1. Estimate the atomic mass of copper. Copper has two isotopes, copper-63 (69.15% abundance) and copper-65 (30.85% abundance). (**HINT:** What is the approximate mass of Cu-63 in amu?)

   $(0.6915 \times 62.9296) + (0.3085 \times 64.9278)$

   $43.\overline{5}16 + 20.0\overline{3}0 = 63.5\overline{5}\text{ amu}$

2. An element exists as four different isotopes. Refer to the data table for the masses and abundances. Calculate the weighted average atomic mass of this element. (Don't forget units.)

Isotope Mass (amu)	Percent Abundance
49.946	4.35
51.941	83.79
52.941	9.50
53.929	2.37

$4.35 \div 100 = 0.0435 \times 49.946 = 2.1\overline{7}2651$

$83.79 \div 100 = 0.8379 \times 51.941 = 43.5\overline{2}\overline{1}3639$

$9.50 \div 100 = 0.0950 \times 52.941 = 5.0\overline{2}9395$

$2.37 \div 100 = 0.0237 \times 53.929 = 1.2\overline{7}8$

$52.0\overline{0}\text{ amu}$

3. Gallium has two isotopes with masses of 68.926 amu and 70.925 amu. What are the natural abundances for these two isotopes? (Answer: 60.13% and 39.87%)

   Let ^{69}Ga be = x, then ^{71}Ga will be $1 - x$. Average atomic mass can be read from PT.

   $x(68.926) + 1 - x(70.925) = 69.723$ **or** **$^{69}Ga = 0.6013 \times 100$**

   $68.926x + 70.925 - 70.925x = 69.723$ **$= 60.13\%$**

   $-1.999x = -1.202$ **$^{71}Ga = 100 - 60.13$**

   $x = 1.202/1.999 = 0.601\overline{3}$ **$= 39.87\%$**

4. Magnesium has three isotopes. One isotope has a mass of 24.986 amu and a natural abundance of 10.00%. Another isotope has a mass of 25.983 amu and a natural abundance of 11.01%. Determine the mass and natural abundance of the third isotope. (Answer: 23.98 amu, 78.99%)

   (natural abundance of the third isotope = 100% – 11.01% – 10.00% = 78.99%

   $(0.1000 \times 24.986) + (0.1101 \times 25.983) + (0.7899 \times x) = 24.3050\text{ amu}$ **(average atomic mass from periodic table)**

   $2.49\overline{8}6 + 2.86\overline{0}7 + 0.789\overline{9}x = 24.3050$

   $0.789\overline{9}x = 24.305\overline{0} - 2.49\overline{8}6 - 2.86\overline{0}7$

   $0.7899x = 18.94\overline{5}7\text{ amu}$

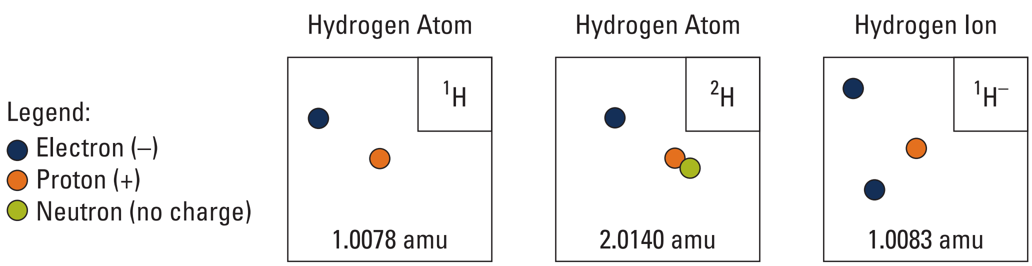
   $x = 18.94\overline{5}7/0.789\overline{9} = 23.9\overline{8}\text{ amu}$

Atomic Structure Review

1. Match the following models/discoveries/theory/law with the related person/experiment:

i. Idea of the atom	**d**	a. Gold foil experiment
ii. Plum Pudding Model	**c**	b. Antoine Lavoisier
iii. Nuclear Atom	**a**	c. Cathode Ray Tube/Discovery of the Electron
iv. Atomic Theory	**e**	d. Democritus
v. Law of Conservation of Mass	**b**	e. John Dalton

2. Each of these drawings includes subatomic particles, which represents an atom or ion. The total mass of the atom or ion is shown in units of atomic mass units in the box (amu).



   a. Which two drawings will help you estimate the mass of an electron? Calculate it and include units.*

   1.0083 amu – 1.0078 amu = 0.0005 amu

   b. Which two drawings will help you estimate the mass of a neutron? Calculate it and include units.*

   2.0140 amu – 1.0078 amu = 1.0062 amu

   c. Which drawing can you use (with the answer to part a or b) to estimate the mass of a proton? Calculate it and include units.*

   1.0078 amu – 0.0005 amu = 1.0073 amu

   d. Based on your answers, where is most of the mass of the atom, inside or outside the nucleus?

   Protons and neutrons

   e. Which of these boxes represents an ion? Explain.

   Box 3

***NOTE:** The precise masses cannot be determined this way due to the binding energy of the nuclear particles.

3. Which of the following descriptions of a subatomic particle is correct?

False **True** a. A proton has a positive charge and a mass of approximately 1 amu.

False b. An electron has a negative charge and a mass of approximately 1 amu.

c. A neutron has no charge and its mass is negligible.

d. A proton has a positive charge and a negligible mass. **False**

4. Write the chemical symbol or name associated with the element and list the numbers of protons, neutrons, and electrons in the neutral atom

Name	Symbol	# of Protons	# of Neutrons	# of Electrons
Phosphorus	$^{31}_{15}P$	**15**	16	**15**
Bromide-81 ion	$^{81}Br^{-}$	**35**	**46**	**36**

5. Describe each element as a metal, semimetal, or non metal (in the first blank). In the second blank, also classify each as one of the following: alkali metal, alkaline earth metal, halogen, noble gas, or transition metal.

   a. Ag **metal** also a **transition metal**

   b. Cl **nonmetal** also a **halogen metal**

6. Name an element in Group 6A: **oxygen**. Name an element in Period 3: **sodium**.

7. Name an element that might have the same chemical properties as calcium. **Strontium, Magnesium**

8. a. What observation in the Gold Foil experiment led to the conclusion that the atom contains a dense nucleus?

   Deflection of alpha particle by ⊕ charged dense center.

   b. What observation in the Gold Foil experiment led to the conclusion that the atom is composed of mostly empty space and the nucleus is very small?

   Most of the alpha particles passed through the atom.

9. Magnesium exists in nature as three different isotopes. Calculate the atomic mass of magnesium from the data given. You must show work for credit. Use significant figures.

Magnesium			
Atomic Mass (amu)	**Mg-24** 23.985045	**Mg-25** 24.985839	**Mg-26** 25.982595
% Abundance	78.90	10.00	11.10

$0.7890 \times 23.985045 = 18.9\bar{2}420051$
$0.1000 \times 24.985839 = 2.49\bar{8}5839$
$0.1110 \times 25.982595 = 2.88\bar{4}068045$
$24.3\bar{0}6$ **Ans: 24.31 amu**

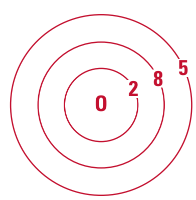
Forming Ions

Complete outer shell.

1. Noble gases are in Group **8A** and are fairly stable compared to other elements. (Why?)
2. Chemical reactivity arises from the organization of the electrons. A simplified model (developed by Niels Bohr) has electrons filling orbits around the nucleus. The number of electrons which fill each orbit is 2, 8, 8, 18 (follows the periods in the periodic table). Use the periodic table to fill in the Bohr diagrams on the next page.
3. Which element on the next page is a noble gas? What is unique about its Bohr diagram compared to the others? **Neon, 2t has completed shell 1 and shell 2. Outer shell is full.**
4. In order to gain stability, atoms gain or lose electrons to fill their outermost electron shells. This shell is called the "valence shell."
5. For each of the elements on the next page, determine how many electrons would be gained or lost to have a filled valence shell. Which is more likely, a loss or gain of electrons?
6. If atoms lose electrons to fill their valence shell, will the resulting ion be positive or negative?
7. If atoms gain electrons, will the resulting ion be a cation or an anion?
8. For each of the elements on the next page, determine the symbol for the most likely ion to result (in order to have a filled valence shell). For example, sodium will likely become a Na^+ cation.

PRACTICE

1. Draw the Bohr diagram for phosphorus.



2. What ion does phosphorus become?

   P atom gains $3e^-$ and becomes P^{3-} ion.

3. Compare phosphorus with nitrogen (on the next page). In what ways are they similar?

   They gain $3e^-$ in outermost shell to make 3 charged ions.

4. What type of ions (on the next page) do you think will bond with phosphorus? (**HINT:** Opposite charges attract!)

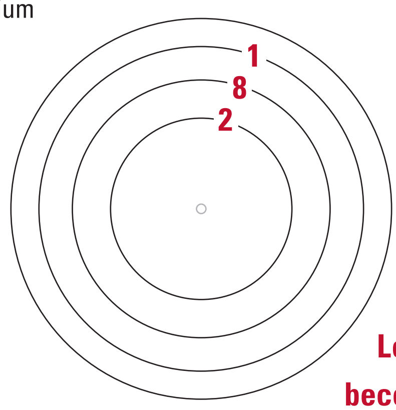
   Cations Na^+, Al^{3+}

5. Is there a way to predict for any element if it becomes a positively charge or a negatively charged ion? (**HINT:** Look at its location on the periodic table)

   Yes—metals make cations because metals lose electrons.
   Nonmetals can gain electrons, they make anions.

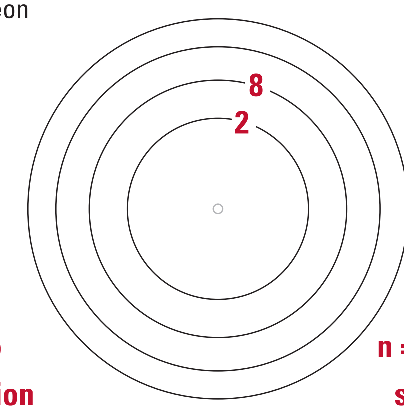
Bohr Model Diagrams

Sodium



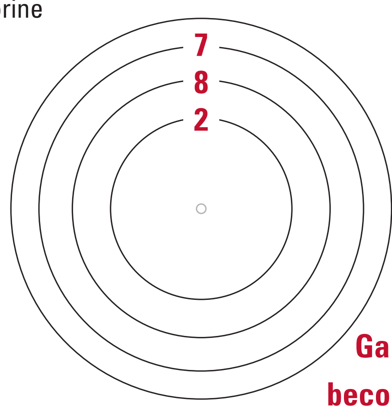
Lose $1e^-$ to become Na^+ ion

Neon



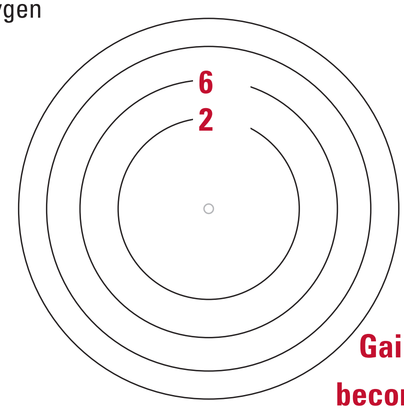
n = 2, valence shell is full

Chlorine



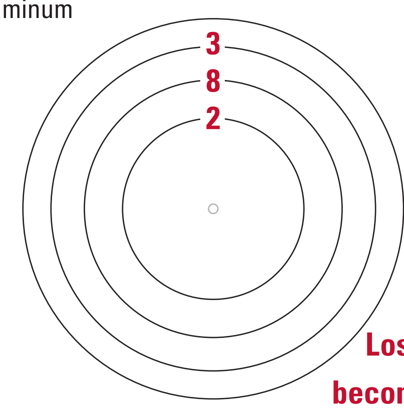
Gain $1e^-$ to become Cl^- ion

Oxygen



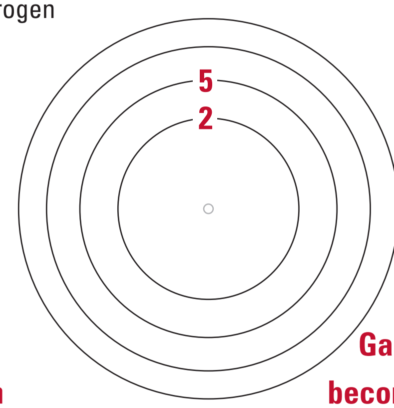
Gain $2e^-$ to become O^{2} ion

Aluminum



Lose $3e^-$ to become Al^{3+} ion

Nitrogen



Gain $3e^-$ to become N^{3-} ion

Charges of Ions

Get familiar with the ionic charge of various elements when they become ions.

1. Draw demarcation line for metals vs. non-metals (exception: hydrogen is a non-metal).
2. Number columns for groups IA–VIIIA (1A–8A).
3. Show fixed ionic charges (for metals and non-metals):

1A, 1+	2A, 2+											3+ 3A	4A	3– 5A	2– 6A	1– 7A	8A
H^+																	He
Li^+														N^{3-}	O^{2-}	F^-	Ne
Na^+	Mg^{2+}										2+	Al^{3+}		P^{3-}	S^{2-}	Cl^-	Ar
K^+	Ca^{2+}										Zn^{2+}	Ga^{3+}			Se^{2-}	Br^-	Kr
Rb^+	Sr^{2+}									Ag^+	Cd^{2+}				Te^{2-}	I^-	Xe
Cs^+	Ba^{2+}																Rn

NAMES OF NON-METAL ANIONS (NEGATIVELY-CHARGED IONS)

For non-metal ions (charged atoms) the –ide ending is used after the first syllable of the element.

H^- hydride	N^{3-} nitride	O^{2-} oxide	F^- fluoride	
	P^{3-} phosphide	S^{2-} sulfide	Cl^- chloride	
	As^{3-} arsenide	Se^{2-} selenide	Br^- bromide	
		Te^{2-} telluride	I^- iodide	

NOTE: H can have a negative charge (in which case it is called "hydride") OR a positive charge (in which case it is called "hydrogen ion"). The elements in Group IVA (4A) can also make anions; C^{4-} is carbide and Si^{4-} is silicide.

NAMES OF METAL CATIONS (POSITIVELY-CHARGED IONS)

For metal ions (charged atoms) the name of the element is followed by "ion."

EXAMPLE

Atom/Element	Ion	Name of Ion
Mg	Mg^{2+}	magnesium ion
Ag	Ag^{+}	silver ion

Some elements can exist in several forms. For example, chlorine can exist as neutral atoms, diatomic molecules, or ions.

chlorine atom Cl chlorine Cl_2 chloride ion Cl^{-}

Ionic charge is always shown as a SUPERSCRIPT for ions, otherwise the atom is assumed to be neutral. When the charge is neutralized, the superscripts are not written (e.g., $Na^{+}Cl^{-}$ becomes NaCl).

PRACTICE

Write the formula and the name of the ion that is formed for each element.

Atom/ Element	Ion Formula (show charge)	Cation?	Anion?	Name of Ion
Na	Na^{+}	X		Sodium ion
O	O^{2-}		X	Oxide ion
Zn	Zn^{2+}	X		Zinc ion
F	F^{-}		X	Fluoride ion
K	K^{+}	X		Potassium ion
N	N^{3-}		X	Nitride ion
Al	Al^{3+}	X		Aluminum ion
P	P^{3-}		X	Phosphide ion
Sr	Sr^{2+}	X		Strontium ion
S	S^{2-}		X	Sulfide ion

Tutorial: Ionic Nomenclature

PART I: FORMULAS AND NOMENCLATURE OF IONIC COMPOUND COMPOSED OF CATIONS AND ANIONS

TYPES OF CATIONS (POSITIVE IONS)

1. Metals lose electrons to form positive ions. These ions are called **monoatomic** because they are made up of only **ONE** ion. They can be of two types: ***Constant charge*** or ***Variable charge.***

   a. ***Constant charge***

   Group IA, IIA, IIIA

   A few transition metals Ag^{+}, Zn^{2+} and Cd^{2+}

   The names of these ions are the same as the name of the atom.

   Examples: Na^{+} = sodium ion, Zn^{2+} = zinc ion

   b. ***Variable charge***

   Most transition metals (except for silver, zinc and cadmium)

   A few representative metals: Sn, Bi, and Pb

Fe^{2+} Iron (II)	Cu^{+} Copper (I)	Sn^{4+} Tin (IV)
Fe^{3+} Iron (III)	Cu^{2+} Copper (II)	Sn^{2+} Tin (II)

2. **A polyatomic ion:** Consists of more than one atom. The most common positive one is ammonium, NH_4^{+}

TYPES OF ANIONS (NEGATIVE IONS)

1. Nonmetals gain electrons to form negative ions. These ions are called **monoatomic** because they are made up of only **ONE** ion. The names of these ions end in –ide.

   Examples: S^{2-} = sulfide, Cl^{1-} = chloride, N^{3-} = nitride, O^{2-} = oxide

2. **A polyatomic ion:** Consists of more than one atom. The names of polyatomic anions often end in –ate or –ite.

   Examples: NO_3^{1-} = nitrate, NO_2^{1-} = nitrite, SO_4^{2-} = sulfate, SO_3^{2-} = sulfite

   Two important polyatomic ions end in –ide: OH^{1-} = hydroxide, CN^{1-} = cyanide

Determine if the compound is ***ionic:*** Contains a cation and an anion—often a metal and a nonmetal.

HINT: Formula begins with a **metal** or **NH_4^+**

IONIC: TO WRITE THE FORMULA

1. Write the symbol of the cation with a SUPERSCRIPT charge.
2. Write the symbol of the anion with a SUPERSCRIPT charge.
3. If the charges are NOT balanced, CRISS-CROSS to find the number of each atom necessary to balance the charges.
4. Use a parenthesis if more than ONE polyatomic ion is necessary.

EXAMPLES

- Aluminum oxide Al^{3+} O^{2-} Al_2O_3
- Iron (II) phosphide Fe^{2+} P^{3-} Fe_3P_2
- Calcium nitrate Ca^{2+} NO_3^{1-} $Ca(NO_3)_2$
- Copper (I) carbonate Cu^{1+} CO_3^{2-} Cu_2CO_3
- Barium oxide Ba^{2+} O^{2-} BaO

  NOTE: The charges are already balanced so the ratio is 1:1.
- Calcium hydroxide Ca^{2+} OH^{1-} $Ca(OH)_2$

IONIC: TO WRITE THE NAME

1. Write the name of the cation.

   If the cation has a variable charge, determine the Roman numeral.
2. Follow with the name of the anion.

EXAMPLES

- **Na_2O** **Sodium oxide** (Group IA—no Roman numeral needed)
- **Cr_2S_3** **Chromium (III) sulfide** (Most transition metals need Roman numerals.)
- **$BaSO_4$** **Barium sulfate** (Group IIA—no Roman numeral needed)
- **$Pb(OH)_2$** **Lead (II) hydroxide** (Pb is a main group metal with variable charge.)
- **Cu_2CO_3** **Copper (I) carbonate** (Most transition metals need Roman numerals.)
- **$FeSO_4$** **Iron (II) sulfate** (Most transition metals need Roman numerals.)

Polyatomic Ions to *Memorize*

Polyatomic Ions to Memorize			
Formula	**Name**	**Formula**	**Name**
NO_2^-	nitrite	HCO_3^-	hydrogen carbonate (or bicarbonate)
NO_3^-	nitrate	ClO^-	hypochlorite
SO_3^{2-}	sulfite	ClO_2^-	chlorite
HSO_3^-	hydrogen sulfite (or bisulfite)	ClO_3^-	chlorate
HSO_4^-	hydrogen sulfate (or bisulfate)	ClO_4^-	perchlorate
SO_4^{2-}	sulfate	BrO^-	hypobromite
PO_3^{3-}	phosphite	BrO_2^-	bromite
PO_4^{3-}	phosphate	BrO_3^-	bromate
HPO_4^{2-}	hydrogen phosphate	BrO_4^-	perbromate
$H_2PO_4^-$	dihydrogen phosphate	IO^-	hypoiodite
OH^-	hydroxide	IO_2^-	iodite
H_3O^+	hydronium	IO_3^-	iodate
NH_4^+	ammonium	IO_4^-	periodate
CO_3^{2-}	carbonate		

Other Polyatomic Ions You May Encounter (not required to memorize)			
Formula	**Name**	**Formula**	**Name**
CrO_4^{2-}	chromate	$C_6H_5O_7^{3-}$	citrate
$Cr_2O_7^{2-}$	dichromate	BO_3^{3-}	borate
CN^-	cyanide	MnO_4^-	permanganate
$C_2H_3O_2^-$ (or CH_3COO^-)	acetate	O_2^{2-}	peroxide
$C_2O_4^{2-}$	oxalate	Hg_2^{2+}	mercury (I)

Memorizing Polyatomic Ions: The T43 Method

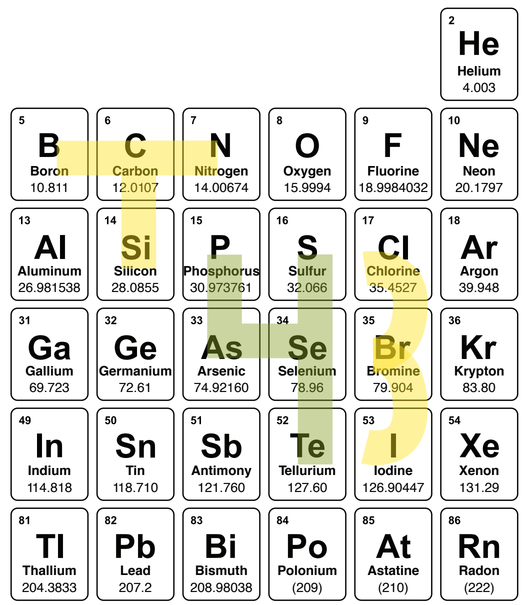
Many of the polyatomic ion charges and number of oxygens can be memorized using the "T43 method." If your instructor does not cover this, try a search on "T43 method" in YouTube to see how it's used.

USING A PERIODIC TABLE

1. Draw a "T" over the B, C, N and Si.

   The "T" elements indicate 3 oxygens and a charge of –3, –2, –1 respectively (left to right).

   For example, the polyatomic ions you get are BO_3^{3-}, CO_3^{2-}, and NO_3^{-}. These end in "ate": borate, carbonate, and nitrate. Ignore silicon—you will not see it very often in this course.



2. Draw a "4" over the P, As/S, Se/Te.

   The "4" elements contain 4 oxygens and a charge of –3 and –2, respectively (left to right). These end in "ate": phosphate and sulfate.

3. Draw a "3" connecting the Cl, Br, and I.

   These "3" elements contain 3 oxygens and a –1 charge: ClO_3^{-}, BrO_3^{-}, and IO_3^{-}. These end in "ate": chlorate, bromate, iodate.

PRACTICE

Provide the formulas for these polyatomic ions using the T43 method (don't forget the charge!)

1. Bromate **BrO_3^{-}**
2. Carbonate **CO_3^{2-}**
3. Nitrate **NO_3^{1-}**
4. Phosphate **PO_4^{3-}**
5. Sulfate **SO_4^{2-}**
6. Iodate **IO_3^{1-}**

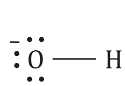
Deriving Polyatomic Ions

You need to know the names and formulas (including charges) for all 27 of these polyatomic ions.

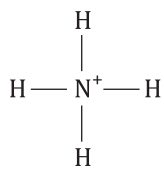
You can do this by memorizing only 7 of them and then deriving the rest! Here's how:

Memorize these two:

Hydroxide OH^-



Ammonium NH_4^+

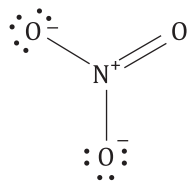
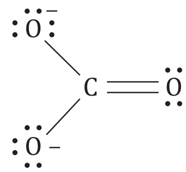
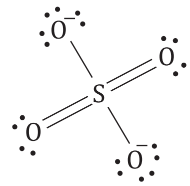
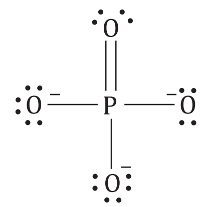


Halogen oxyanions: $XO_\#^-$

(X can be Cl, Br, or I; names will use "chlor" or "brom" or "iod")

→ per …ate	**perchlorate**
increase –ate formula by one more O atom	ClO_4^-
Memorize this one: **chlorate**	$\mathbf{ClO_3^-}$
→ –ite	**chlorite**
reduce –ate formula by one O atom	ClO_2^-
→ hypo …ite	**hypochlorite**
reduce –ite formula by one more O atom	ClO^-

Formula	Name
NO_2^-	nitrite
NO_3^-	nitrate
SO_3^{2-}	sulfite
HSO_3^-	hydrogen sulfite (bisulfite)
HSO_4^-	hydrogen sulfate (bisulfate)
SO_4^{2-}	sulfate
PO_3^{3-}	phosphite
PO_4^{3-}	phosphate
HPO_4^{2-}	hydrogen phosphate
$H_2PO_4^-$	dihydrogen phosphate
OH^-	hydroxide
H_3O^+	hydronium
NH_4^+	ammonium
CO_3^{2-}	carbonate
HCO_3^-	hydrogen carbonate (bicarbonate)
ClO^-	hypochlorite
ClO_2^-	chlorite
ClO_3^-	chlorate
ClO_4^-	perchlorate
BrO^-	hypobromite
BrO_2^-	bromite
BrO_3^-	bromate
BrO_4^-	perbromate
IO^-	hypoiodite
IO_2^-	iodite
IO_3^-	iodate
IO_4^-	periodate

	Memorize These; Use to Derive Related Ions			
	Nitrate	**Carbonate**	**Sulfate**	**Phosphate**
	NO_3^-	CO_3^{2-}	SO_4^{2-}	PO_4^{3-}
				
Related Ions				
→ –ite	nitrite	—	sulfite	phosphite
reduce –ate formula by one O atom	NO_2^-	—	SO_3^{2-}	PO_3^{3-}
→ hydrogen ...ate	—	hydrogen carbonate (or bicarbonate)	hydrogen sulfate (or bisulfate)	hydrogen phosphate
add H^+ to –ate formula (add H and increase charge by 1)	—	HCO_3^-	HSO_4^-	HPO_4^{2-}
→ dihydrogen ...ate	—	—	—	dihydrogen phosphate
add two H^+ to –ate formula	—	—	—	$H_2PO_4^-$
→ hydrogen ...ite	—	—	hydrogen sulfite	—
reduce –ate formula by one O atom, add one H^+	—	—	HSO_3^-	—
Related Acids				
→ acid form of ...ate –ate ending changes to –ic acid	nitric acid	carbonic acid	sulfuric acid	phosphoric acid
add enough H^+ to –ate formula until neutral charge	HNO_3	H_2CO_3	H_2SO_4	H_3PO_4
→ acid form of ...ite –ite ending changes to –ous acid	nitrous acid	—	sulfurous acid	phosphorous acid
add enough H^+ to –ate formula until neutral charge	HNO_2	—	H_2SO_3	H_3PO_3

Combining Ions

1. Give the charges on the ions of the following elements:

   a. Ca

   b. Rb

   c. Al

   d. N

   e. Se

   f. Br

2. What will the charge be on elements of the following groups?

   a. Group 1

   b. Group 13

   c. Group 15

   d. Group 16

3. Write the names of the following metal cations: (**NOTE:** Use Roman numerals in parenthesis to indicate charge on metals with a variable charge. Roman numerals are not used for metals with a constant charge.)

   a. Li^{+} ________________

   b. Ca^{2+} ________________

   c. Fe^{2+} ________________

   d. Cu^{2+} ________________

   e. Cu^{+} ________________

   f. Zn^{2+} ________________

4. Write the names of the following nonmetal cations: (**NOTE:** Use the root name of the nonmetal and the ending *-ide*.)

   a. O^{2-} ________________

   b. Cl^{-} ________________

   c. N^{3-} ________________

   d. S^{2-} ________________

   e. F^{-} ________________

   f. P^{3-} ________________

5. Give the formula of the ionic compounds formed between the following ions and name them.)

   a. Na^{+} and Br^{-} ______________________

   b. Sr^{2+} and I^{-} ______________________

   c. Li^{+} and O^{2-} ______________________

   d. Cs^{+} and N^{3-} ______________________

   e. Aluminum and sulfur ______________

   f. Calcium and oxygen ______________

   g. Barium and chlorine ______________

   h. Potassium and phosphorus ________

Ionic Compound Matrix

Fill in the matrix with the chemical formula derived from the pairing of each cation and anion. Write its name underneath the chemical formula.

anion → cation ↓	chloride Cl^-	oxide S^{2-}	nitrate SO_4^{2-}	sulfite SO_3^{2-}	phosphate CO_3^{2-}
sodium Na^+	$NaCl$	Na_2S	Na_2SO_4	Na_2SO_3	Na_2CO_3
calcium Ca^{2+}					
aluminum Al^{3+}					
iron (II) Fe^{2+}					
iron (III) Fe^{3+}					
copper (I) Cu^+	$CuCl$	Cu_2S	Cu_2SO_4	Cu_2SO_3	Cu_2CO_3
copper (II) Cu^{2+}	$CuCl_2$	CuS	$CuSO_4$	$CuSO_3$	$CuCO_3$
chromium (II) Cr^{2+}					
silver Ag^+					
zinc Zn^{2+}					
ammonium NH_4^+	NH_4Cl	$(NH_4)_2S$	$(NH_4)_2SO_4$	$(NH_4)_2SO_3$	$(NH_4)_2CO_3$

Ionic Nomenclature

1. Binary Ionic Compounds with Constant Charge Metal Cations
   a. NaBr **Sodium bromide**
   b. Zn_3P_2 **Zinc phosphide**
   c. CaS **Calcium sulfide**
   d. AlP **Aluminum phosphide**
   e. Na_2O **Sodium oxide**
   f. Rb_3N **Rubidium nitride**
   g. AlF_3 **Aluminum fluoride**
   h. cesium iodide **CsI**
   i. zinc nitride $Zn^{2+}N^{3-} \longrightarrow Zn_3N_2$
   j. potassium sulfide K_2S
   k. silver oxide Ag_2O
   l. calcium bromide $CaBr_2$
   m. strontium chloride $SrCl_2$
   n. sodium hydride **NaH**

2. Binary Ionic Compounds with Variable Charge Metal Cations
   a. Cu_2O **Copper (I) oxide**
   b. SnS_2 **Tin (IV) sulfide**
   c. Pb_3N_2 **Lead (II) nitride**
   d. $NiCl_2$ **Nickle (II) chloride**
   e. $FeCl_3$ **Iron (III) chloride**
   f. $SnCl_4$ **Tin (IV) chloride**
   g. CoS **Cobalt (II) sulfide**
   h. Cr_2S_3 **Chromium (III) sulfide**
   i. manganese (VII) arsenide $Mn^{7+}As^{3-} \longrightarrow Mn_3As_7$
   j. platinum (II) sulfide **PtS**
   k. vanadium (V) nitride $V^{5+}N^{3-} \longrightarrow V_3N_5$
   l. gallium (III) oxide $Ga^{3+}O^{2-} \longrightarrow Ga_2O_3$
   m. tin (II) fluoride $Sn^{2+}F^{-} \longrightarrow SnF_2$
   n. copper (II) chloride $Cu^{2+}Cl^{1-} \longrightarrow CuCl_2$
   o. chromium (III) oxide $Cr^{3+}O^{2-} \longrightarrow Cr_2O_3$
   p. titanium (II) selenide $Ti^{2+}Se^{2-} \longrightarrow TiSe$

3. 7 Polyatomic Ions
   a. NO_3^- **Nitrate**
   b. CO_3^{2-} **Carbonate**
   c. SO_4^{2-} **Sulfate**
   d. PO_4^{3-} **Phosphate**
   e. ClO_4^- **Perchlorate**
   f. OH^- **Hydroxide**
   g. NH_4^+ **Ammonium**

4. Mixed Polyatomic Ions
   a. NO_2^- **Nitrite**
   b. PO_4^{3-} **Phosphate**
   c. SO_3^{2-} **Sulfite**
   d. ClO_3^- **Chlorate**
   e. NH_4^+ **Ammonium**
   f. HSO_4 **Hydrogen sulfate (bisulfate)**
   g. BrO_4^- **Perbromate**
   h. IO_2^- **Iodite**
   i. SO_4^{2-} **Sulfate**
   j. BrO^- **Hypobromite**
   k. PO_3^{3-} **Phosphite**
   l. ClO^- **Hypochlorite**
   m. BrO_2^- **Bromite**
   n. NO_3^- **Nitrate**
   o. HPO_4^{2-} **Hydrogen phosphate**
   p. IO_3^- **Iodate**
   q. $H_2PO_4^-$ **Dihydrogen phosphate**
   r. BrO_3^- **Bromate**

s. OH^- **Hydroxide**
t. IO^- **Hypoiodite**
u. HCO_3^- **Hydrogen carbonate (bicarbonate)**
v. IO_4^- **Periodate**
w. ClO_2^- **Chlorite**
x. CO_3^{2-} **Carbonate**
y. HSO_3^- **Hydrogen sulfite**
z. ClO_4^- **Perchlorate**

5. Ionic Compounds with Polyatomic Ions & Constant Charge Ions
   a. $CaCO_3$ **Calcium carbonate**
   b. $KClO_4$ **Potassium perchlorate**
   c. $CdSO_3$ **Cadmium sulfite**
   d. Li_2SO_3 **Lithium sulfite**
   e. Ag_3PO_4 **Silver phosphate**
   f. NH_4F **Ammonium fluoride**
   g. $Al(NO_2)_3$ **Aluminum nitrite**
   h. $Ba(HCO_3)_2$ **Barium hydrogencarbonate**
   i. potassium hydroxide **KOH**
   j. beryllium hydroxide **$Be(OH)_2$**
   k. strontium acetate **$Sr(CH_3COO)_2$**
   l. ammonium oxide **$(NH_4)_2O$**
   m. silver cyanide **$AgCN$**
   n. zinc carbonate **$ZnCO_3$**
   o. aluminum bicarbonate **$Al(HCO_3)_3$**
   p. magnesium phosphate **$Mg_3(PO_4)_2$**

6. Ionic Compounds with Variable Charge Metal Cations & Polyatomic Ions
   a. $Fe(HCO_3)_2$ **Iron (II) hydrogen carbonate**
   b. $NiPO_4$ **Nickel (III) phosphate**
   c. $V_2(SO_4)_3$ **Vanadium (III) sulfate**
   d. $Ti(CN)_4$ **Titanium (IV) cyanide**
   e. $Sc(OH)_3$ **Scandium hydroxide**
   f. $Cu(NO_2)_2$ **Copper (II) nitrite**
   g. $CoCO_3$ **Cobalt (II) carbonate**
   h. $Sn(HPO_4)_2$ **Tin (II) hydrogen phophate**
   i. copper (II) chlorate **$Cu(ClO_3)_2$**
   j. cobalt (III) chromate **$Co_2(CrO_4)_3$**
   k. lead (IV) sulfate **$Pb(SO_4)_2$**
   l. iron (II) phosphate **$Fe_3(PO_4)_2$**
   m. chromium (II) carbonate **$CrCO_3$**
   n. iron (II) cyanide **$Fe(CN)_2$**
   o. nickel (II) sulfite **$NiSO_3$**
   p. tin (IV) chlorate **$Sn(ClO_3)_4$**

7. Mixed Ionic Compounds
   a. tin (II) phosphate **$Sn^{2+}PO_4^{3-} \longrightarrow Sn_3(PO_4)_2$**
   b. zinc oxide **$Zn^{2+}O^{2-} \longrightarrow ZnO$**
   c. barium sulfate **$Ba^{2+}SO_4^{2-} \longrightarrow BaSO_4$**
   d. lithium nitride **$Li^{2+}N^{3} \longrightarrow Li_3N$**
   e. silver fluoride **$Ag^{+}F^{-} \longrightarrow AgF$**
   f. barium hydroxide **$Ba^{2+}OH^{-} \longrightarrow Ba(OH)_2$**
   g. lead (II) iodide **$Pb^{2+}I^{1-} \longrightarrow PbI_2$**
   h. mercury (II) chloride **$Hg^{2+}Cl^{1-} \longrightarrow HgCl_2$**
   i. cobalt (II) nitrate **$Co^{2+}NO_3^{1-} \longrightarrow Co(NO_3)_2$**
   j. lithium bromide **$Li^{+}Br^{-} \longrightarrow LiBr$**
   k. $AlBr_3$ **Aluminum bromide**
   l. FeS **$Fe^{2+}S^{2-}$, iron (II) sulfide**
   m. $(NH_4)_3PO_4$ **Ammonium phosphate**
   n. $Hg(NO_3)_2$ **$Hg^{2+}NO_3^{1-}$, mercury (II) nitrate**
   o. Ag_2S **$Ag^{+}S^{2-}$ silver sulfide**
   p. $KMnO_4$ **Potassium permanganate**
   q. $MgCl_2$ **Magnesium chloride**
   r. Cr_2O_3 **$Cr^{3+}O^{2-}$, chromium (III) oxide**
   s. K_3P **Potassium phosphide**
   t. $TiCl_2$ **$Ti^{2+}Cl^{1-}$, titanium (II) chloride**

Nomenclature Practice Test 1

1. Write the formulas of the following ionic compounds:
   - a. tin (II) phosphate $Sn^{2+}PO_4^{3-}$ **$Sn_3(PO_4)_2$**
   - b. zinc oxide $Zn^{2+}O^{2-}$ **ZnO**
   - c. barium sulfate $Ba^{2+}SO_4^{2-}$ **$BaSO_4$**
   - d. lithium nitride $Li^{+}N^{3-}$ **Li_3N**
   - e. silver fluoride $Ag^{+}F^{-}$ **AgF**
   - f. barium hydroxide $Ba^{2+}OH^{-}$ **$Ba(OH)_2$**
   - g. lead (II) iodide $Pb^{2+}I^{-}$ **PbI_2**
   - h. mercury (II) chloride $Hg^{2+}Cl^{1-}$ **$HgCl_2$**
   - i. cobalt (II) nitrate $Co^{2+}NO_3^{-}$ **$Co(NO)_3)_2$**
   - j. lithium bromide $Li^{+}Br^{-}$ **LiBr**

2. Write the names of the following ionic compounds:
   - a. $AlBr_3$ **Aluminum bromide**
   - b. FeS $Fe^{2+}S^{2-}$ **Iron (II) sulfide**
   - c. $(NH_4)_3PO_4$ **Ammonium phosphate**
   - d. $Hg(NO_3)_2$ **Mercury (II) nitrate**
   - e. Ag_2S **Silver sulfide**
   - f. $KMnO_4$ **Potassium permanganate**
   - g. $MgCl_2$ **Magnesium chloride**
   - h. Cr_2O_3 **Chromium (III) oxide**
   - i. K_3P **Potassium phosphide**
   - j. $TiCl_2$ **Titanium (II) chloride**

Nomenclature Practice Test 2

1. Write the formulas of the following ionic compounds:
   - a. iron (II) arsenide $Fe^{2+}As^{3-}$ **Fe_3As_2**
   - b. lead (II) sulfate $Pb^{2+}SO_4^{2-}$ **$PbSO_4$**
   - c. lead (IV) hydroxide $Pb^{4+}OH^{1-}$ **$Pb(OH)_4$**
   - d. copper(II)acetate $Cu^{2+}CH_3COO^{-}$ **$Cu(CH_3COO)_2$**
   - e. beryllium chloride $Be^{2+}Cl^{1-}$ **$BeCl_2$**
   - f. ammonium chromate $NH_4^{+}CrO_4^{2-}$ **$(NH_4)_2CrO_4$**
   - g. silver oxide $Ag^{+}O^{2-}$ **Ag_2O**
   - h. potassium sulfide $K^{+}S^{2-}$ **K_2S**

2. Write the names of the following ionic compounds:
   - a. KI **Potassium iodide**
   - b. $Mn_2(SO_3)_7$ $Mn^{7+}SO_3^{2-}$ **Manganese (VII) sulfite**
   - c. $SnBr_4$ $Sn^{4+}Br^{-}$ **Tin (IV) bromide**
   - d. Mg_3P_2 **Manganese phosphide**
   - e. NaF **Sodium fluoride**
   - f. $Sr(MnO_4)_2$ **Strontium permanganate**
   - g. $Cr_3(PO_4)_2$ $Cr^{2+}PO_4^{3-}$ **Chromium (II) phosphate**
   - h. Al_2Se_3 **Aluminum selenide**

Covalent Compounds

Covalent compounds (also called "molecular" compounds)

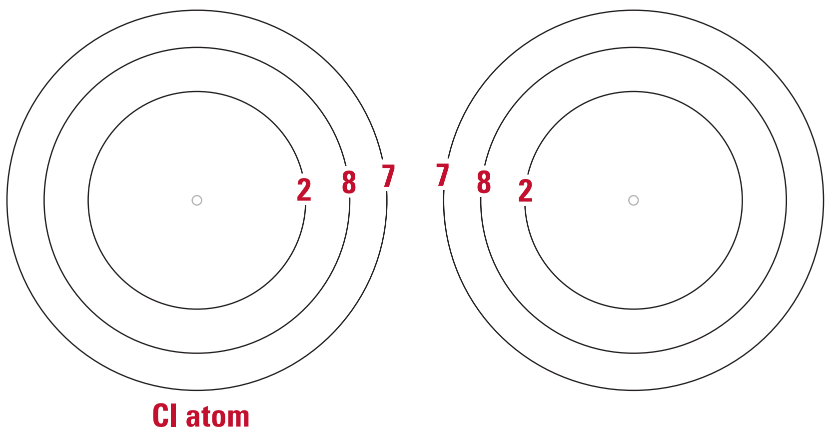
Now we will learn how to name COVALENT compounds of this form:

Non-Metal + Non-Metal

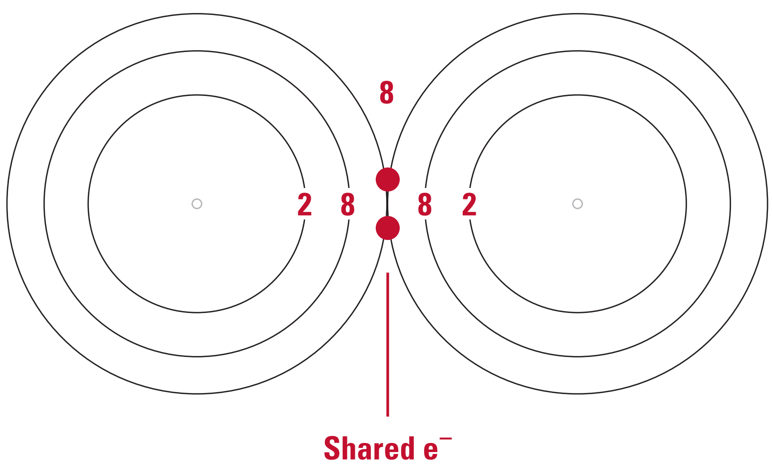
Covalent Compounds do not form from ions. They are from atoms which share electrons in order to mutually fill their valence shells. You will learn more about this toward the end of the quarter.

For example, covalent bonding explains why some elements are diatomic (BrINClHOF).

Use these templates to complete a Bohr model diagram for Cl_2. Treat each of these two templates as a separate chlorine atom.



Notice that both chlorine atoms do not have a filled shell. They can obtain a filled shell by sharing electrons with each other. There are two electrons shared between them.



Covalent Nomenclature

Covalent compounds do not form from ions. Therefore, you simply have to state the chemical formula using words and prefixes, and continue using binary endings.

EXAMPLE

$SiCl_4$ is silicon tetrachloride (**NOTE:** Mono is usually omitted from the first element.)

Prefixes for Covalent Compounds (memorize)			
Number	**Prefix**	**Number**	**Prefix**
1	mono	6	hexa
2	di	7	hepta
3	tri	8	octa
4	tetra	9	nona
5	penta	10	deca

You should also know the common names for three compounds:

- H_2O (water)
- CH_4 (methane)
- NH_3 (ammonia)

1. XeO_3 **Xenon trioxide**
2. OF_2 **Oxygen difluoride**
3. SO_3 **Sulfur trioxide**
4. P_2O_5 **Diphosphorus pentoxide**
5. Cl_2O **Dichlorine monoxide**
6. NI_3 **Nitrogen triiodide**
7. NH_3 **Silicon dichloride**
8. dihydrogen monoxide **H_2O**
9. xenon tetrafluoride **XeF_4**
10. diiodine pentoxide **I_2O_5**
11. chlorine monoxide **ClO**
12. carbon tetrafluoride **CF_4**
13. tetraphosphorus triselenide **P_4Se_3**
14. methane **PBr_5**

Ionic or Molecular?

For each of the following, indicate whether it is an **ionic compound (I)** or a **molecule (M),** and give the corresponding formula or name as appropriate.

	I or M	Chemical Formula/Name
1. aluminum acetate	**I**	**$Al(CH_3COO)_3$**
2. silver phosphate	**I**	**Ag_3PO_4**
3. ammonia	**I**	**$Cu(OH)_2$**
4. copper (II) hydroxide	**I**	**$(NH_4)_2SO_4$**
5. ammonium sulfate	**I**	**$Ca(ClO_4)_2$**
6. calcium perchlorate	**M**	**PCl_3**
7. phosphorus trichloride	**I**	**TiO_2**
8. titanium (IV) oxide	**I**	**$Zn(ClO)_2$**
9. zinc hypochlorite	**I**	**Sodium oxide**
10. diarsenic pentasulfide	**M**	**Tetraphosphorus heptoxide**
11. Na_2O_2	**I**	**Mercury (I) carbonate**
12. P_4O_7	**I**	**Nickel (III) hydroxide**
13. Hg_2CO_3	**M**	**Dinitrogen pentoxide**
14. $Ni(OH)_3$	**I**	**Tin (IV) chromate**
15. N_2O_5	**M**	**Sulfur trioxide**
16. $Sn(CrO_4)_2$		
17. SO_3		
18. Mg_3N_2		
19. GeO_2		
20. $Al_2(SO_4)_3$		

The Mole

PART 1

Here are some common counting units relating to the numbers of things.

1 dozen = 12 **1 score = 20** **1 ream = 500** **1 gross = 144**

Practice converting between units of measure and the items they represent. Use dimensional analysis.

1. 24 eggs = **2** dozen eggs
2. 4 score of years = **80** years
3. 750 sheets = **1.5** reams of paper
4. 2 gross of pencils = **288** pencils

PART 2

In chemistry, it is convenient to group small particles (atoms, molecules, ions) into a counting unit called the **mole.**

1 mole = 6.022×10^{23} objects (This number is called Avogadro's number.)

When you use your calculator, try using the EE, E or EXP function as a shortcut for the exponent in scientific notation " × 10 ^ " (times 10 to the power of). ***Use dimensional analysis for each one.***

5. 24 atoms of sodium = **4.00×10^{-23}** moles of sodium atoms

$$\begin{array}{c|c} \textbf{24 atoms Na} & \textbf{1 mol Na} \\ \hline & \textbf{6.02} \times \textbf{10}^{23}\textbf{ atoms} \end{array}$$

6. 15.0 moles of Pb^{2+} ions = **9.03×10^{24}** Pb^{2+} ions

$$\begin{array}{c|c} \textbf{15 mol Pb}^{2+}\textbf{ ions} & \textbf{6.02} \times \textbf{10}^{23}\textbf{ ions Pb}^{2+} \\ \hline & \textbf{1 mol Pb}^{2+}\textbf{ ion} \end{array}$$

7. 950 atoms of silver = **1.6×10^{-21}** moles of silver atoms
2 s.f.

$$\begin{array}{c|c} \textbf{950 atoms Ag} & \textbf{1 mol Ag} \\ \hline & \textbf{6.02} \times \textbf{10}^{23}\textbf{ atoms} \end{array}$$

8. 5.36 moles of NH_3 = **3.23×10^{24}** molecules of NH_3 molecules

$$\begin{array}{c|c} \textbf{5.36 mol} & \textbf{6.02} \times \textbf{10}^{23}\textbf{ mcs} \\ \hline & \textbf{1 mol} \end{array}$$

9. 5 molecules of chlorine (Cl_2) gas = **8×10^{-24}** moles of chlorine (Cl_2) molecules
1 s.f.

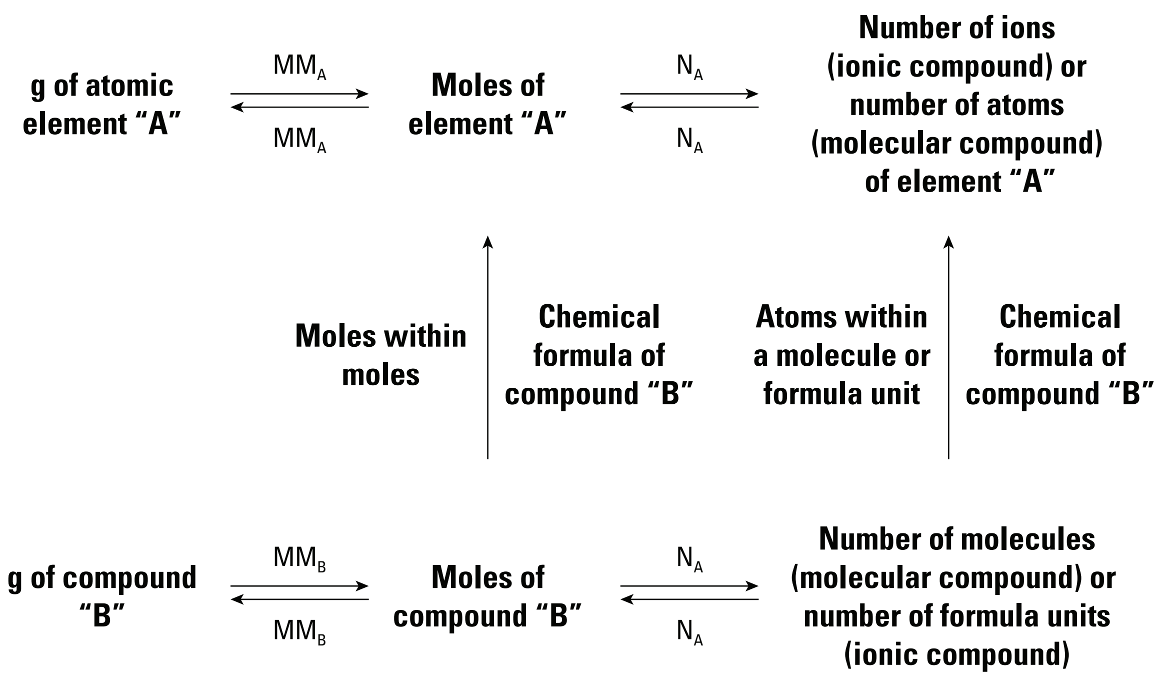
$$\begin{array}{c|c} \textbf{5 mcs} & \textbf{1 mol} \\ \hline & \textbf{6.02} \times \textbf{10}^{23}\textbf{ mcs} \end{array}$$

10. 4.00 moles of barium atoms = **2.41×10^{24}** atoms of barium

$$\begin{array}{c|c} \textbf{4 mol Ba} & \textbf{6.02} \times \textbf{10}^{23}\textbf{ atoms} \\ \hline & \textbf{1 mol Ba} \end{array}$$

11. Why is a mole such a big counting unit compared to ones we use every day (such as the dozen)?
Because atoms and molecules are very small, there has to be a large number of them to be visible to the naked eye.

Mass, Moles, Avogadro's Number, and Moles Within Moles Conversions



ABBREVIATIONS USED

MM = Molar mass

N_A = Avogadro's number = 6.02×10^{23} items

Moles = Mol

Molecules = mc/mcs

Molar Mass

- The molar mass of a substance = the mass (in grams) of one mole of the substance.
- One mole of an element = the atomic mass of that element (on the periodic table) in grams.
- One mole of a compound = the sum of the atomic masses of the atoms present in the compound in grams.
  - Example: $H_2O = (2 \times 1.008) + 16.00 = 18.016$ g/mol.
- **The units of molar mass are always grams per mole (g/mol).**
- **NOTE:** "Mole" may be abbreviated "mol," but not "m" ("m" means meter).

1. What is the mass of one mole (molar mass) of Ar? **39.95 g Ar/mol**

2. What is the molar mass of Na? **22.99 g Na/mol**

3. What is the mass of one mole (molar mass) of H_2O? **18.016 g H_2O/mol**

   2 H = 2 × 1.008 = 2.016
   1 O → = 16
   18.016

4. What is the molar mass of NaCl? **22.99 + 35.45 = 58.44 g/mol**

5. How many moles of H_2O are in 22.5 g of H_2O?

   $$? \text{ mol } H_2O = \frac{22.5 \text{ g } H_2O}{} \left| \frac{1 \text{ mol } H_2O}{18.016 \text{ g } H_2O} \right. = 1.25 \text{ mol } H_2O$$

6. How many grams are in 0.250 moles of NaCl?

   $$? \text{ g NaCl} = \frac{0.250 \text{ mol NaCl}}{} \left| \frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} \right. = 14.6 \text{ g NaCl}$$

7. What is the molar mass of C_2H_5OH (ethanol)?

   2 C 2 × 12.01 = 24.02
   6 H 6 × 1.008 = 6.048
   O = 16.00
   46.068 ≈ 46.07 g/mol

8. How many moles are in 25.0 mL of ethanol, C_2H_5OH (the density of ethanol is 0.785 g/mL)? (Answer: 0.426 mol)

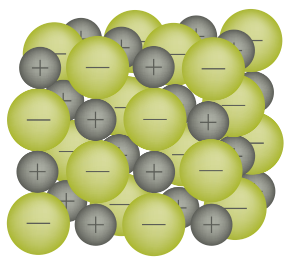
   $$? \text{ mol } C_2H_5OH = \frac{25.0 \text{ mol } C_2H_5OH}{} \left| \frac{0.785 \text{ g} = 19.625 \text{ g}}{1 \text{ mL}} \right| \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g}} = 0.426 \text{ mol } C_2H_5OH$$

Types of Substances

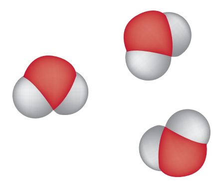
Most elements occur as monatomic elements (single atoms, such as Cu or Na). Some elements are diatomic and exist as molecules of two atoms each (BriNClHOF). Other elements may be molecular as well (S_8 or C_{60})—but you don't need to know those.

Compounds are either ionic or covalent. If ionic, their basic unit is the "formula unit" because they consist of extending arrays of ions in a crystal lattice (diagram below, left). Covalent or molecular compounds have a basic unit of "molecules" because they are discrete or separated (diagram below, right).

NaCl crystal of Na^+ and Cl^- ions exist as ***formula units*** of "NaCl" (simplest ratio).



Water ***molecules*** which exist as "H_2O."



The mass of one atom, molecule, or formula unit is most appropriate in units of atomic mass units (amu). The molar mass of these substances is in units of grams per mole. The numbers are from the atomic masses on the periodic table.

Find the formula and molar mass of the following. Use at least two decimal places to ensure enough significant figures for most problems. Indicate the simplest unit as atom, molecule, or formula unit. The first is completed as an example.

Substance	Formula Mass (amu)	Molar Mass (g/mol)	Simplest Unit
CO_2	44.01 amu	44.01 g/mol	Molecule
$Al_2(SO_4)_3$	**342.17 amu**	**342.17 g/mol**	**Formula unit**
$C_6H_{10}O_5$	**162.14 amu**	**162.14 g/mol**	**Molecule**
Br_2	**159.8 amu**	**159.8 g/mol**	**Molecule**
$CaCl_2$	**110.98 amu**	**110.98 g/mol**	**Formula unit**
H_2O	**18.016 amu**	**18.016 g/mol**	**Molecule**
$CaCl_2 \cdot 2\ H_2O$ (a hydrate; dot means to add 2 H_2O)	**147.01 amu**	**147.01 g/mol**	**Formula unit**

Moles, Grams, Number of Particles, and Elements in Compounds

Complete this concept map by filling in "Avogadro's number" or "Molar Mass" over the arrows.

# of particles (atom, ions, molecules, formula units) $\longleftrightarrow$ **moles** $\longleftrightarrow$ **mass (grams)**

1. How many moles of magnesium are there in 10.00 g?

$$? \text{ mol Mg} = \frac{10.00 \text{ g Mg}}{} \bigg| \frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}} = 0.4114 \text{ mol Mg}$$

2. How many grams of iron are there in 34.77 moles of iron?

$$? \text{ g Fe} = \frac{34.77 \text{ mol Fe}}{} \bigg| \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}} = 1942 \text{ g Fe}$$

3. How many moles are there in 3.493 grams of lye, NaOH?

MM NaOH

Na = 22.99

O = 16.00

H = 1.008

39.998 ≈ 40.00 g

$$? \text{ mol NaOH} = \frac{3.493 \text{ g NaOH}}{} \bigg| \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} = 0.08733 \text{ mol NaOH}$$

4. How many grams are there in 0.275 moles of Na_3PO_4?

MM Na_3PO_4

3 Na $3 \times 22.99 = 68.97$

1 P $= 30.97$

4 O $4 \times 16.0 = 64.00$

163.94

$$? \text{ g } Na_3PO_4 = \frac{0.275 \text{ mol } Na_3PO_4}{} \bigg| \frac{163.94 \text{ g } Na_3PO_4}{1 \text{ mol } Na_3PO_4} = 45.1 \text{ g } Na_3PO_4$$

5. How many molecules are in 7.47 moles of NH_3? How many total atoms are there? (Answer: 1.80×10^{25} atoms)

4.50×10^{24} mcs NH^3

$$? \text{ mcs } NH_3 \text{ (\& atoms)} = \frac{7.47 \text{ mol } NH_3}{} \left| \frac{6.02\text{ E}23}{1 \text{ mol } NH_3} \right| \frac{4 \text{ atoms total}}{1 \text{ mc } NH_3} = 1.80 \times 10^{25} \text{ total atoms}$$

6. What is the mass in grams of 1.000×10^{12} (1.000 trillion) atoms of gold?

$$? \text{ g Au} = \frac{1.000 \times 10^{12} \text{ atoms Au}}{} \left| \frac{1 \text{ mol Au}}{6.02\text{ E}23} \right| \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 3.272 \times 10^{-10} \text{ g Au}$$

7. How many grams of ammonium nitrate, NH_4NO_3, contain 8.3×10^{25} formula units?

$$? \text{ g } NH_4NO_3 = \frac{8.3 \times 10^{25} \text{ form. units}}{} \left| \frac{1 \text{ mol } NH_4NO_3}{6.02\text{ E}23 \text{ form. units}} \right| \frac{80.05 \text{ g } NH_4NO_3}{1 \text{ mol } NH_4NO_3} = 1.1 \times 10^{4} \text{ g } NH_4NO_3$$

MM NH_4NO_3

2N = 2 × 14.01 = 28.02

4H = 4 × 1.008 = 4.032

30 = 3 × 16.00 = 48.00

80.052

8. How many molecules of water are there in 0.034 grams of water?

$$? \text{ mcs } H_2O = \frac{0.034 \text{ g } H_2O}{} \left| \frac{1 \text{ mol } H_2O}{18.016 \text{ g } H_2O} \right| \frac{6.02 \times 10^{23} \text{ mcs } H_2O}{1 \text{ mol } H_2O} = 1.1 \times 10^{21} \text{ mcs } H_2O$$

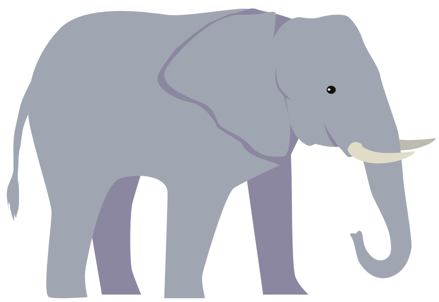
Moles: Calculate the Missing Quantities

Fill in the blanks.

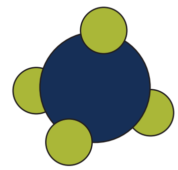
You need to use 1–2 extra digits (beyond # of sig figs in given quantity) in your molar mass to not cause a rounding error.

Chemical Formula	Molar Mass (g/mol)	Mass	Moles	# of Particles (atoms, molecules, or formula units)	Type of Substance (monatomic element, molecular element, covalent compound, ionic compound)
N_2	**28.02 g/mol**	**7.3×10^{-3} g**	2.6×10^{-4} moles	**1.6×10^{20} mcs**	**Molecular element**
CH_3OH	**32.05 g/mol**	1.25 g	**0.0390 mol**	**2.35×10^{22} mcs**	**Covalent or molecular compound**
Na	**22.99 g/mol**	**3119 g Na**	**135.7 mol Na**	8.170×10^{25} Na atoms	**Monatomic element**
CaF_2	**78.08 g/mol**	**6.59 g**	0.0844 moles	**5.08×10^{22} formula units**	**Ionic compound**
KNO_3	**101.11 g/mol**	19.3 kg	**191 mol**	**1.15×10^{26} formula units**	**Ionic compound**
Br_2	**159.81 g/mol**	0.4395 g	2.75×10^{-3} moles	**1.66×10^{21} mcs**	diatomic molecular element
Kr	**83.7 g/mol**	314 g	**3.75 mol**	2.26×10^{24} atoms	monatomic element

Mole Ratios Using Chemical Formulas



An elephant has
one trunk and four legs.



A molecule of methane, CH_4
contains one carbon and four hydrogens.

A ratio can be constructed for these two objects in the following ways:

$$\frac{\textbf{4 legs}}{\textbf{1 elephant}} \qquad \frac{\textbf{4 H atoms}}{\textbf{1 } \mathbf{CH_4} \textbf{ molecule}}$$

1. Can you think of another ratio for each of the above? Write it next to the ones above.

2. Now picture a dozen elephants and a dozen methane molecules. What ratios can you construct using the term "dozen" in the numerator and denominator? Fill in the numerators below:

$$\frac{\textbf{4 dozen legs}}{\textbf{1 } \textbf{\textit{dozen}} \textbf{ elephants}} \qquad \frac{\textbf{4 dozen "H" atoms}}{\textbf{1 } \textbf{\textit{dozen}} \ \mathbf{CH_4} \textbf{ molecules}}$$

3. Now picture you have a mole (rather than a dozen) elephants and methane molecules. Fill in the numerator now:

$$\frac{\textbf{4 mol legs}}{\textbf{1 } \textbf{\textit{mole}} \textbf{ of elephants}} \qquad \frac{\textbf{4 mol "H" atoms}}{\textbf{1 } \textbf{\textit{mole}} \textbf{ of } \mathbf{CH_4} \textbf{ molecules}}$$

4. Practice constructing a ***mole ratio*** between the indicated element in ***one mole*** of each of the following compounds (first one is completed as an example):

   a. Oxygen atoms (O) in CO_2
   Answer: 2 mole O atoms/1 mole CO_2

   b. Carbon in Glucose ($C_6H_{12}O_6$)
   6 mol C/1 mol $C_6H_{12}O_6$, glucose

   c. Chloride ions (Cl^-) in $CaCl_2$
   2 mol Cl^-/1 mol $CaCl_2$

5. Practice using the conversions above for the given amounts. Use dimensional analysis!

   a. Moles of oxygen (O) in 4 moles of CO_2

$$? \text{ mol O} = \begin{array}{c|c} 4\text{ mol CO}_2 & 2\text{ mol O} \\ \hline & 1\text{ mol CO}_2 \end{array} = 8\text{ mol O}$$

   b. Moles of carbon in 0.45 moles of glucose, $C_6H_{12}O_6$

$$\text{mol C} = \begin{array}{c|c} 0.45\text{ mol glu} & 6\text{ mol C} \\ \hline & 1\text{ mol glu} \end{array} = 2.7\text{ mol O}$$

   c. Moles of chloride ions in 5.00 grams of $CaCl_2$

$$\text{mol Cl}^- = \begin{array}{c|c|c} 5.00\text{ g CaCl}_2 & 1\text{ mol CaCl}_2 & 2\text{ mol Cl}^- \\ \hline & 110.98\text{ g} & 1\text{ mol CaCl}_2 \end{array} = 0.0901\text{ mol Cl}^-$$

Practicing Mole Ratios

1. One mole of carbon dioxide contains **1** moles of C and **2** moles of O.

   Rewrite this statement as two mole ratios: $\frac{1\text{ mol C}}{1\text{ mol }CO_2}$, $\frac{2\text{ mol O}}{1\text{ mol }CO_2}$

2. 88.02 grams of carbon dioxide is **2 mole(s)** of CO_2, which is **2 mole(s)** of carbon atoms, and **4 mole(s)** of oxygen atoms. $\frac{88.02\text{ g }CO_2 \mid 1\text{ mol }CO_2}{\mid 44.01\text{ g }CO_2} = 2\text{ mol }CO_2$ $\frac{2\text{ mol }CO_2 \mid 1\text{ mol C}}{\mid 1\text{ mol }CO_2} = 2\text{ mol C}$

3. a. Perform this calculation, using some information you already provided in the previous question.

   88.02 grams of carbon dioxide is also **24.02** grams of carbon atoms and **64.00** grams of oxygen atoms.

   Show your work: $\frac{2\text{ mol }CO_2 \mid 12.01\text{ g C}}{\mid 1\text{ mol }CO_2} = 24.02\text{ g C}$ $\frac{2\text{ mol }CO_2 \mid 32.00\text{ g O}}{\mid 1\text{ mol }CO_2} = 64.00\text{ g O}$

   b. Do the grams of carbon atoms and grams of oxygen atoms add up to the grams of carbon dioxide? Should they?

   Yes, 24.02 g C + 64.00 g O = 88.02 g CO_2. Law of conservation of mass

4. Indicate whether each of the following statements is true or false, and **explain** your reasoning.

   a. One mole of NH_3 weighs more than one mole of H_2O. **False, weighs less, see calc.**

   NH_3 14.01 H_2O 2H = 2 × 1.008 = 2.016
   3 × 1.008 3.024   O = 16.000
   17.03 g/mol   18.016 g/mol

   b. There are more carbon atoms in 48 grams of carbon dioxide than in 12 grams of diamond (pure carbon). **True**

   $\frac{12\text{ g C} \mid 1\text{ mol C}}{\mid 12.01\text{ g C}} \approx 1\text{ mol C}$ $\frac{48\text{ g }CO_2 \mid 1\text{ mol }CO_2 \mid 1\text{ mol C}}{\mid 44.01\text{ g }CO_2 \mid 1\text{ mol }CO_2} = 1.09\text{ mol C or }1.1\text{ mol C}$

   c. There are equal numbers of nitrogen atoms in one mole of NH_3 and one mole of N_2.
   False. 1 mol NH_3 has 1 mol N atoms, 1 mol N_2 has 2 mol N atoms.

   d. The number of Cu atoms in 100 grams of Cu (*s*) is the same as the number of Cu atoms in 100 grams of copper (II) oxide, CuO. **False**

   $\frac{100\text{ g Cu} \mid 1\text{ mol Cu}}{\mid 63.55\text{ g Cu}} = 1.57\text{ mol Cu}$ $\frac{100\text{ g CuO} \mid 1\text{ mol CuO} \mid 1\text{ mol C}}{\mid 79.55\text{ g CuO} \mid 1\text{ mol CuO}} = 1.26\text{ mol Cu}$

   e. The number of Ni atoms in 100 moles of Ni (*s*) is the same as the number of Cl atoms in 100 moles of nickel (II) chloride, $NiCl_2$. **False**

   100 mol Ni versus $\frac{100\text{ mol }NiCl_2 \mid 2\text{ mol Cl}}{\mid 1\text{ mol }NiCl_2} = 200\text{ mol Ni}$

   f. There are more hydrogen atoms in 2 moles of NH_3 than in 2 moles of CH_4. **False**
   has 6 mol "H"   **has 8 mol "H"**

Moles within Moles

We can interpret 1 mole of a compound or molecules in terms of moles of elements.

EXAMPLES

- 1 mole of $(NH_4)_2SO_4$ = 2 moles of nitrogen
- 1 mole of $(NH_4)_2SO_4$ = 8 moles of hydrogen
- 1 mole of $(NH_4)_2SO_4$ = 1 mole of sulfur
- 1 mole of $(NH_4)_2SO_4$ = 4 moles of oxygen

We can use the analogy of an elephant. We can say 1 mole of elephants contains 1 mole of trunks, 2 moles of ears and 4 moles of legs.

1. How many moles of oxygen atoms are there in 2 moles of KNO_3? (6 mol)

$$? \text{ mol O} = \frac{2 \text{ mol } KNO_3 \;\Big|\; 3 \text{ mol O}}{\;\Big|\; 1 \text{ mol } KNO_3} = 6 \text{ mol O}$$

2. How many hydrogen atoms are there in 2.50 moles of $(NH_4)_2SO_4$ formula units? (1.20×10^{25} H atoms)

$$? \text{ H atoms} = \frac{2.50 \text{ mol } (NH_4)_2SO_4 \;\Big|\; 8 \text{ mol H} \;\Big|\; 6.02 \times 10^{23} \text{ atoms H}}{\;\Big|\; 1 \text{ mol } (NH_4)_2SO_4 \;\Big|\; 1 \text{ mol H}} = 1.20 \times 10^{25} \text{ H atoms}$$

3. A mole of H_2O and a mole of O_2

   a. have the same mass.

   b. contain one molecule each.

   c. have a mass of 1 g each.

   d. contain the same number of molecules.

4. One molecule of phosphorus contains 4 phosphorus atoms. Then one mole of phosphorus molecules will contain

   a. 4 g of phosphorus.

   b. 4 moles of phosphorus atoms.

   c. 6.02×10^{23} phosphorus atoms.

   d. 4 phosphorus atoms.

5. How many nitrogen atoms are there in 48.0 g of N and 48.0 g of N_2?

$$\#\ \text{O atom} = \frac{48.0\text{ g O}}{} \left| \frac{1\text{ mol O}}{16.00\text{ g O}} \right| \frac{6.02 \times 10^{23}}{1\text{ mol O}} = 1.81 \times 10^{24}\text{ atoms O}$$

$$\#\ \text{O atom} = \frac{48.0\text{ g }O_2}{} \left| \frac{1\text{ mol }O_2}{32.00\text{ g }O_2} \right| \frac{2\text{ mol O}}{1\text{ mol }O_2} \left| \frac{6.02 \times 10^{23}\text{ atoms}}{1\text{ mol O}} \right. = 1.81 \times 10^{24}\text{ atoms O}$$

6. How many atoms of oxygen are there in 0.0327 moles of Na_2SO_4? (**HINT:** There are 4 moles of O in each mole of Na_2SO_4.) (Answer: 7.87×10^{22} atoms)

$$?\ \text{atoms O} = \frac{0.0327\text{ mol }Na_2SO_4}{} \left| \frac{4\text{ mol O}}{1\text{ mol }Na_2SO_4} \right| \frac{6.02 \times \text{E23}}{1\text{ mol O}} = 7.87 \times 10^{22}\text{ O atoms}$$

7. How many atoms of nitrogen are there in 35.0 moles of $(NH_4)_2SO_4$? (**HINT:** How many moles of N are in each mole of the compound?) (Answer: 4.21×10^{25} atoms)

$$?\ \text{atoms N} = \frac{35.0\text{ mol }(NH_4)_2SO_4}{} \left| \frac{2\text{ mol N}}{1\text{ mol }(NH_4)_2SO_4} \right| \frac{6.02\text{ E23}}{1\text{ mol N}} = 4.21 \times 10^{25}\text{ N atoms}$$

8. How many grams of oxygen are there in 43.6 grams of $CaCO_3$? (**HINT:** How many moles of O are in a mole of calcium carbonate?) (Answer:20.9 g)

$$?\ \text{g O} = \frac{43.6\text{ g }CaCO_3}{} \left| \frac{1\text{ mol }CaCO_3}{100.09\text{ g }CaCO_3} \right| \frac{3\text{ mol O}}{1\text{ mol }CaCO_3} \left| \frac{16.00\text{ g O}}{1\text{ mol O}} \right. = 20.9\text{ g Oxygen}$$

Ca = 40.08
C = 12.01
3O = 3 × 16.00 = 48.00
109.09

Percent Composition by Mass

The **percent composition** is the percent by mass of each element in a given mass of compound.

$$\text{\% by mass} = \frac{\textit{mass of elements in one mole of the compound}}{\textit{mass of one mole of the compound (molar mass)}} \times 100\%$$

Calculate the percent composition of hydrogen in the following compounds:

1. C_2H_6

$2C = 2 \times 12.01 = 24.02$
$6H = 6 \times 1.008 = 6.048$
30.068 g/mol

$$\% C = \frac{24.02 \text{ g C} \mid 100}{30.068 \text{ g } C_2H_6 \mid} = 79.89\% \text{ C}$$

$$\% H = 100 - 79.89\% = 20.11\% \text{ H}$$

2. $Ca(C_2H_3O_2)_2$

$Ca = 40.08$
$4C = 48.04$
$6H = 6.048$
$4O = 64.00$
158.168 g/mol

$$\% Ca = \frac{40.08 \text{ g Ca} \mid 100}{158.168 \text{ g} \mid} = 25.34\% \text{ Ca}$$

$$\% H = \frac{6.048 \text{ g H} \mid 100}{158.168 \text{ g} \mid} = 3.82\%$$

$$\% C = \frac{48.04 \text{ g C} \mid 100}{158.168 \text{ g} \mid} = 30.37\% \text{ C}$$

$$\% O = 100 - (25.34 - 30.37 - 3.82)$$

3. H_2O

2.016 g H
16.00 O
18.016 g/mol

$$\frac{2.016 \text{ g H} \mid 100}{18.016 \text{ g} \mid} = 11.2\% \text{ H}$$

$$\% O = 100 - 88.8\% \text{ O}$$

Calculate the mass of the specified element in the given mass of compound:

1. Mass of hydrogen in 350. g C_2H_6 (Answer: 70.5 g)

$$? \text{ g H} = \frac{350 \text{ g } C_2H_6 \mid 1 \text{ mol } C_2H_6 \mid 6 \text{ mol H} \mid 1.01 \text{ g H}}{\mid 30.068 \text{ g} \mid 1 \text{ mol } C_2H_6 \mid 1 \text{ mol H}} = 70.5 \text{ g H}$$

2. Mass of hydrogen in 124 g of $Ca(C_2H_3O_2)_2$ (Answer: 4.75 g)

$$? \text{ g H} = \frac{124 \text{ g } Ca(C_2H_3O_2)_2 \mid 1 \text{ mol } Ca(C_2H_3O_2)_2 \mid 6 \text{ mol H} \mid 1.01 \text{ g H}}{\mid 158.168 \text{ g } Ca(C_2H_3O_2)_2 \mid 1 \text{ mol } Ca(C_2H_3O_2)_2 \mid 1 \text{ mol H}}$$

3. Mass of oxygen in 100. g H_2O (Answer: 88.8 g)

$$? \text{ g O} = \frac{100 \text{ g } H_2O \mid 1 \text{ mol } H_2O \mid 1 \text{ mol O} \mid 16.00 \text{ g O}}{\mid 18.016 \text{ g } H_2O \mid 1 \text{ mol } H_2O \mid 1 \text{ mol O}} = 88.8 \text{ g O}$$

Empirical and Molecular Formulas

REVIEW

The **empirical formula** is the simplest whole-number ratio of elements in a compound. The **molecular formula** is the chemical formula of the chemical compound of interest.

1. Given the molecular formulas below, determine the empirical formula.
   a. Ribose, $C_5H_{10}O_5$, a sugar molecule in RNA. **$C_1H_2O_1$**
   b. Ethyl butanoate, $C_6H_{12}O_2$, a cmpd w/ the odor of pineapple. **$C_3H_6O_1$**
   c. Chlorophyll, $C_{55}H_{72}MgN_4O_5$, part of photosynthesis. **$C_{55}H_{77}MgN_4O_5$**
   d. DEET, $C_{12}H_{17}ON$, an insect repellent. **$C_{12}H_{17}ON$**
   e. Oxalic acid $H_2C_2O_4$, found in spinach and tea. **$H_1C_1O_2$**
2. True or False? The empirical formula of a compound can be the same as its molecular formula. **True. For example, refer to c and d molecules in 1 above.**
3. Determine the empirical formulas for each of the following compounds based on the information provided.
   a. A sample of a compound contains 2.23 g Fe and 1.93 g S. (Answer: Fe_2S_3)

$$\frac{2.23\text{ g Fe}}{} \left| \frac{1\text{ mol Fe}}{55.85\text{ g Fe}} \right. = 0.0399\text{ mol Fe} \quad \div 0.0399 = 1\text{ mol Fe}$$

$$\frac{1.93\text{ g S}}{} \left| \frac{1\text{ mol S}}{32.07\text{ g S}} \right. = 0.0602 \quad \div 0.0399 = 1.5\text{ mol S}$$

1 mol Fe × 2 : 1 mol S × 2

2 mol Fe : 3 mol S

∴ EF is Fe_2S_3

   b. A compound is 63.11% C, 11.92% H and 24.97% F. (Answer: C_4H_9F)

$$\frac{63.11\text{ g C}}{} \left| \frac{1\text{ mol C}}{12.01\text{ g C}} \right. = 5.255\text{ mol C}/1.314 = 4\text{ mol C}$$

$$\frac{11.92\text{ g H}}{} \left| \frac{1\text{ mol H}}{1.008\text{ g H}} \right. = 11.83\text{ mol H}/1.314 = 9\text{ mol H}$$

$$\frac{24.97\text{ g F}}{} \left| \frac{1\text{ mol F}}{19.00\text{ g F}} \right. = 1.314\text{ mol F}/1.314 = 1\text{ mol F}$$

∴ EFis C_4H_9F

   c. A compound is 40.0% C, 6.67% H and 53.3% O. (Answer: CH_2O)

$$\frac{40.0\text{ g C}}{} \left| \frac{1\text{ mol C}}{12.01\text{ g C}} \right. = 3.33\text{ mol C}/3.33 = 1\text{ mol C}$$

$$\frac{6.67\text{ g H}}{} \left| \frac{1\text{ mol H}}{1.008\text{ g H}} \right. = 6.60\text{ mol H}/3.33 = 2\text{ mol H}$$

$$\frac{53.3\text{ g O}}{} \left| \frac{1\text{ mol O}}{16.00\text{ g O}} \right. = 3.33\text{ mol O}/3.33 = 1\text{ mol F}$$

∴ EF is CH_2O

Concept Question: This compound (Part c above) has a molar mass of 60 g/mol. What is its molecular formula?

Emperical formula mass CH_2O = 12.01 + 2.016 + 16 = 30.0 g

$$n = \frac{MM}{EFM} = \frac{60}{30} = 2 \quad \therefore \text{MF is } (CH_{2(\times 2)}O) = \text{Molecular formula} = C_2H_4O_2$$

Review: Moles and Chemical Composition

Use significant figures and show your work!

1. How many calcium cations are in 1.6×10^{-3} moles of Ca^{2+}?

$$? \ Ca^{2+} = \frac{1.6 \times 10^{-3} \text{ mol } Ca^{2+}}{} \left| \frac{6.02\text{ E23}}{1 \text{ mol } Ca^{2+}} \right. = 9.6 \times 10^{20} \ Ca^{2+} \text{ cations}$$

2. How many moles of argon atoms would be in a balloon containing 8.19×10^{24} atoms of Ar?

$$? \text{ mol Ar} = \frac{8.19 \times 10^{24} \text{ atoms Ar}}{} \left| \frac{1 \text{ mol Ar}}{6.02\text{ E23}} \right. = 13.6 \text{ mol Ar}$$

3. How many atoms of C are in 1.22 moles of C_6H_5Cl? (Answer: 4.41×10^{24} atoms)

$$? \text{ atoms C} = \frac{1.22 \text{ mol } C_6H_5Cl}{} \left| \frac{6 \text{ mol C}}{1 \text{ mol } C_6H_5Cl} \right| \frac{6.02\text{ E23 cations}}{1 \text{ mol C}} = 4.41 \times 10^{24} \text{ C atoms}$$

4. What is the mass of 1.31×10^{28} molecules of water (H_2O)? (Answer: 3.92×10^{5} g)

$$? \text{ g } H_2O = \frac{1.31 \times 10^{28} \text{ mcs } H_2O}{} \left| \frac{1 \text{ mol } H_2O}{6.02\text{ E23 mcs } H_2O} \right| \frac{18.016 \text{ g } H_2O}{1 \text{ mol } H_2O} = 3.92 \times 10^{5} \text{ g } H_2O$$

5. How many molecules of carbon dioxide are in 0.550 g? (Answer: 7.53×10^{21} mcs)

$$? \text{ mcs } CO_2 = \frac{0.550 \text{ g } CO_2}{} \left| \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \right| \frac{6.02\text{ E23 mcs } CO_2}{1 \text{ mol } CO_2} = 0.0753 \times 10^{23} \text{ mcs}$$

or 7.53×10^{21} mcs CO_2

6. How many moles of O are in 8.0 g of N_2O_3? (Answer: 0.32 mol)

$$? \text{ mol O} = \frac{8.0 \text{ g } N_2O_3}{} \left| \frac{1 \text{ mol } N_2O_3}{76.02 \text{ g } N_2O_3} \right| \frac{3 \text{ mol O}}{1 \text{ mol } N_2O_3} = 0.32 \text{ mol O}$$

$2N = 2 \times 14.01 = 28.02$
$3O = 3 \times 16.00 = 48.00$
76.02 g/ mol

7. How many F atoms are there in 11.2 g of the chlorofluorocarbon, CF_2Cl_2? (Answer: 1.12×10^{23} atoms)

$$? \text{ F atoms} = \frac{11.2 \text{ g } CF_2Cl_2}{} \left| \frac{1 \text{ mol } CF_2Cl_2}{120.91 \text{ g } CF_2Cl_2} \right| \frac{2 \text{ mol F}}{1 \text{ mol } CF_2Cl_2} \left| \frac{6.02 \times 10^{23} \text{ atoms F}}{1 \text{ mol F}} \right. = 1.12 \times 10^{23} \text{ F atoms}$$

MM $CF_2Cl_2 = 120.91$ g/mol

8. What is the % composition of each element in CF_2Cl_2?

12.01 g C
38.00 g F
70.90 g Cl
120.91 g/mol

$$\% \text{ C} = \frac{12.01 \text{ g C}}{120.91 \text{ g total}} \left| \frac{100}{} \right. = 9.930\% \text{ C}$$

$$\% \text{ F} = \frac{38.00 \text{ g F}}{120.91 \text{ g total}} \left| \frac{100}{} \right. = 31.43\% \text{ F}$$

$$\% \text{ Cl} = \frac{70.90 \text{ g Cl}}{120.91 \text{ g total}} \left| \frac{100}{} \right. = 56.64\% \text{ Cl}$$

9. A compound that contains only C, H, and O is 48.38% C and 8.12% H. Determine the empirical formula. (Answer: $C_3H_6O_2$)

$$C = \frac{48.38 \text{ g C}}{} \left| \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right. = 4.028 \text{ mol C} = 1.5 \text{ mol C} \times 2 = 3 \text{ mol C}$$

$$H = \frac{8.12 \text{ g H}}{} \left| \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right. = 8.04 \text{ mol H} = 3 \text{ mol H} \times 2 = 6 \text{ mol H}$$

Emperical formula = $C_3H_6O_2$

$$O = \frac{43.5 \text{ g O}}{} \left| \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right. = 2.72 \text{ mol O} = 1 \text{ mol O} \times 2 = 2 \text{ mol O}$$

10. Determine the empirical formula for a compound that gives the following composition: 71.65% Cl; 24.27% C; 4.07% H. The formula mass is known to be 98.96 amu. What is its molecular formula? (Answer: $C_2H_4Cl_2$)

$$\frac{71.65 \text{ g Cl}}{} \left| \frac{1 \text{ mol Cl}}{35.45 \text{ g Cl}} \right. = 2.02 \text{ mol Cl} \div 2.02 = 1 \text{ mol Cl}$$

$$\frac{24.27 \text{ g C}}{} \left| \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right. = 2.02 \text{ mol C} \div 2.02 = 1 \text{ mol C}$$

EF mass = CH_2Cl

35.45
12.01
2.02
49.48 g/mol

$$\frac{4.07 \text{ g H}}{} \left| \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right. = 4.03 \text{ mol H} \div 2.02 = 2 \text{ mol H}$$

$$n = \frac{MM}{EF \text{ mass}} = \frac{98.96}{49.48} = 2 \quad \therefore \text{ MF is } C_2H_4Cl_2$$

Review: Nomenclature and Chemical Composition

1. Identify the following as ionic or molecular compounds.

   a. FeS ionic / molecular

   b. SO_2 ionic / molecular

   c. CH_3OH ionic / molecular

   d. $ZnSO_4$ onic / molecular

2. Fill in the blanks with names or formulas for the compounds.

   a. $CaCl_2$ **Calcium chloride**

   b. N_2H_4 **Dinitrogen tetrahydride**

   c. $CuBr_2$ **Copper (II) bromide**

   d. sodium phosphate **Na_3PO_4**

   e. chromium (III) chloride **$CrCl_3$**

   f. boron trifluoride **BF_3**

3. Calculate the mass in grams of 5.94×10^{20} H_2O_2 molecules. (Answer: 0.0336 g)

$$? \text{ g } H_2O_2 = \frac{\text{given } 5.94 \times 10^{20} \text{ } H_2O_2 \text{ mcs}}{} \left| \frac{1 \text{ mol } H_2O_2}{6.02\text{ E23}} \right| \frac{34.02 \text{ g } H_2O_2}{1 \text{ mol } H_2O_2} = 0.0336 \text{ g } H_2O_2$$

or 3.36×10^{-2} g

4. Calculate the mass percent of each element present in aspirin, $C_9H_8O_4$.

MM Calc

$9 \text{ C} = 9 \times 12.01 = 108.09 \text{ g C}$

$8 \text{ H} = 8 \times 1.01 = 8.08 \text{ g H}$

$4 \text{ O} = 4 \times 16.00 = 64.00 \text{ g O}$

Aspirin 180.17 g/mol

$$\% \text{ C} = \frac{108.09 \text{ g C}}{180.17 \text{ g total}} \left| \frac{100 \text{ g total}}{} \right. = 59.99\% \text{ C}$$

$$\% \text{ H} = \frac{8.08 \text{ g H}}{180.17 \text{ g total}} \left| \frac{100 \text{ g total}}{} \right. = 4.48\% \text{ H}$$

$$\% \text{ O} = 100 - (59.99\% \text{ C} + 4.48\% \text{ H}) = 35.53\% \text{ C}$$

5. Vitamin C is known chemically by the name ascorbic acid. Determine the empirical formula of ascorbic acid if it is composed of 40.92% carbon, 4.58% hydrogen, and 54.50% oxygen. (Answer: $C_3H_4O_3$)

$$\frac{40.92 \text{ g C}}{} \left| \frac{1 \text{ mol C}}{12.01 \text{ g C}} \right. \qquad \frac{4.58 \text{ g H}}{} \left| \frac{1 \text{ mol H}}{1.01 \text{ g H}} \right. \qquad \frac{54.50 \text{ g O}}{} \left| \frac{1 \text{ mol O}}{16.00 \text{ g O}} \right.$$

$= 3.41 \text{ mol C} \div 3.41 \rightarrow 1 \text{ mol C}$, $\times 3 = 3 \text{ mol C}$

$= 4.53 \text{ mol H} \div 3.41 \rightarrow 1.33 \text{ mol H}$, $\times 3 = 4 \text{ mol H}$

$= 3.41 \text{ mol O} \div 3.41 \rightarrow 1 \text{ mol O}$, $\times 3 = 3 \text{ mol O}$

$\therefore$ EF is $C_3H_4O_3$

6. What is the molecular formula of a compound given the molar mass of the compound is 30.04 grams/mole and the empirical formula is NH?

EFM

N = 14.01

H = 1.01

15.02

$$n = \frac{\text{NM}}{\text{EFM}} = \frac{30.04}{15.02} = 2 \quad \therefore \text{ MF is } (N_{1(\times 2)}H_{1(\times 2)}) = N_2H_2$$

Balancing Chemical Equations 1

While balancing equations, take into consideration the following:

1. **Never change the formula and the subscripts in a formula.**
2. Remember the seven **diatomic molecules are always written as H_2, N_2, O_2, F_2, Cl_2, Br_2, I_2. Balance the atoms** on either side of the equation by **changing the coefficients**.
3. Make sure the **coefficients are the smallest set of a whole number.**

Now try balancing these equations:

1. $\mathbf{2}\,Cu + \mathbf{1}\,O_2 \rightarrow \mathbf{2}\,CuO$

2. $\mathbf{1}\,S_8 + \mathbf{24}\,F_2 \rightarrow \mathbf{8}\,SF_6$

3. $\mathbf{1}\,P_4O_{10} + \mathbf{6}\,H_2O \rightarrow \mathbf{4}\,H_3PO_4$

4. $\mathbf{1}\,ZnS + \mathbf{2}\,HCl \rightarrow \mathbf{1}\,ZnCl_2 + \mathbf{1}\,H_2S$

5. $\mathbf{1}\,C_5H_{12} + \mathbf{8}\,O_2 \rightarrow \mathbf{5}\,CO_2 + \mathbf{6}\,H_2O$

6. $\mathbf{1}\,C_2H_5OH + \mathbf{3}\,O_2 \rightarrow \mathbf{2}\,CO_2 + \mathbf{3}\,H_2O$

7. $\mathbf{2}\,NH_3 + \mathbf{1}\,O_2 \rightarrow \mathbf{2}\,NO + \mathbf{3}\,H_2$

8. $\mathbf{2}\,Fe_2S_3 + \mathbf{9}\,O_2 \rightarrow \mathbf{2}\,Fe_2O_3 + \mathbf{6}\,SO_2$

9. $\mathbf{2}\,C_6H_6 + \mathbf{15}\,O_2 \rightarrow \mathbf{12}\,CO_2 + \mathbf{6}\,H_2O$

10. _______ $NaN_3 \rightarrow$ _______ Na + _______ N_2

Balancing Chemical Equations 2

1. **1** N_2 + **3** H_2 → **2** NH_3

2. **2** P_2O_3 → **1** P_4 + **3** O_2

3. **2** $AgNO_3$ + ____ Cu → ____ $Cu(NO_3)_2$ + **2** Ag

4. **1** CF_4 + **2** Br_2 → **1** CBr_4 + **2** F_2

5. **1** CH_4 + **2** O_2 → **1** CO_2 + **2** H_2O

6. **1** MgF_2 + **1** Li_2CO_3 → **1** $MgCO_3$ + **2** LiF

7. **2** $RbNO_3$ + **1** BeF_2 → ____ $Be(NO_3)_2$ + **2** RbF

8. **2** HCN + **1** $CuSO_4$ → **1** H_2SO_4 + **1** $Cu(CN)_2$

9. **1** GaF_3 + **3** Cs → **3** CsF + **1** Ga

10. **1** BaS + **1** PtF_2 → **1** BaF_2 + **1** PtS

11. **2** NaF + **1** Br_2 → **2** NaBr + **1** F_2

12. **2** $AlBr_3$ + **3** K_2SO_4 → **6** KBr + **1** $Al_2(SO_4)_3$

13. **2** Al + **6** HCl → **3** H_2 + **2** $AlCl_3$

14. **1** $Pb(OH)_2$ + **2** HCl → **2** H_2O + **1** $PbCl_2$

15. **2** NH_3 + **1** H_2SO_4 → **1** $(NH_4)_2SO_4$

Balancing Word Equations

1. Solid lithium carbonate reacts with aqueous hydrochloric acid (HCl) to produce liquid water, carbon dioxide gas, and aqueous lithium chloride.

   $Li_2CO_3 + 2\ HCl \rightarrow 2\ LiCl + H_2O + CO_2$

2. Aqueous sodium sulfide will react with aqueous magnesium iodide to produce aqueous sodium iodide and a precipitate of magnesium sulfide.

   $Na_2S + MgI_2 \rightarrow 2\ NaI + MgS$

3. Liquid ethanol (CH_3CH_2OH or C_2H_5OH) will burn in oxygen gas to produce liquid water and carbon dioxide gas.

   $C_2H_5OH + 3\ O_2 \rightarrow 3\ H_2O + 2\ CO_2$

4. Hydrogen gas and oxygen gas can be burned, under certain conditions, to form liquid hydrogen peroxide (officially called dihydrogen dioxide).

   $H_2 + O_2 \rightarrow H_2O_2$

5. Solid calcium chloride can react with liquid water to produce hydrogen chloride gas and aqueous calcium oxide.

   $CaCl_2 + H_2O \rightarrow 2\ HCl + CaO$

6. Aqueous aluminum nitrate can react with aqueous potassium fluoride to produce aqueous potassium nitrate and a precipitate of aluminum fluoride.

   $Al(NO_3)_3 + 3\ KF \rightarrow 3\ KNO_3 + AlF_3$

7. Solid silicon and oxygen gas can react to form a solid disiliconhexaoxide.

   $2\ Si + 3\ O_2 \rightarrow Si_2O_6$

8. Aqueous ammonium chloride will react with aqueous calcium hydroxide to form aqueous calcium chloride and aqueous ammonium hydroxide.

   $2\ NH_4Cl + Ca(OH)_2 \rightarrow CaCl_2 + 2\ NH_4OH$

9. Aqueous iron (II) bromide can react with liquid water to produce aqueous iron (II) oxide and hydrogen bromide gas.

   $FeBr_2 + H_2O \rightarrow FeO + 2\ HBr$

10. Silver metal can react with hydrochloric acid (HCl) to form a precipitate of silver chloride and hydrogen gas.

    $2\ Ag + 2HCl \rightarrow 2\ AgCl + H_2$

Tutorial: Classifying Types of Reactions

Synthesis/combination	$A + B \rightarrow AB$
Decomposition	$AB \rightarrow A + B$
Single-replacement	$A + BC \rightarrow AC + B$
Double-replacement	$AB + CD \rightarrow AD + CB$
Combustion	$CxHy + O_2 \rightarrow CO_2 + H_2O$

1. Balance these reactions:

   a. $\mathbf{2}\,Na\,(s) \quad + \quad Cl_2\,(g) \quad \rightarrow \quad \mathbf{2}\,NaCl\,(s)$

   b. $\mathbf{2}\,HgO\,(s) \quad \rightarrow \quad \mathbf{2}\,Hg\,(l) \quad + \quad O_2\,(g)$

   c. $AgNO_3\,(aq) \quad + \quad NaCl\,(aq) \quad \rightarrow \quad AgCl\,(s) \quad + \quad NaNO_3\,(aq)$

   d. $CaO\,(s) \quad + \quad CO_2\,(g) \quad \rightarrow \quad CaCO_3\,(s)$

   e. $HCl\,(aq) \quad + \quad NaOH\,(aq) \quad \rightarrow \quad H_2O\,(l) \quad + \quad NaCl\,(aq)$

   f. $\mathbf{2}\,C_2H_6\,(g) \quad + \quad \mathbf{7}\,O_2\,(g) \quad \rightarrow \quad \mathbf{4}\,CO_2\,(g) \quad + \quad \mathbf{6}\,H_2O\,(g)$

   g. $\mathbf{2}\,HCl\,(aq) \quad + \quad CaS\,(aq) \quad \rightarrow \quad H_2S\,(g) \quad + \quad CaCl_2\,(aq)$

   h. $\mathbf{2}\,Ca\,(s) \quad + \quad O_2\,(g) \quad \rightarrow \quad \mathbf{2}\,CaO\,(s)$

   i. $\mathbf{2}\,H_2O\,(l) \quad \rightarrow \quad \mathbf{2}\,H_2\,(g) \quad + \quad O_2\,(g)$

   j. $Zn\,(s) \quad + \quad CuCl_2\,(aq) \quad \rightarrow \quad ZnCl_2\,(aq) \quad + \quad Cu\,(s)$

2. Which of the equations above describe each type of reaction? List all that apply.

   a. Synthesis **1, 4, 8**

   b. Single-replacement **10**

   c. Decomposition **2, 9**

   d. Double-replacement **3, 5, 7**

   e. Combustion **6, 8**

Stoichiometry Practice

1. Balance the following chemical equations.

   a. $Zn + \mathbf{2}\ HCl \rightarrow ZnCl_2 + H_2$

   b. $\mathbf{3}\ CO + Fe_2O_3 \rightarrow \mathbf{2}\ Fe + \mathbf{3}\ CO_2$

   c. $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + \mathbf{2}\ NaCl$

   d. $\mathbf{2}\ Al + \mathbf{3}\ CuSO_4 \rightarrow \mathbf{3}\ Cu + Al_2(SO_4)_3$

   e. Which of the reactions above use BrINClHOF for diatomic molecules?

   a / b / c / d

2. Classify each of the following balanced equations as combination, decomposition, single or double displacement, combustion (may be more than one type for each).

   a. $N_2 + 3H_2 \rightarrow 2\ NH_3$
   Synthesis/combination

   b. $BaCl_2 + K_2CO_3 \rightarrow BaCO_3 + 2\ KCl$
   Double displacement

   c. $H_2CO_3 \rightarrow H_2O + CO_2$
   Decomposition

   d. $2\ K + Br_2 \rightarrow 2\ KBr$
   Synthesis/combination

   e. $N_2 + 2\ O_2 \rightarrow 2\ NO_2$
   Combination/synthesis

   f. $Cu + 2\ AgNO_3 \rightarrow 2\ Ag + Cu(NO_3)_2$
   Single displacement

   g. Which of the reactions above include BrINClHOF for diatomic molecules?

   a / b / c / d / e / f

3. Complete and balance the combustion reaction for each of the following fuels:

   a. Methane (natural gas, formula CH_4)

   $\mathbf{CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)}$

   b. Isooctane (major component of gasoline, formula C_8H_{18})

   $\mathbf{2\ C_8H_{18}\ (l) + 25\ O_2\ (g) \rightarrow 16\ CO_2\ (g) + 18\ H_2O\ (g)}$

   c. Ethanol (alternative fuel and alcoholic beverage component, formula C_2H_5OH)

   $\mathbf{C_2H_5OH\ (l) + 3\ O_2\ (g) \rightarrow 2\ CO_2\ (g) + 3\ H_2O\ (g)}$

Stoichiometry 1

1. Ammonia gas reacts with oxygen gas according to the following equation:

MM 17.03 g/mol 32.00 g/mol 30.01 g/mol 18.02 g/mol

$$4\ NH_3 + 5\ O_2 \longrightarrow 4\ NO + 6\ H_2O$$

a. How many moles of oxygen gas are needed to react with 23 moles of ammonia? (29 mole)

$$? \text{ mol } O_2 = \frac{23 \text{ mol } NH_3 \mid 5 \text{ mol } O_2}{\mid 4 \text{ mol } NH_3} = 28.75 \text{ mol } O_2 = 29 \text{ mol } O_2$$

b. How many grams of NO are produced when 25 moles of oxygen gas react with an excess of ammonia? (6.0×10^2 g)

$$? \text{ g NO} = \frac{25 \text{ mol } O_2 \mid 4 \text{ mol NO} \mid 30.01 \text{ g NO}}{\mid 5 \text{ mol } O_2 \mid 1 \text{ mol NO}} = 6.0 \times 10^2 \text{ g NO}$$

c. If 24 grams of water are produced, how many moles of nitrogen monoxide are formed? (0.89 mole)

$$? \text{ mol NO} = \frac{24 \text{ g } H_2O \mid 1 \text{ mol } H_2O \mid 4 \text{ mol NO}}{\mid 18.02 \text{ g } H_2O \mid 6 \text{ mol } H_2O} = 0.8879 \text{ mol NO} \approx 0.89 \text{ mol NO}$$

d. How many grams of oxygen are needed to react with 6.78 grams of ammonia? (16.0 g)

$$? \text{ g } O_2 = \frac{6.78 \text{ g } NH_3 \mid 1 \text{ mol } NH_3 \mid 5 \text{ mol } O_2 \mid 32.00 \text{ g } O_2}{\mid 17.03 \text{ g } NH_3 \mid 4 \text{ mol } NH_3 \mid 1 \text{ mol } O_2} = 15.9 \text{ g } O_2$$

2. The compound calcium carbide, CaC_2, is made by reacting calcium carbonate with carbon at high temperatures. The UNBALANCED EQUATION for the reaction is:

MM 100.09 g/mol 12.01 g/mol 64.1 g/mol 44.01 g/mol

$$2\ CaCO_3 + 5\ C \longrightarrow 2\ CaC_2 + 3\ CO_2$$

a. Balance the equation.

b. How many moles of carbon are required to produce 5.0 moles CO_2? (8.3 mole)

$$?\ \text{mol C} = \frac{\text{given } 5.0\ \text{mol } CO_2}{} \left| \frac{5\ \text{mol C}}{3\ \text{mol } CO_2} \right. = 8.3\ \text{mol C}$$

c. How many grams of calcium carbide are produced when 4.00 moles of carbon react with an excess of calcium carbonate? (102 g)

$$?\ \text{g } CaC_2 = \frac{\text{given } 4.0\ \text{mol C}}{} \left| \frac{2\ \text{mol } CaC_2}{5\ \text{mol C}} \right| \frac{64.1\ \text{g } CaC_2}{1\ \text{mol } CaCl_2} = 102.56\ CaC_2 = 1.0 \times 10^2\ \text{g } CaC_2$$

d. How many moles of carbon dioxide are produced when 55 grams of calcium carbonate react with an excess of carbon? (0.83 mole)

$$?\ \text{mol } CO_2 = \frac{\text{given } 55\ \text{g } CaCO_3}{} \left| \frac{1\ \text{mol } CaCO_3}{100.09\ \text{g } CaCO_3} \right| \frac{3\ \text{mol } CO_2}{2\ \text{mol } CaCO_3} = 0.82\ \text{mol } CO_2$$

e. How many grams of carbon are needed to react with 453 grams of calcium carbonate? (136 g)

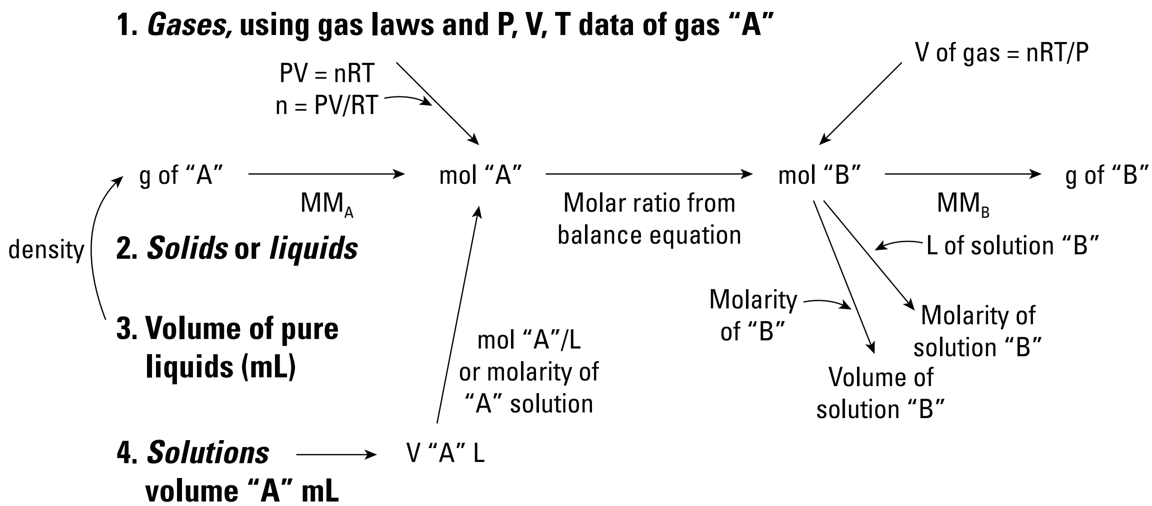
$$?\ \text{g C} = \frac{453\ \text{g } CaCO_3}{} \left| \frac{1\ \text{mol } CaCO_3}{100.09\ \text{g } CaCO_3} \right| \frac{5\ \text{mol C}}{2\ \text{mol } CaCO_3} \left| \frac{12.01\ \text{g C}}{1\ \text{mol C}} \right. = 136\ \text{g C}$$

f. How many grams of calcium carbonate are needed to form 598 grams of calcium carbide? (934 g)

$$?\ \text{g } CaCO_3 = \frac{598\ \text{g } CaC_2}{} \left| \frac{1\ \text{mol } CaC_2}{64.1\ \text{g } CaC_2} \right| \frac{2\ \text{mol } CaCO_3}{2\ \text{mol } CaC_2} \left| \frac{100.09\ \text{g } CaCO_3}{1\ \text{mol } CaCO_3} \right. = 934\ \text{g } CaCO_3$$

Stoichiometry Calculations Using Balanced Equations

A TOOL FOR STOICHIOMETRY



ABBREVIATIONS USED

MM = Molar mass

V = volume

P = pressure

T = temperature

n = moles

NOTE: Gases may not be covered in CHEM& 140 but will be covered in CHEM& 162. Molarity will be covered later in this course.

Stoichiometry: Iron Ore Smelting

Iron ore is converted to iron by heating it with coal and oxygen in a process called smelting:

$$2\ Fe_2O_3\ (s) + 4\ C\ (s) + 1\ O_2\ (g) \longrightarrow 4\ Fe\ (s) + 4\ CO_2\ (g)$$

unbalanced equation!

1. Calculate the number of moles of Fe_2O_3 that would be needed to react exactly with 0.49 mol of O_2. (Answer: 0.98 mol)

$$?\ \text{mol } Fe_2O_3 = \frac{0.49\ \text{mol } O_2}{} \times \frac{2\ \text{mol } Fe_2O_3}{1\ \text{mol } O_2} = 0.98\ \text{mol } Fe_2O_3$$

2. Calculate the number of moles of C that would need to react to produce 6429 g of Fe. (Answer: 115.1 mol)

$$?\ \text{mol C} = \frac{6429\ \text{g Fe}}{} \times \frac{1\ \text{mol Fe}}{55.85\ \text{g Fe}} \times \frac{4\ \text{mol C}}{4\ \text{mol Fe}} = 115.1\ \text{mol C}$$

3. How many grams of CO_2 would be produced from smelting 2385 g of Fe_2O_3? (Answer: 1315 g)

$$?\ \text{g } CO_2 = \frac{2385\ \text{g } Fe_2O_3}{} \times \frac{1\ \text{mol } Fe_2O_3}{159.7\ \text{g } Fe_2O_3} \times \frac{4\ \text{mol } CO_2}{2\ \text{mol } Fe_2O_3} \times \frac{44.01\ \text{g } CO_2}{1\ \text{mol } CO_2} = 1315\ \text{g } CO_2$$

MM Fe_2O_3

$2 \times 55.85 = 111.7$

$3 \times 16.00 = 48.0$

159.7 g/mol

4. How many kilograms of Fe can be smelted from 7.545 kg of Fe_2O_3? (Answer: 5.227 kg)

$$?\ \text{kg Fe} = \frac{7.545\ \text{kg } Fe_2O_3}{} \times \frac{10^3\ \text{g}}{1\ \text{kg}} \times \frac{1\ \text{mol } Fe_2O_3}{159.7\ Fe_2O_3} \times \frac{4\ \text{mol Fe}}{2\ \text{mol } Fe_2O_3} \times \frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}} \times \frac{1\ \text{kg}}{10^3\ \text{g}} = 5.277\ \text{kg Fe}$$

Stoichiometry: Removing CO_2 in Space

The Contaminate Control Cartridge (CCC), which contains lithium hydroxide (LiOH), is used to remove carbon dioxide (CO_2) from the space shuttle cabin.

The removal of CO_2 is represented by the following equation:

$$\mathbf{2\ LiOH}\ (s) + \mathbf{CO_2}\ (g) \longrightarrow \mathbf{Li_2CO_3}\ (s) + \mathbf{H_2O}\ (g)$$

A typical space shuttle crew consists of six individuals and each CCC contains 750 g of LiOH. Assuming that each crew member expels 42.0 g of CO_2 per hour on average, and that a mission is scheduled to last 18 days, how many CCCs must be carried on board the space shuttle?

PROBLEM AND SOLUTION KEY (ONE APPROACH)

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$$2\ \text{LiOH}\ (s) + CO_2\ (g) \longrightarrow Li_2CO_3\ (s) + H_2O\ (g)$$

1. A typical space shuttle crew consists of six individuals and each CCC contains 750 g of LiOH. Assuming that each crew member expels 42.0 g of CO_2 per hour on average, and that a mission is scheduled to last 18 days, how many CCCs must be carried on board the space shuttle?

$$18\ \text{days} \cdot \frac{24\ \text{hr}}{\text{day}} \cdot \frac{42.0\ \text{g}}{1\ \text{hr}} = 18{,}100\ \text{g per person}$$

For a crew of six:

$6(18{,}100\ \text{g}) = 109{,}000\ \text{g}\ CO_2$ per person

$$108{,}000\ \text{g}\ CO_2 \cdot \frac{1\ \text{mol}\ CO^2}{44.0\ \text{g}\ CO_2} \cdot \frac{2\ \text{mol LiOH}}{1\ \text{mol CO2}} \cdot \frac{23.9\ \text{g LiOH}}{1\ \text{mol LiOH}} \cdot \frac{1\ \text{CCC}}{750\ \text{g LiOH}} = 158\ \text{CCCs}$$

From Nasa.gov "Carbon dioxide removal – Stoichiometry"

Limiting Reactant by Moles

1. A recipe for cookies states that 3 cups of flour, 2 eggs, and 1.5 tsp of baking soda can make 24 cookies. Let's say that you have 5 cups of flour, 3 eggs, and 4 tsp of baking soda.

   a. How many cookies can be made (what is the theoretical yield)?

   3 cup F + 2 E + 1.5 BS → 24 cookies

   Using flour $? \text{ cookies} = \frac{5 \text{ cups F} \mid 24 \text{ cookies}}{\mid 3 \text{ cups F}} = 40 \text{ cookies}$

   Using eggs $? \text{ cookies} = \frac{3 \text{ eggs} \mid 24 \text{ cookies}}{\mid 2 \text{ eggs}} = 36 \text{ cookies}$ can be made with the supplies on hand. This is also the Theoretical yield.

   Using BS $? \text{ cookies} = \frac{4 \text{ tsp BS} \mid 24 \text{ cookies}}{\mid 1.5 \text{ BS}} = 64 \text{ cookies}$

   b. What is the limiting reactant? How can you tell?

   Least amount of cookies (product) made by using the eggs available on hand.

   "Egg" is the limiting reactant as it limits the amount of cookies that can be made.

2. Ammonia gas reacts with oxygen gas according to the following equation:

$$4\ NH_3 + 5\ O_2 \longrightarrow 4\ NO + 6\ H_2O$$

   a. Calculate the number of moles of oxygen gas needed to react with 7.00 moles of NH_3.

   $? \text{ mol } O_2 = \frac{7.00 \text{ mol } NH_3 \mid 5 \text{ mol } O_2}{\mid 4 \text{ mol } NH_3} = 8.75 \text{ mol } O_2 \text{ needed}$

   b. If you had 7.00 moles of NH_3 and 6.00 moles of oxygen gas, what would be the limiting reactant?

   7.00 mol NH_3 need 8.75 mol O_2

   You have 6.00 mol O_2, or see calculations to make H_2O using NH_3.

   Using NH_3 = $\frac{7.00 \text{ mol } NH_3 \mid 6 \text{ mol } H_2O}{\mid 4 \text{ mol } NH_3}$ = Not made

   Gives least number of moles H_2O, so O_2 is LR.

   Using O_2 = $\frac{6.00 \text{ mol } NH_3 \mid 6 \text{ mol } H_2O}{\mid 4 \text{ mol } NH_3}$ = 9 mol H_2O

   O_2 is the limiting reactant.

   c. If you had 7.00 moles of NH_3 and 10.0 moles of oxygen gas, what would be the limiting reactant?

   7.00 mol NH_3 need 8.75 mol O_2

   You have 7.00 mol NH_3, you have 10.00 mol O_2 → Excess is O_2 now

   LR would be NH_3

3. Aluminum reacts with aqueous sulfuric acid to produce an aluminum sulfate aqueous solution and hydrogen gas.

$$2\ Al\ (s) + 3\ H_2SO_4\ (aq) \rightarrow Al_2(SO_4)_3\ (aq) + 3\ H_2\ (g)$$

a. What would be the limiting reactant if you had 2 moles of Al and 2 moles of sulfuric acid?

2 mol Al needs 3 mol H_2SO_4.
You have 2 mol Al and 2 mol H_2SO_4. ← Less of this means H_2SO_4 is LR per 2 mol Al

$$\frac{2\text{ mol Al} \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 2\text{ mol Al}} = 1\text{ mol } Al_2(SO_4)_3 \qquad \frac{2\text{ mol } H_2SO_4 \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 3\text{ mol } H_2SO_4} = 0.66\text{ mol } Al_2(SO_4)_3$$

b. What would be the limiting reactant if you had 3 moles of Al and 3 moles of sulfuric acid?

2 mol Al needs 3 mol H_2SO_4.
You have 3 mol Al (excess) and 3 mol H_2SO_4 (LR).

$$\frac{3\text{ mol Al} \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 2\text{ mol Al}} = 1.5\text{ mol } Al_2(SO_4)_3 \qquad \frac{3\text{ mol } H_2SO_4 \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 3\text{ mol } H_2SO_4} = 1\text{ mol } Al_2(SO_4)_3$$

c. What would be the limiting reactant if you had 5 moles of Al and 7 moles of sulfuric acid?

$$?\text{ mol } Al_2(SO_4)_3 \text{ made using Al} = \frac{5\text{ mol Al} \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 2\text{ mol Al}} = 2.8\text{ mol } Al_2(SO_4)_3 \text{ (Not made)}$$

$$?\text{ mol } Al_2(SO_4)_3 \text{ made using } H_2SO_4 = \frac{7\text{ mol } H_2SO_4 \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 3\text{ mol } H_2SO_4} = 2.33\text{ mol } Al_2(SO_4)_3$$

d. What would be the limiting reactant if you had 0.50 moles of Al and 0.75 moles of sulfuric acid?

$$\text{Using Al} = \frac{0.5\text{ mol Al} \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 2\text{ mol Al}} = 0.25\text{ mol } Al_2(SO_4)_3$$

$$\text{Using } H_2SO_4 = \frac{0.75\text{ mol } H_2SO_4 \mid 1\text{ mol } Al_2(SO_4)_3}{\mid 3\text{ mol } H_2SO_4} = 0.25\text{ mol } Al_2(SO_4)_3$$

No LR here

4. Ammonia can be synthesized by the following reaction.

$$2\ NO\ (g) + 5\ H_2\ (g) \rightarrow 2\ NH_3\ (g) + 2\ H_2O\ (g)$$

a. If 10.0 moles of NO and 20.0 moles of hydrogen react, what is the theoretical yield (in moles) of ammonia? (Answer: 8 mol)

$$?\text{ mol } NH_3 \text{ made using NO} = \frac{10\text{ mol NO} \mid 2\text{ mol } NH_3}{\mid 2\text{ mol NO}} = 10\text{ mol } NH_3 \text{ not produced}$$

$$?\text{ mol } NH_3 \text{ made using } H_2 = \frac{20\text{ mol } H_2 \mid 2\text{ mol } NH_3}{\mid 5\text{ mol } H_2} = 8\text{ mol } NH_3 \text{ produced (Theoretical yield)}$$

b. If 0.500 moles of NO and 1.50 moles of hydrogen react, what is the theoretical yield (in moles) of ammonia? (Answer: 0.5 mol)

$$?\text{ mol } NH_3 \text{ made using NO} = \frac{0.5\text{ mol NO} \mid 2\text{ mol } NH_3}{\mid 2\text{ mol NO}} = 0.5\text{ mol } NH_3 \text{ (Theoretical yield)}$$

$$?\text{ mol } NH_3 \text{ made using } H_2 = \frac{1.5\text{ mol } H_2 \mid 2\text{ mol } NH_3}{\mid 5\text{ mol } H_2} = 0.6\text{ mol } NH_3 \text{ not produced}$$

Activity: Limiting Reactant

PART 1: REVIEW STOICHIOMETRY

There are three steps to doing gram-to-gram conversions in stoichiometry:

1. Convert grams of the starting substance to moles. (Use molar mass!)
2. Convert the moles of the starting substance to moles of the ending substance. (Use the mole ratio given in the balanced chemical equation.)
3. Convert the moles of the ending substance to grams. (Use molar mass!)

USE SIGNIFICANT FIGURES! USE UNITS! SHOW YOUR WORK!

REVIEW

1. Balance this chemical equation of the combustion of methane:

MM	16.05 g/mol		32.00 g/mol		44.01 g/mol		18.02 g/mol
	1 CH_4	+	**2** O_2	$\rightarrow$	**1** CO_2	+	**2** H_2O

2. What is the theoretical yield or the maximum mass of water (in grams) that may be formed from the reaction of 5.00 grams of methane, CH_4? Follow the three steps above and show your work below. (Answer: 11.2 g)

   (**NOTE:** In this problem, oxygen is "in excess." When there is an excess of one reactant that means you can assume all of the other reactant gets converted to product. (In other words, you don't have to worry about it in the calculation.) We can say "methane reacts completely" or "methane is the limiting reactant" since the calculation is based on this assumption.)

   Solution map: –g $CH_4 \xrightarrow{1}$ mol $CH_4 \xrightarrow{2}$ mol $H_2O \xrightarrow{3}$ g H_2O

   $$? \text{ g } H_2O = \frac{5.00\text{ g } CH_4}{} \left|\frac{1\text{ mol } CH_4}{16.05\text{ g } CH_4}\right|\frac{2\text{ mol } H_2O}{1\text{ mol } CH_4}\left|\frac{18.02\text{ g } H_2O}{1\text{ mol } H_2O}\right. = 11.2\text{ g } H_2O$$

   Theoretical yield of water: **11.2** grams **H_2O**

3. Now suppose there is an excess of methane and only 5.00 grams of O_2. (Oxygen is now the limiting reactant.) What is the theoretical yield of water in this case? Follow the three steps above and show your work below. (Answer: 2.82 g)

   g $O_2 \xrightarrow{1}$ mol $O_2 \xrightarrow{2}$ mol $H_2O \xrightarrow{3}$ g H_2O

   $$? \text{ g } H_2O = \frac{5.00\text{ g } O_2}{} \left|\frac{1\text{ mol } O_2}{32.00\text{ g } O_2}\right|\frac{2\text{ mol } H_2O}{2\text{ mol } O_2}\left|\frac{18.02\text{ g } H_2O}{1\text{ mol } H_2O}\right. = 2.82\text{ g } H_2O$$

   Theoretical yield of water: **2.82** grams **H_2O**

PART 2: THE LIMITING REACTANT

The Limiting Reactant is the reactant that determines the theoretical yield. The Limiting Reactant is completely consumed (or "used up") while some of the other reactant (in excess) will be "left-over" at the end of the reaction.

1. Suppose you have a reaction with 5.00 grams of CH_4 **AND** 5.00 grams of O_2 in one container. Unlike in the Review Questions #2 and #3, you are not told which reactant is limiting and which is in excess so you are not sure which to use in the calculation, 5.00 grams of methane OR 5.00 grams of oxygen.

   In this situation, **YOU** have to determine which is the limiting reactant. You do this by setting up two stoichiometry problems—one for each reactant. Luckily, you already did this on the previous page. Let's compare the amount of water formed from 5.00 grams of each reactant:

   Amount (grams) of water formed from reacting 5.00 grams of methane: **11.2** grams **H_2O**

   Amount (grams) of water formed from reacting 5.00 grams of oxygen: **2.82** grams **H_2O**

2. Which one "limits" the amount of water formed, the 5.00 grams of CH_4 or 5.00 grams of O_2? (Circle one.) The one you circle is called "the limiting reactant."

   How many grams of water is formed from this limiting reactant? **2.82** grams **H_2O**

   How many grams of the limiting reactant should remain after reaction? **zero** grams
   (LR used up completely.)

3. To summarize, fill in the blanks:

   When 5.00 grams of CH_4 and 5.00 grams of O_2 react, **2.82** grams of water can be formed.

   In this case, **O_2** is the limiting reactant and **CH_4** is the excess reactant.

4. Which substances should be present after the reaction? Circle all that apply:

   ER **LR** **Products**

   CH_4 / ~~O_2~~ / H_2O / CO_2

The theoretical yield is based on calculations assuming all of the limiting reactant is converted to product. In reality, this may not be the case. The amount of product actually formed in a reaction is called the "actual yield" and is usually less than the amount expected in theory.

$$\textbf{\% yield} \ \frac{\textit{actual yield}}{\textit{theoretical yield}} \times 100$$

5. If this reaction of methane and oxygen had a 75% yield, what is the **actual** mass of water produced?

75% of 2.82 = $\frac{2.82 \;|\; 75}{\;|\; 100}$ **= 2.1 g actual yield**

Or we use 75% = $\frac{AY \;|\; 100}{2.82\text{ g} \;|\;}$ **; ∴ AY =** $\frac{75 \times 2.82}{100}$ **= 2.1 g H_2O formed**

Theoretical Yield and Limiting Reagents

1. For the reaction $3\ H_2\ (g) + N_2\ (g) \longrightarrow 2\ NH_3\ (g)$, 3 mol H_2 is reacted with 6 mol N_2.

   a. **2** mol of NH_3 is produced

   b. **0** mol of H_2 remains

   c. **5** mol N_2 remains

$$\frac{3\ \text{mol}\ H_2}{} \left| \frac{2\ \text{mol}\ NH_3}{3\ \text{mol}\ H_2} \right. = 2\ \text{mol}\ NH_3 \quad \text{(Theoretical yield)}$$

H_2 is LR

$$\frac{6\ \text{mol}\ N_2}{} \left| \frac{2\ \text{mol}\ NH_3}{1\ \text{mol}\ N_2} \right. = 12\ \text{mol}\ NH_3$$

$$\frac{3\ \text{mol}\ H_2}{} \left| \frac{1\ \text{mol}\ N_2}{3\ \text{mol}\ H_2} \right. = 1\ \text{mol}\ N_2\ \text{used}$$

Given 6 mol N_2—1 mol used = 5 mol N_2 remains

2. For the reaction $2\ N_2H_4\ (l) + N_2O_4\ (l) \longrightarrow 3\ N_2\ (g) + 4\ H_2O\ (l)$, 160 g N_2H_4 is mixed with 160 g N_2O_4.
(Answer: 125 g H_2O)

   a. **N_2O_4** is the limiting reagent

   b. **125** g H_2O is produced

Using N_2H_4: $$\frac{160\ \text{g}\ N_2H_4}{} \left| \frac{1\ \text{mol}\ N_2H_4}{32.05\ \text{g}\ N_2H_4} \right| \frac{4\ \text{mol}\ H_2O}{2\ \text{mol}\ N_2H_4} = 9.98\ \text{mol}\ H_2O$$

Using N_2O_4: $$\frac{160\ \text{g}\ N_2O_4}{} \left| \frac{1\ \text{mol}\ N_2O_4}{92.02\ \text{g}\ N_2O_4} \right| \frac{4\ \text{mol}\ H_2O}{1\ \text{mol}\ N_2O_4} = 6.96\ \text{mol}\ H_2O\ \text{formed}$$

MM N_2H_4	MM N_2O_4
28.02	28.02
4.032	64.00
32.052	92.02

$$\frac{6.96\ \text{mol}\ H_2O}{} \left| \frac{18.02\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O} \right. = 125\ \text{g}\ H_2O$$

3. For the reaction $Fe_2O_3\ (s) + 3\ CO\ (g) \longrightarrow 2\ Fe\ (g) + 3\ CO_2$, 224 g of CO is available to react with 400. g Fe_2O_3.
(Answer: 278 g Fe, 331 g CO_2)

   a. **Fe_2O_3** is the limiting reagent

   b. **278** g iron is produced

   c. **331** g CO_2 is produced

$$\frac{224\ \text{g CO}}{} \left| \frac{1\ \text{mol CO}}{28.01\ \text{g CO}} \right| \frac{2\ \text{mol Fe}}{3\ \text{mol CO}} \left| \frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}} \right. = 298\ \text{g Fe}$$

$$\frac{400\ \text{g}\ Fe_2O_3}{} \left| \frac{1\ \text{mol}\ Fe_2O_3}{159.7\ \text{g}\ Fe_2O_3} \right| \frac{2\ \text{mol Fe}}{1\ \text{mol}\ Fe_2O_3} \left| \frac{55.85\ \text{g Fe}}{1\ \text{mol Fe}} \right. = 278\ \text{g Fe formed}$$

$$\frac{400\ \text{g}\ Fe_2O_3}{} \left| \frac{1\ \text{mol}\ Fe_2O_3}{159.7\ \text{g}\ Fe_2O_3} \right| \frac{3\ \text{mol}\ CO_2}{1\ \text{mol}\ Fe_2O_3} \left| \frac{44.01\ \text{g}\ CO_2}{1\ \text{mol}\ CO_2} \right. = 331\ \text{g}\ CO_2\ \text{formed}$$

MM CO	MM Fe_2O_3	MM CO_2
12.01	2 × 55.85 = 111.7	12.01
16.00	3 × 16.00 = 48.0	32.00
28.01 g/mol	159.7 g/mol	44.01 g/mol

4. For the reaction $2\ C_4H_{10}\ (g) + 13\ O_2\ (g) \longrightarrow 8\ CO_2\ (g) + 10\ H_2O\ (l)$, 300 g of C_4H_{10} is combusted in 1000 g of O_2.
(Answer: 400 g H_2O)

   a. **O_2** is the limiting reagent

   b. **400 (1 s.f. used)** g H_2O is formed

$$\frac{300\ \text{g}\ C_4H_{10}}{} \left| \frac{1\ \text{mol}\ C_4H_{10}}{58.12\ \text{g}\ C_4H_{10}} \right| \frac{10\ \text{mol}\ H_2O}{2\ \text{mol}\ C_4H_{10}} = 25.8\ \text{mol}\ H_2O\ \text{not formed}$$

$$\frac{1000\ \text{g}\ O_2}{} \left| \frac{1\ \text{mol}\ O_2}{32.00\ \text{g}\ O_2} \right| \frac{10\ \text{mol}\ H_2O}{13\ \text{mol}\ O_2} = 24.0\ \text{mol}\ H_2O\ \text{formed}$$

$$\frac{24.0\ \text{mol}\ H_2O}{} \left| \frac{18.02\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O} \right. = 433 \approx 400\ \text{g}$$

MM C_4H_{10}
4 × 12.01 = 48.04
10 × 1.008 = 10.08
58.12 g/mol

From ThoughtCo. Chemistry www.thoughtco.com

Percent Yield

1. For the reaction of nitrogen and hydrogen to produce ammonia:

$$3\,H_2\,(g) + N_2\,(g) \longrightarrow 2\,NH_3\,(g)$$

   a. If 10.0 g of N_2 were reacted, what is the theoretical yield of ammonia, in grams? (Answer: 12.2 g)

   Actual yield

   b. If 10.0 g of N_2 were reacted, and 9.1 g of ammonia were obtained, what is the % yield? (Answer: 74.6%)

$$? \text{ g } NH_3 = \frac{10.0 \text{ g } N_2}{} \Big| \frac{1 \text{ mol } N_2}{28.02 \text{ g } N_2} \Big| \frac{2 \text{ mol } NH_3}{1 \text{ mol } N_2} \Big| \frac{17.03 \text{ g } NH_3}{1 \text{ mol } NH_3} = 12.2 \text{ g } NH_3 \text{ formed}$$

Theoretical yield

$$\% \text{ yield} = \frac{9.1 \text{ g } NH_3}{12.2 \text{ g } NH_3} \Big| \frac{100}{} = 74.6\% \text{ yield}$$

2. Potassium superoxide, KO_2, reacts with carbon dioxide to form potassium carbonate and oxygen. This reaction makes potassium superoxide useful in a self-contained breathing apparatus. Write a balanced equation before doing calculations.

   a. How much oxygen could be produced from 2.50 g of KO_2 and 4.50 g of carbon dioxide? (Answer: 0.845 g)

   MM KO_2
   39.10
   32.00
   71.10

   Actual yield

   b. If only 0.500 g of oxygen is produced from 2.50 g of KO_2 and 4.50 g of carbon dioxide, what is the % yield? (Answer: 59.2%)

Balanced equation: $4KO_2 + 2CO_2 \longrightarrow 2K_2CO_3 + 3O_2$

$$? \text{ g } O_2 \text{ using } KO_2 = \frac{2.50 \text{ g } KO_2}{} \Big| \frac{1 \text{ mol } KO_2}{71.10 \text{ g } KO_2} \Big| \frac{3 \text{ mol } O_2}{4 \text{ mol } KO_2} = 0.0264 \text{ mol } O_2; \text{ produced} \qquad \frac{0.0264 \text{ mol } O_2}{} \Big| \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} = 0.845 \text{ g } O_2 \text{ formed}$$

$$? \text{ g } O_2 \text{ using } CO_2 = \frac{4.50 \text{ g } CO_2}{} \Big| \frac{1 \text{ mol } CO_2}{44.01 \text{ g } CO_2} \Big| \frac{3 \text{ mol } O_2}{2 \text{ mol } CO_2} = 0.\cancel{1}53 \text{ mol } O_2$$

$$\% \text{ yield} = \frac{0.500 \text{ g } O_2}{0.845 \text{ g } O_2} \Big| \frac{100}{} = 59.2\% \text{ yield}$$

3. After 3.8 mL of liquid ethanol (C_2H_5OH, density = 0.789 g/mL) is allowed to burn in the presence of 12.5 g of oxygen, 3.10 mL of water (density = 1.00 g/mL) is collected. Determine the limiting reactant, theoretical yield of H_2O and percent yield for the reaction. Write the balanced chemical equation for the combustion of ethanol before doing calculations. (Answer: 3.52 g H_2O, 88.1%)

   MM C_2H_5OH
   2 C = 24.02
   6 H = 6.048
   1 O = 16
   46.068

$C_2H_5OH\,(l) + 3O_2\,(g) \longrightarrow 2CO_2\,(g) + 3H_2O\,(l)$

= 2.99 g; = 0.0651 mol; = 0.195 mol H_2O formed

$$? \text{ g } H_2O \text{ using } C_2H_5OH = \frac{3.8 \text{ mL } C_2H_5OH}{} \Big| \frac{0.789 \text{ g } C_2H_5OH}{1 \text{ mL } C_2H_5OH} \Big| \frac{1 \text{ mol } C_2H_5OH}{46.07 \text{ g } C_2H_5OH} \Big| \frac{3 \text{ mol } H_2O}{1 \text{ mol } C_2H_5OH} \Big| \frac{18.02 \text{ g } H_2O}{1 \text{ mol } H_2O} = 3.52 \text{ g } H_2O$$

Limiting reactant

Theoretical yield

$$? \text{ g } H_2O \text{ using } O_2 = \frac{12.5 \text{ g } O_2}{} \Big| \frac{1 \text{ mol } O_2}{32.00 \text{ g } O_2} \Big| \frac{3 \text{ mol } H_2O}{3 \text{ mol } O_2} = 0.391 \text{ mol } H_2O \text{ not formed}$$

ER

$$\text{Actual yield} = \frac{3.10 \text{ mL}}{} \Big| \frac{1.00 \text{ g}}{1 \text{ mL}} = 3.10 \text{ g } H_2O; \quad \% \text{ yield} = \frac{3.10}{3.52} \Big| \frac{100}{} = 88.1\% \text{ yield}$$

Review: Stoichiometry

1. For this generic chemical reaction: 2 A + 3 B $\longrightarrow$ 3 C

   How many moles of B are required to react with: (a) 6 mol A, (b) 2 mol A, (c) 7 mol A

   a) $? \text{ mol B} = \dfrac{6 \text{ mol A}}{} \left| \dfrac{3 \text{ mol B}}{2 \text{ mol A}} \right. = 9 \text{ mol B}$ b) $? \text{ mol B} = \dfrac{2 \text{ mol A}}{} \left| \dfrac{3 \text{ mol B}}{2 \text{ mol A}} \right. = 3 \text{ mol B}$ c) $? \text{ mol B} = \dfrac{7 \text{ mol A}}{} \left| \dfrac{3 \text{ mol B}}{2 \text{ mol A}} \right. = 10.5 \text{ mol B}$

2. a. Balance this equation: **2** Ca + O_2 $\longrightarrow$ **2** CaO

   b. How many moles of product form from 0.112 mol of oxygen and excess calcium?

   $$? \text{ mol CaO} = \frac{0.112 \text{ mol } O_2}{} \left| \frac{2 \text{ mol CaO}}{1 \text{ mol } O_2} \right. = 0.224 \text{ mol CaO}$$

3. a. Write a balanced chemical equation for the combustion of C_3H_8.

   $$C_3H_8\,(g) + 5\,O_2\,(g) \longrightarrow 3\,CO_2\,(g) + 4\,H_2O\,(g)$$

   b. How many moles of each product form from 4.6 mol of C_3H_8 and excess O_2?

   $? \text{ mol } CO_2 = \dfrac{4.6 \text{ mol } C_3H_8}{} \left| \dfrac{3 \text{ mol } CO_2}{1 \text{ mol } C_3H_8} \right. = 14 \text{ mol } C_3H_8$ $? \text{ mol } H_2O = \dfrac{4.6 \text{ mol } C_3H_8}{} \left| \dfrac{4 \text{ mol } H_2O}{1 \text{ mol } C_3H_8} \right. = 18 \text{ mol } H_2O$

   c. How many moles of oxygen are required?

   $$? \text{ mol } O_2 = \frac{4.6 \text{ mol } C_3H_8}{} \left| \frac{5 \text{ mol } O_2}{1 \text{ mol } C_3H_8} \right. = 23 \text{ mol } O_2$$

4. For the decomposition of potassium chlorate ($KClO_3$) to form potassium chloride and oxygen gas:

   a. Write a balanced chemical equation

   $$2\,KClO_3\,(s) \longrightarrow 2\,KCl\,(s) + 3\,O_2\,(g)$$

   b. What is the theoretical yield of oxygen if 2.00 grams of potassium chlorate are used?

   (Answer: 0.783 g)

   MM $KClO_3$
   39.10
   35.45
   48.00
   122.55 g/mol

   $$? \text{ g } O_2 = \frac{2.00 \text{ g } KClO_3}{} \left| \frac{1 \text{ mol } KClO_3}{122.55 \text{ g } KClO_3} \right| \frac{3 \text{ mol } O_2}{2 \text{ mol } KClO_3} \left| \frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2} \right. = 0.783 \text{ g } O_2$$

   Theoretical yield

5. In this reaction: 2 S + 3 O_2 $\longrightarrow$ 2 SO_3)

   If 5 mol of S and 9 mol of O_2 are used, how many mole of SO_3 can be produced? Which is the limiting reactant, S or O_2?

   $$\text{Using S} = \frac{5 \text{ mol S}}{} \left| \frac{2 \text{ mol } SO_3}{2 \text{ mol S}} \right. = 5 \text{ mol } SO_3$$

   Theoretical yield “S” is Limiting reactant

   $$\text{Using } O_2 = \frac{9 \text{ mol } O_2}{} \left| \frac{2 \text{ mol } SO_3}{3 \text{ mol } O_2} \right. = 6 \text{ mol } SO_3$$

6. What is the limiting reactant and theoretical yield if 0.483 mol of Ti and 0.911 mol Cl_2 are used in this reaction: $Ti + 2\ Cl_2 \longrightarrow TiCl_4$

$$? \text{ mol } TiCl_4 = \frac{0.483 \text{ mol Ti} \mid 1 \text{ mol } TiCl_4}{\mid 1 \text{ mol Ti}} = 0.483 \text{ mol } TiCl_4$$

$$? \text{ mol } TiCl_4 = \frac{0.911 \text{ mol } Cl_2 \mid 1 \text{ mol } TiCl_4}{\mid 2 \text{ mol } Cl_2} = 0.456 \text{ mol } TiCl_4 \longleftarrow \text{Theoretical yield} \quad Cl_2 \text{ is LR}$$

7. In the reaction: $4\ HCl + O_2 \longrightarrow 2\ H_2O + 2\ Cl_2$, when 63.1 g of HCl react with 17.2 g of O_2, Actual yield — 49.3 g of Cl_2 are collected. Determine the limiting reactant, theoretical yield of Cl_2 and percent yield of the reaction. (Answer: 61.4 g, 80.3%)

$$? \text{ g } Cl_2 \text{ using HCl} = \frac{63.1 \text{ g HCl} \mid 1 \text{ mol HCl} \mid 2 \text{ mol } Cl_2 \mid 70.90 \text{ g } Cl_2}{\mid 36.46 \text{ g HCl} \mid 4 \text{ mol HCl} \mid 1 \text{ mol } Cl_2} = 61.4 \text{ g } Cl_2 \longleftarrow \text{Theoretical yield} \quad \text{HCl is LR}$$

$$? \text{ g } Cl_2 \text{ using } O_2 = \frac{17.2 \text{ g } O_2 \mid 1 \text{ mol } O_2 \mid 2 \text{ mol } Cl_2 \mid 70.90 \text{ g } Cl_2}{\mid 32.00 \text{ g } O_2 \mid 1 \text{ mol } O_2 \mid 1 \text{ mol } Cl_2} = 76.\cancel{2} \text{ g } Cl_2$$

$$\% \text{ yield} = \frac{49.3 \text{ g } Cl_2 \mid 100}{61.4 \text{ g } Cl_2 \mid} = 80.3\% \text{ yield}$$

8. If the theoretical yield of a reaction is 0.118 g and the actual yield is 0.104 g, what is the percent yield?

$$\% \text{ yield} = \frac{\text{Actual yield} \mid 100}{\text{Theoretical yield} \mid} = \frac{0.104 \text{ g} \mid 100}{0.118 \text{ g} \mid} = 88.1\% \text{ yield}$$

Molarity 1

There are several ways of expressing solution concentration. The most common is molarity, M (note that is a capital M). The unit M is defined as:

moles solute per liter of solution   or   M = mols solute/L solution

Note that the denominator is liters of *solution*, not liters of *solvent*. So this volume is the **final** volume of the solution, **after** the solute has been added.

Concentrations can be used like conversion factors:

$$\textbf{Molarity} = \frac{\text{mol solute}}{\text{L of solution}}$$

- allows conversion of moles to liters or liters to moles

$$\textbf{Mass percent} = \frac{\text{g of solute} \times 100}{\text{g of solution}}$$

- allows conversion of g solution to g of solute or g solute to g solution

PRACTICE

1. Calculate the molarity of each solution

   $KCl = 39.10 + 35.45 = 74.55$ g/mol

   a. 33.2 g of KCl in 0.895 L of solution (Answer: 0.498 M)

   $$? \text{ M or mol/L} = \frac{33.2 \text{ g KCl}}{} \left| \frac{1 \text{ mol KCl}}{74.55 \text{ g KCl}} \right| \frac{1}{0.895 \text{ L}} = 0.498 \text{ mol/L or M KCl}$$

   b. 61.3 g of C_2H_6O in 3.4 L of solution (Answer: 0.39 M)

   MM $C_2H_6O = 2 \times 12.01 + 6 \times 1.008 + 16.00 = 46.068$
   $24.02 + 6.048 + 16.00 = 46.07$ g/mol

   $$? \text{ M} = \frac{61.3 \text{ g } C_2H_6O}{} \left| \frac{1 \text{ mol } C_2H_6O}{46.07 \text{ g } C_2H_6O} \right| \frac{1}{3.4 \text{ L}} = 0.391 \text{ M } C_2H_6O$$

   c. 38.2 mg of KI in 112 mL of solution (Answer: 2.05×10^{-3} M)

   Note: M = mol/L = mmol/mL, so we can take smaller units to do calculation.

   MM $KI = 39.10 + 126.9 = 166.00$ g/mol

   $$? \text{ M} = \frac{38.2 \text{ g mg KI}}{} \left| \frac{1 \text{ mmol}}{166.00 \text{ mg KI}} \right| \frac{1}{112 \text{ mL}} = 2.05 \times 10^{-3} \text{ mmol/mL} = 2.05 \times 10^{-3} \text{ mol/L}$$

2. Determine how many grams of sucrose ($C_{12}H_{22}O_{11}$) are contained in 1.72 L of 0.758 M sucrose solution. (Answer: 445 g)

   MM Sucrose
   $12 \text{ C} = 12 \times 12.01 = 144.12$
   $22 \text{ H} = 22 \times 1.008 = 22.176$
   $11 \text{ O} = 11 \times 16.00 = 176.00$
   342.3 g

   $$n = M \times V = \text{mol/L} \times \text{L} = \frac{0.758 \text{ mol}}{\text{L}} \left| \frac{1.72 \text{ L}}{} \right. = 1.30 \text{ mol Sucrose}$$

   $$\text{g Sucrose} = \frac{1.30 \text{ mol Sucrose}}{} \left| \frac{342.3 \text{ g Sucrose}}{1 \text{ mol Sucrose}} \right. = 445 \text{ g Sucrose}$$

3. The concentratoin of hydrochloric acid in the stomach is 0.100 M. What volume of stomach acid contains 0.500 g of hydrochloric acid? (Answer: 137 mL)

Molarity 2

Answers are given in parenthesis.

1. A 1.50 mole sample NaOH was dissolved in water to make a volume of 2.0 L. What is the molarity of this solution? (Answer: 0.75 M)

$$M = mol/L = \frac{1.50 \text{ mol NaOH}}{2.0 \text{ L solution}} = 0.75 \text{ M NaOH}$$

2. A 0.300 mole sample NaOH was dissolved in water to make a volume of 150 mL. What is the molarity of this solution? (Answer: 2.0 M)

$$? = mol/L = \frac{0.300 \text{ mol}}{150 \text{ mL} \cdot \frac{1 \text{ L}}{1000 \text{ mL}}} = \frac{0.300 \text{ mol}}{0.150 \text{ L}} = 2.0 \text{ M NaOH}$$

3. A 2.00 g sample of NaOH was dissolved in water to make a volume of 100. mL. What is the molarity of this solution? (Answer: 0.500 M)

$\rightarrow \div 1000 = 0.100$ L

$$mol/L = \frac{2.00 \text{ g NaOH}}{} \left| \frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}} \right|_{= 0.05 \text{ moles}} \frac{1}{0.100 \text{ L}} = 0.500 \text{ mol/L or M}$$

4. How many milliliters of 0.250 M NaOH are needed to provide 0.200 moles of NaOH? (Answer: 800. mL)

$V = n/M$

$$V(L) = mol/(mol/L) = \frac{mol}{} \left| \frac{L}{mol} \right. = \frac{0.200 \text{ mol}}{} \left| \frac{1 \text{ L}}{0.250 \text{ mol}} \right| \frac{1000 \text{ mL}}{1 \text{ L}} = 800. \text{ mL solution}$$

5. How many grams of NaOH are there in 80.0 mL of 0.400 M NaOH solution? (Answer: 1.28 g)

$$n = M \times V = \frac{0.400 \text{ mol}}{L} \left| \frac{0.080{,}0 \text{ L}}{} \right. = \frac{0.032 \text{ mol NaOH}}{} \left| \frac{40.00 \text{ g Na OH}}{1 \text{ mol NaOH}} \right. = 1.28 \text{ g NaOH}$$

6. How many moles of NaOH are contained in 150 mL of 3.00 M NaOH solution? (Answer: 0.45 mol)

$$n = M \times V = \frac{3.00 \text{ mol}}{L} \left| \frac{0.150 \text{ L}}{} \right. = 0.45 \text{ mol}$$

7. What is the molarity of a solution when 0.135 moles of LiOH are dissolved in water to give 50.0 mL of solution? (Answer: 2.70 M)

$$M = \frac{0.135 \text{ mol}}{0.0500 \text{ L}} = 2.70 \text{ M}$$

8. How many grams of $KMnO_4$ would you need to prepare 520 mL of 0.58 M solution? (Answer: 48 g)

$n = M \times V(L)$

$$n = \frac{0.58 \text{ mol}}{L} \left| \frac{0.520 \text{ L}}{} \right. = \frac{0.3016 \text{ mol } KMnO_4}{} \left| \frac{158.04 \text{ g}}{1 \text{ mol } KMnO_4} \right. = 47.66 \approx 48 \text{ g } KMnO_4$$

MM $KMnO_4$
K = 39.10
Mn = 54.94
4 O = 64.00
158.04 g/mol

9. How many milliliters of 5.3 M KBr do you need to get 0.112 moles of KBr? (Answer: 21 mL)

$$V = n/M = \frac{0.112 \text{ mol}}{5.3 \text{ mol/L}} = \frac{0.0211 \text{ L}}{} \left| \frac{1000 \text{ mL}}{1 \text{L}} \right. = 21.1 \text{ mL}$$

Molarity: Applications

1. Consider a solution of $MgCl_2$ in water.
   a. What is the solvent? **Water (H_2O)**
   b. What is the solute? **$MgCl_2$**
   c. Is the solution a homogeneous or a heterogeneous mixture? **Homogeneous**
   d. What molecules, atoms, and/or ions are present in the solution? Write their formulas.

   Mg^{2+} (*aq*) Cl^- (*aq*) ← **H_2O (*l*)**

   e. If the solution had a concentration of 1.0 M and a volume of 1.0 L, what are the specific amounts of the molecules, atoms, and/or ions present in the solution?

   $$n = \frac{1.0\ \text{mol}}{\text{L}} \,\Big|\, 1.0\ \text{L} = 1\ \text{mol} \qquad \underset{1.0\ \text{mol}}{MgCl_2\,(s)} \xrightarrow{H_2O\,(l)} \underset{1\ \text{mol}}{Mg^{2+}\,(aq)} + \underset{2\ \text{mol}}{2\,Cl^-\,(aq)}$$

   Use mol ratios here.

2. A sports beverage has a concentration of sodium phosphate of 0.0020 M.
   a. How many **moles of sodium ions** are there in 1.00 moles of Na_3PO_4?
   b. What is the **molarity of sodium** in this beverage?

   a. $\underset{1\ \text{mol}}{Na_3PO_4\,(s)} \xrightarrow{H_2O\,(l)} \underset{3\ \text{mol}}{3\,Na^+\,(aq)} + \underset{1\ \text{mol}}{PO_4^{3-}\,(aq)}$, **$Na^+$ ion moles are 3.00 mol Na^+.**

   b. $? \ Na^+ \text{ ions} = \frac{0.0020\ \text{M}\ Na_3PO_4 \,|\, 3\ \text{mol}\ Na^+ \text{ ions}}{1\ \text{mol}\ Na_3PO_4} = 0.0060\ \text{M}\ Na^+ \text{ ions}$

3. A 75 mL of 0.211 M NaOH is diluted to a final volume of 125 mL.
   0.075 L
   a. How many moles of NaOH are there?
   b. What is the final molarity of NaOH? (Answer: 0.127 M)

   a. $n = M \times V = \frac{0.211\ \text{mol}}{\text{L}} \,\Big|\, 0.075\ \text{L} = 0.0158\ \text{mol NaOH}$

   b. $M = \frac{0.0158\ \text{mol NaOH}}{0.125\ \text{L}} = 0.127\ \text{M}$

4. An aqueous glucose solution has a volume of 0.800 L and a concentration of 1.2 M.
   a. How many moles of glucose are present in the solution? (Answer: 0.96 mol)
   b. What volume of water should you add to dilute the concentration to 1.0 M? (Answer: 0.96 L)

   a. $n = M \times V = \frac{1.2\ \text{mol}}{\text{L}} \,\Big|\, 0.800\ \text{L} = 0.96\ \text{mol glucose}$

   b. $\frac{1.0\ \text{mol}}{\text{L}} = \frac{0.96\ \text{mol}}{?\ \text{L}}$ or $V(\text{L}) = \frac{n\ \text{moles}}{M(\text{mol/L})}$

   $? \ \text{L} = \frac{0.96\ \text{mol} \,|\, 1\ \text{L}}{1.0\ \text{mol}} = 0.96\ \text{L}$ **or 960 mL**

   Take 0.96 mol of glucose and add water to make it to 960 mL, it will be 1.0 M solution.

Molarity: Fill-in-the-Blanks Table

Complete the table for the following aqueous solutions.

Solute	Mass Solute (g)	Volume (liters)	Molarity (mol/L or M)
LiBr **MM 86.84 g/mol**	43.3 **= 0.498 moles**	0.5	**0.997 ≈ 1 M**
KCl **MM 74.55 g/mol**	37.5 **g** **= 0.503 mol**	**4.02 L**	0.125
$CuSO_4$ **MM 159.62 g/mol**	**0.20 moles** **= 31.9 ≈ 32 g**	1.6	0.125
NaCl **MM 58.44 g/mol**	146.0 **g** **= 2.498 mol**	25.0	**0.0999 M**
$Al(NO_3)_3$ **MM 213.01 g/mol**	**0.3 mol** **= 63.9 ≈ 60 g**	0.5	0.6
$BaCl_2$ **MM 208.23 g/mol**	385.5 **= 1.851 mol**	**0.74 L**	2.5
KOH **MM 56.11 g/mol**	28.0 **g** **= 0.499 moles**	0.25	**2.0 M**
$(NH_4)_2SO_4$ **MM 132.17 g/mol**	**80 g**	0.3	2.0
NaOH **MM 40.00 g/mol**	80.0	**1.3 L**	1.6
NH_4Cl **MM 53.50 g/mol**	160.2	4.0	**0.75 M**

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Answers (from top to bottom): 1 M, 4.02 L, 32 g, 0.999 M, 60 g, 0.74 L, 2.0 M, 80 g, 1.3 L, 0.75 M.
4.02 L 32 g 0.0999 M 60 g 0.74 L 80 g 1.3 L
1 s.f. 3 s.f. 2 s.f. 3 s.f.

Concentration: Mass Percent (% *m/m*)

Concentration: A measure of the amount of solute in a given amount of solution. There are many different ways of expressing concentration; some are more appropriate than others depending on the purpose. Percent concentrations are frequently used with commercial products.

Important Concepts to Remember

- Solution = solute + solvent
- % is $\frac{\text{part}}{\text{whole}} \times 100$
- Masses are additive, volumes are not

Construct an equation for % (*m/m*):

$$\% \text{ m/m} = \frac{\text{g solute} \mid 100}{\text{g solution}}$$

(g solute = g solvent)

PRACTICE

1. What is the percent-by-mass, % (*m/m*), concentration of sucrose in a solution made by dissolving 7.6 g of sucrose in 83.4 g of water?

$$\% \text{ m/m} = \frac{7.6 \text{ g sucrose} \mid 100}{(\underbrace{7.6 \text{ g sucrose}}_{\text{solute}} + \underbrace{83.4\ H_2O}_{\text{solvent}})} = \frac{7.6 \text{ g sucrose} \mid 100}{91.0 \text{ g solution}} = 8.4\% \text{ m/m}$$

2. If 162.35 g aluminum hydroxide are dissolved in 6750 g of solution, what is the % (*m/m*) concentration of the solution?

$$\% \text{ m/m} = \frac{162.35 \text{ g} \mid 100}{6750 \text{ g solution}} = 2.41\% \text{ m/m}$$

3. Calculate the concentration in % (*m/m*) of a 450.0 g solution containing 0.0762 moles of iodine (I_2). (Answer: 4.30%)

$$? \text{ g } I_2 = \frac{0.0762 \text{ mol } I_2 \mid 253.81 \text{ g } I_2}{1 \text{ mol } I_2} = 19.3 \text{ g } I_2$$

$$\% \text{ m/m} = \frac{19.3 \text{ g } I_2 \mid 100}{450.0 \text{ g solution}} = 4.30\% \text{ m/m}$$

Percent Concentration Practice

1. Calculate the mass % for 1.00 mg of Nutrasweet **(NS)** added to 50.0 g of water.

$$\frac{1.00\text{ mg}}{} \left| \frac{1\text{ g}}{1000\text{ mg}} \right. = 1.00 \times 10^{-3}\text{ g}$$

$$\%\text{ m/m} = \frac{1.00 \times 10^{-3}\text{ g NS}}{(50.0\text{ g} + 1.00 \times 10^{-3}\text{ g})} \left| \frac{100}{} \right. = 1.999 \times 10^{-3} \approx 2.00 \times 10^{-3}\%\text{ m/m or } 0.00200\%\text{ m/m}$$

2. Calculate the mass % for 6.6 g of sodium nitrate added to 6.0 L of water. (Assume the density of water is 1.0 g/mL.) (Answer: 0.11%)

$$\frac{6.0\text{ L }H_2O}{} \left| \frac{1000\text{ mL}}{1\text{ L}} \right| \frac{1.0\text{ g}}{1\text{ mL}} = 6000.\text{ g}$$

$$\%\text{ m} = \frac{6.6\text{ g }NaNO_3}{6\bar{0}00 + 6.6\text{ g}} \left| \frac{100}{} \right. = \frac{660}{6\bar{0}06.6} = 0.11\%\text{ m/m}$$

3. Calculate the grams of glucose present in 400.0 g of a solution that is 8.0% (*m*/*m*) glucose. (Answer: 32 g)

$$?\text{ g glucose} = \frac{400.0\text{ g solution}}{} \left| \frac{8.0\text{ g glucose}}{100.\text{ g solution}} \right. = 32\text{ g glucose}$$

means $\frac{8.0\text{ g glucose}}{100\text{ g solution}}$

4. How many grams of a 0.600% (*m*/*m*) penicillin solution should be given to patient who requires 1.0 g of penicillin? (Answer: 170 g)

$$?\text{ g solution} = \frac{1.0\text{ g penicillin}}{} \left| \frac{100\text{ g solution}}{0.600\text{ g penicillin}} \right. = 166.66\text{ g} \approx 170\text{ g solution}$$

5. A glass of milk is about 8.00 fluid ounces. How many grams of fat are in this glass of milk, assuming the milk is 2.0% (*m*/*m*) fat. (Answer: 4.5 g)

$$8\text{ fl oz} = \frac{28.35\text{ g}}{1\text{ oz}} \left| \times 8\text{ oz} = 226.8\text{ g (solution) milk} \right.$$

$$?\text{ g fat} = \frac{226.8\text{ g milk}}{} \left| \frac{2.0\text{ g fat}}{100\text{ g milk}} \right. = 4.5\text{ g fat}$$

Solution Stoichiometry 1

1. Determine the volume in liters of 0.150 M NaOH solution required to neutralize 55 mL of a 0.055 M hydrochloric acid solution. The neutralization reaction is: $NaOH\ (aq) + HCl\ (aq) \longrightarrow H_2O\ (l) + NaCl\ (aq)$ (Answer: 0.020 L)

55 mL → L HCl →
0.055 mol/L HCl → mol HCl → mol NaOH $\xrightarrow{\div M}$ L HCl

$$? \text{ mol HCl} = \frac{0.055 \text{ L solution}}{} \left| \frac{0.055 \text{ mol}}{\text{L solution}} \right. = 0.003\bar{0}25 \text{ mol HCl;}$$

$$? \text{ mol NaOH} = \frac{0.003\bar{0}25 \text{ mol HCl}}{} \left| \frac{1 \text{ mol NaOH}}{1 \text{ mol HCl}} \right. = 0.003\bar{0} \text{ mol NaOH}$$

$$? V(L) = n/M = \frac{0.003\bar{0}25 \text{ mol NaOH}}{} \left| \frac{1 \text{ L}}{0.150 \text{ mol NaOH}} \right. = 0.02\bar{0}2 \text{ L} = 0.020 \text{ L or 20 mL}$$

2. For this reaction: $2\ HNO_3\ (aq) + Na_2CO_3\ (aq) \longrightarrow H_2O\ (l) + CO_2\ (g) + 2\ NaNO_3\ (aq)$

   a. How many milliliters of 0.112 M Na_2CO_3 are needed to completely react with 27.2 mL (= 0.0272 L) of 0.135 M HNO_3? (Answer: 16.4 mL)

   = $0.003\bar{6}7$ mol HNO_3 = 0.00183 mol Na_2CO_3

$$? \text{ mL } Na_2CO_3 = \frac{0.0272 \text{ L solution}}{} \left| \frac{0.135 \text{ mol } HNO_3}{\text{L solution}} \right| \frac{1 \text{ mol } Na_2CO_3}{2 \text{ mol } HNO_3} \left| \frac{1 \text{ L}}{0.112 \text{ mol } Na_2CO_3} \right| \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 16.4 \text{ mL } Na_2CO_3$$

   b. A 25.0-mL (= 0.0250 L) sample of HNO_3 solution requires 35.7 mL (= 0.0357 L) of 0.108 M Na_2CO_3 to completely react with all of the HNO_3 solution. What was the concentration of the HNO_3 solution? (Answer: 0.308 M)

   = 0.00386 mol Na_2CO_3 = 0.00771 mol HNO_3

$$? \text{ M } HNO_3 = \frac{0.0357 \text{ L solution}}{} \left| \frac{0.108 \text{ mol } Na_2CO_3}{\text{L solution}} \right| \frac{2 \text{ mol } HNO_3}{1 \text{ mol } Na_2CO_3} \left| \frac{1}{0.0250 \text{ L solution}} \right.$$

$$= 0.308 \text{ M } HNO_3$$

3. What volume in milliliters of 0.225 M K_3PO_4 solution is needed to completely react with 134 mL of 0.0112 M $NiCl_2$? (Answer: 4.45 mL)

0.134 L

$$2\ K_3PO_4\ (aq) + 3\ NiCl_2 \longrightarrow Ni_3(PO_4)_2\ (s) + 6\ KCl\ (aq)$$

= 0.00150 mol $NiCl_2$ = 0.00100 mol K_3PO_4

$$? V \text{ mL } K_3PO_4 = \frac{0.134 \text{ L solution} \times}{} \left| \frac{0.0112 \text{ mol } NiCl_2}{\text{L solution}} \right| \frac{2 \text{ mol } K_3PO_4}{3 \text{ mol } NiCl_2} \left| \frac{1 \text{ L}}{0.225 \text{ mol } K_3PO_4} \right| \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 4.45 \text{ mL } K_3PO_4 \text{ solution}$$

4. A 10.0-mL sample of an unknown H3PO4 solution requires 112 mL (= 0.112 L) of 0.100 M KOH to completely react with the H3PO4 according to the following reaction. What was the concentration of the unknown H3PO4 solution? (Answer: 0.373 M)

$$H_3PO_4\,(aq) + 3\,KOH\,(aq) \longrightarrow 3\,H_2O\,(l) + K_3PO_4\,(aq)$$

$$? \text{ M } H_3PO_4 = \frac{0.112 \text{ L solution}}{} \left| \frac{0.100 \text{ mol KOH}}{\text{L solution}} \right|_{= 0.0112 \text{ mol KOH}} \frac{1 \text{ mol } H_3PO_4}{3 \text{ mol KOH}} \Bigg|_{= 0.00373 \text{ mol } H_3PO_4} \frac{1}{0.0100 \text{ L } H_3PO_4 \text{ solution}}$$

$$= 0.373 \text{ M } H_3PO_4$$

5. What is the minimum amount in liters of 6.0 M H_2SO_4 necessary to produce 15.0 g of hydrogen gas according to the following reaction? (Answer: 1.2 L)

$$2\,Al\,(s) + 3\,H_2SO_4\,(aq) \longrightarrow Al_2(SO_4)_3\,(aq) + 3\,H_2\,(g)$$

$$? \text{ V(L)} = \frac{15.0 \text{ g } H_2}{} \left| \frac{1 \text{ mol } H_2}{2.016 \text{ g } H_2} \right|_{= 7.44 \text{ g } H_2} \frac{3 \text{ mol } H_2SO_4}{3 \text{ mol } H_2} \Bigg|_{= 7.44 \text{ mol } H_2SO_4} \frac{1 \text{ L solution}}{6.0 \text{ mol } H_2SO_4} = 1.\bar{2}4 \text{ L} = 1.2 \text{ L}$$

Answers: (1) 0.020 L (2a) 16.4 mL (2b) 0.308 M (3) 4.45 mL (4) 0.373 M (5) 1.2 L

Solution Stoichiometry 2

Solve the following solutions stoichiometry problems. Write a balanced chemical equation for each and use dimensional analysis and significant figures.

M = mol solute/Liter solution

1. What volume of 0.496 M HCl is required to neutralize $\overset{=\,0.0200\ \text{L}}{20.0}$ mL of 0.809 M sodium hydroxide? (0.0326L HCl)

$NaOH\ (aq) + HCl\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$

$$?\ \text{V HCl} = \frac{0.0200\ \text{L solution}}{}\ \Bigg|\ \overset{=\,0.0162\ \text{mol NaOH}}{\frac{0.809\ \text{mol NaOH}}{\text{L solution}}}\ \Bigg|\ \frac{1\ \text{mol HCl}}{1\ \text{mol NaOH}}\ \Bigg|\ \frac{\text{L solution}}{0.496\ \text{mol HCl}} = 0.0326\ \text{L HCl}$$

2. How many mL of 0.715 M HCl are required to neutralize 1.25 grams of sodium carbonate? (Producing carbonic acid, H_2CO_3) (0.0330 L HCl)

$Na_2CO_3\ (s) + 2\ HCl\ (aq) \longrightarrow 2\ NaCl\ (aq) + H_2CO_3\ (aq)$

$$?\ \text{mL HCl} = \frac{1.25\ \text{g}\ Na_2CO_3}{}\ \Bigg|\ \overset{=\,0.011\overline{8}\ \text{mol}}{\frac{1\ \text{mol}\ Na_2CO_3}{106.01\ \text{g}\ Na_2CO_3}}\ \Bigg|\ \overset{=\,0.0236\ \text{mol HCl}}{\frac{2\ \text{mol HCl}}{1\ \text{mol}\ Na_2CO_3}}\ \Bigg|\ \frac{1\ \text{L}}{0.715\ \text{mol HCl}}\ \Bigg|\ \frac{1000\ \text{mL}}{1\ \text{L}} = 32.98\ \text{mL HCl} \approx 33.0\ \text{mL HCl}$$

MM Na_2CO_3

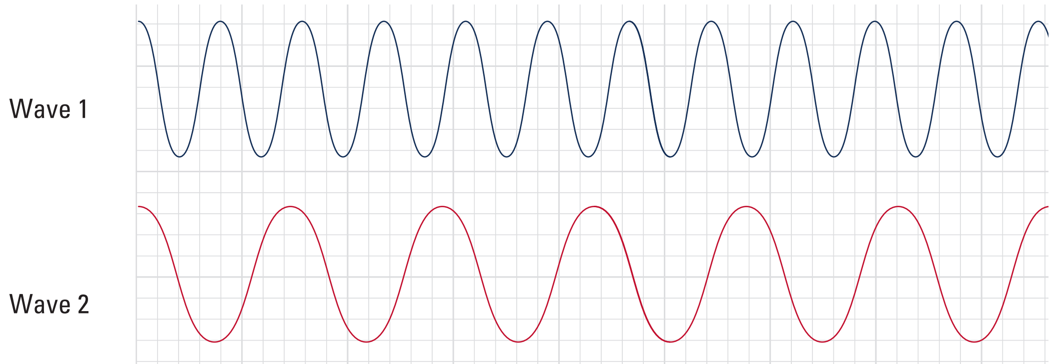
46.00
12.01
48.00
106.01

3. A 20.00-mL sample of H_2SO_4 requires 25.00 mL of 0.500 M NaOH to completely neutralize it. What is the molarity of H_2SO_4? (Answer: 0.313 M H_2SO_4)

4. A sample of $NaHCO_3$ weighing 0.1015 g requires 47.21 mL of HCl(aq) to completely neutralize it. What is the molarity of HCl? (Answer: 0.02559 M)

Activity: Electromagnetic Radiation

ELECTROMAGNETIC RADIATION (LIGHT WAVES)



1. Frequency is the number of waves (cycles) which travel in a given amount of time. Which wave has a higher frequency, wave 1 or wave 2? **Wave 1**

2. Wavelength is the distance between two crests (maxima) or two troughs (minima)—also known as the distance of a repeating unit. Mark this distance for each wave on the picture above. Which wave has a longer wavelength, wave 1 or wave 2? **Wave 2**

3. The higher the frequency, the higher the ***energy*** of the wave. Which wave represents higher energy radiation, wave 1 or wave 2? **Wave 1**

4. True or false? Higher energy light has ~~longer~~ **shorter** wavelengths than lower energy light. **False**

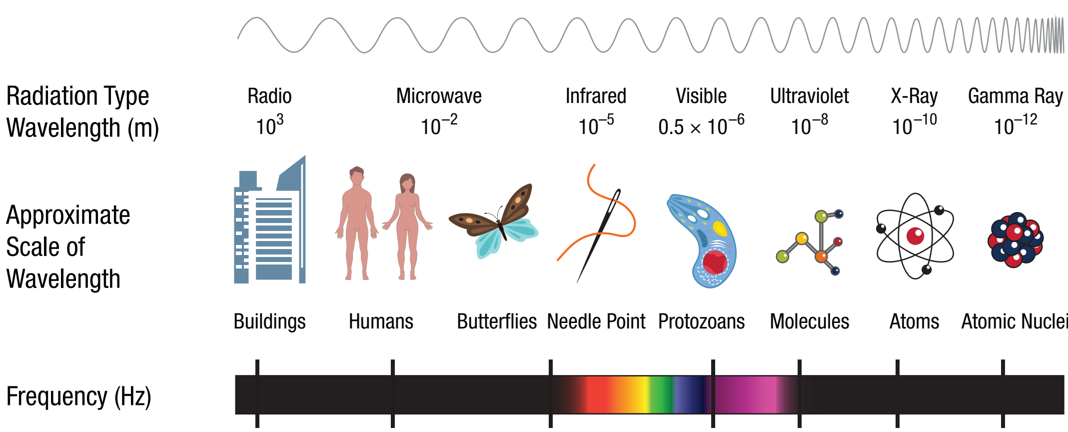


Figure 1. The Electromagnetic Radiation Spectrum

5. List these types of electromagnetic radiation in order of low energy to high energy. (Choices: gamma rays, radio waves, microwaves, ultraviolet, visible light, infrared, x-rays)

Low energy **Radio** **Micro** **Infrared**

Visible **UV** **X-rays** **Gamma** High energy

6. The colors of visible light can be remembered by this mnemonic: ROY G. BIV. What are the colors represented by the letters of this mnemonic?

   R**ed** O**range** Y**ellow** G**reen** B**lue** I**ndigo** V**iolet**

7. If red light has a wavelength of 700 nm (nanometers) and violet light has a wavelength of 400 nm, which color of light is higher in energy? **Violet**

Summary (circle the best answer choice): The **shorter** the wavelength, the ***lower / higher*** its frequency and the ***lower / higher*** its energy. In other words, wavelength and frequency are ***directly / inversely*** proportional, wavelength and energy are ***directly / inversely*** proportional, and the frequency and energy of a wave are ***directly / inversely*** proportional.

Energy, Wavelength, and Frequency

1. Sketch a diagram of a wave and label a crest, a trough, a wavelength, and amplitude.

2. Define frequency of a wave. **Number of wavecycles passing through a point per second.**

3. What are the units and symbol/abbreviations for each of these wave properties?

   a. energy
   Joule (J) or Calorie (Cal)

   b. wavelength
   λ (lambda) in m, nm, etc.

   c. frequency
   ν, 1/s or Hertz (Hz)

4. a. If the frequency (ν) of a wave is increased, what happens to the energy (E) of the wave?

   increases or decreases

   b. Based on your answer, which relationship is correct? (∝ means "proportional to")

   $E \propto \nu$ or $E \propto (1/\nu)$

5. a. If the frequency (ν) of a wave is increased, what happens to the wavelength (λ) of the wave?

   increases or decreases

   b. Based on your answer, which relationship is correct? (c = speed of light)

   $\nu \propto \lambda$ or $\nu \propto (1/\lambda)$

6. These variables can be written into an equation but require constants in order for the mathematics to work. The constants used are Planck's constant **($h = 6.626 \times 10^{-34}$ Js)** and the speed of light **($c = 2.998 \times 10^{8}$ m/s).**

   $E = h\nu$ and $c = \lambda\nu$

   a. Verify that the first equation is consistent with your answer to 4b above. Verify that the units of Planck's constant (h) must be Js. **J = J · s · (1/s)**

   b. Verify that the second equation is consistent with your answer to 5b above. Verify that the units of the speed of light (c) must be m/s. **m/s = m · (1/s)**

7. Use the two equations above to derive a third equation, $E = \mathbf{hc}/\lambda$. This is also a useful equation to use.

   $E = h\nu$ **$c = \lambda\nu$** **$\nu = c/\lambda$** **$E = hc/\lambda$**

EXERCISES

1. The yellow light given off by a sodium vapor lamp used for public lighting has a wavelength of 589 nm. What is the frequency of this radiation?

(Answer: 5.09×10^{14} Hz or 1/s)

$$c = \nu\lambda \text{ or } \nu = c/\lambda = \frac{2.998 \times 10^{8}\ \text{m}}{\text{s}} \cdot \frac{1}{589\ \text{nm}} \cdot \frac{1\ \text{nm}}{10^{-9}\ \text{m}} = 5.09 \times 10^{14}\ \text{1/s or Hz}$$

2. A radio station broadcasts at a frequency of 590. KHz. What is the wavelength of the radio waves?

(Answer: 508 M)

$$\lambda = c/\lambda = \frac{2.998 \times 10^{8}\ \text{m}}{\text{s}} \cdot \frac{1}{590 \times 10^{3}\ \text{Hz}} \cdot \frac{1\ \text{Hz}}{1/\text{s}} = 508\ \text{m}$$

3. A certain microwave has a wavelength of 0.032 meters. Calculate the frequency of this microwave.

(Answer: 9.4×10^{9} 1/s

$$\nu = c/\lambda = \frac{2.998 \times 10^{8}\ \text{m}}{\text{s}} \cdot \frac{1}{0.032\ \text{m}} = 93.7 \times 10^{8} = 9.4 \times 10^{9}\ \text{1/s}$$

4. Microwave ovens emit microwave energy with a wavelength of 12.9 cm. What is the energy of exactly one photon of this microwave radiation?

(Answer: 1.54×10^{-24} J)

$$E = hc/\lambda = \frac{6.626 \times 10^{-34}\ \text{J} \cdot \text{s} \times 2.998 \times 10^{8}\ \text{m}}{\text{s}} \cdot \frac{1}{12.9\ \text{cm}} \cdot \frac{1\ \text{cm}}{10^{-2}\ \text{m}} = 1.54 \times 10^{-24}\ \text{J}$$

5. What is the wavelength of light of a photon with an energy of 4.90×10^{-15} J.

(Answer: 4.05×10^{-11} m)

$$E = hc/\lambda \text{ or } \lambda = hc/E = \frac{6.626 \times 10^{-34}\ \text{J} \cdot \text{s} \times 2.998 \times 10^{8}\ \text{m}}{4.90 \times 10^{-15}\ \text{J} \qquad \text{s}} = 4.05 \times 10^{-11}\ \text{m}$$

6. Calculate the frequency of a gamma ray photon whose energy is 3.33×10^{-13} J.

(Answer: 5.02×10^{20} 1/s)

$$E = h\nu \text{ or } \nu = E/h = \frac{3.33 \times 10^{-13}\ \text{J}}{6.626 \times 10^{-34}\ \text{J} \cdot \text{s}} = 5.02 \times 10^{20}\ \text{1/s}$$

7. Calculate the energy of a photon with frequency of 8.5×10^{14} Hz.

$$E = h\nu = 6.626 \times 10^{-34}\ \text{J} \cdot \text{s} \times 8.5 \times 10^{14}\ \text{1/s} = 5.6 \times 10^{-19}\ \text{J}$$

(Answer: 5.6×10^{-19} J)

8. Calculate the energy of a photon of radiation with a wavelength of 6.4×10^{-7} m.

(Answer: 3.1×10^{-19} J)

$$E = hc/\lambda = \frac{6.626 \times 10^{-34}\ \text{J} \cdot \text{s} \times 2.998 \times 10^{8}\ \text{m}}{6.4 \times 10^{-7}\ \text{m} \qquad \text{s}} = 3.1 \times 10^{-19}\ \text{J}$$

Activity: The Bohr Model

Use a periodic table for this activity.

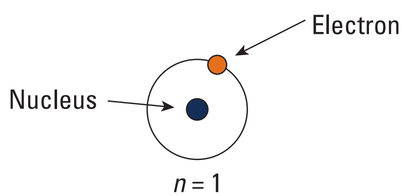
PART 1: BACKGROUND INFORMATION

Ernest Rutherford showed with the **Gold Foil Experiment** that atoms contained a nucleus of positive charge. It was assumed the electrons "floated around" outside the nucleus. What prevented these negatively charged electrons from being sucked into the positively charged nucleus?

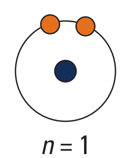
Danish physicist Niels Bohr proposed that electrons traveled on orbits (rings) that prevented them from spiraling into the nucleus. The number of electrons in each orbit can be determined from the periodic table as will be discussed below. These orbits are **QUANTIZED.** This means electrons occupy orbits, and do not exist in between orbits.

Each PERIOD (row) of the periodic table shows a layer of an atom. The atom can be thought of as an onion, with layers of orbits (or shells). For example, hydrogen is in period 1, which represents the first orbit of the atom (closest to the nucleus, $n = 1$). Helium is also in period 1. ***n*** is the orbit number (also called an "energy level").

The Bohr diagrams for each element are shown below:



Bohr Diagram for Hydrogen



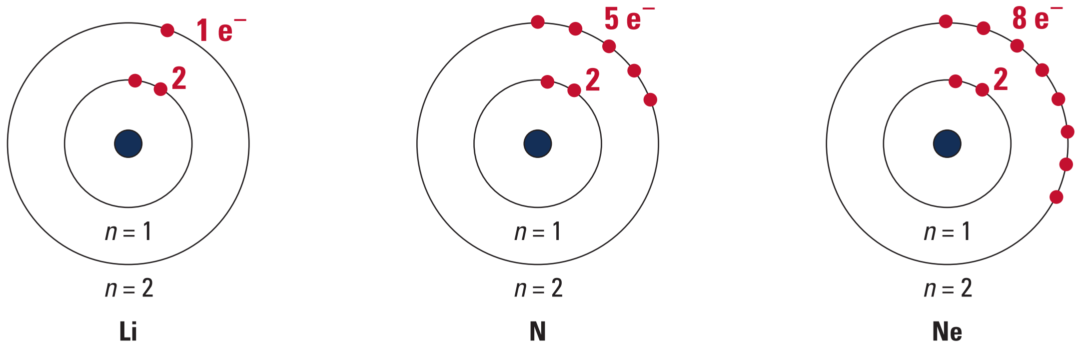
Bohr Diagram for Helium

1. Review: Using a periodic table, how do you know how many electrons are in an atom? For example, how do you determine that helium contains two electrons, as shown above?

   Atomic number = number of protons = number of electrons for an atom

   Elements in PERIOD two will have two layers of an atom, the first energy level being $n = 1$, and the second energy level being $n = 2$. As you can see above, two electrons fit in the first energy level, then the second level begins to fill.

2. Fill in the Bohr diagrams for lithium, nitrogen, and neon below by drawing in the electrons. **HINT:** (1) Determine the number of electrons each element has, then (2) fill in the first orbit, then (3) fill in the second orbit. Use the periodic table to guide you on when to switch from the first to the second "layer." **NOTE:** It only matters how many electrons are in each orbit, not how they are placed within an orbit. Generally we pair the electrons, as shown with helium above.

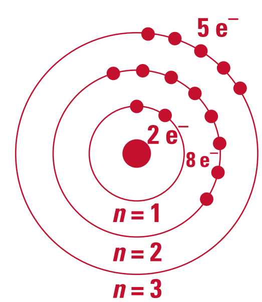


3. Can you guess how to draw a Bohr diagram for phosphorus?

Draw it in the box provided to the right.

*Don't forget to draw the nucleus and the orbits, label the orbits (*n* = 1, etc.), and draw in the electrons.*

Use the periodic table to guide you!

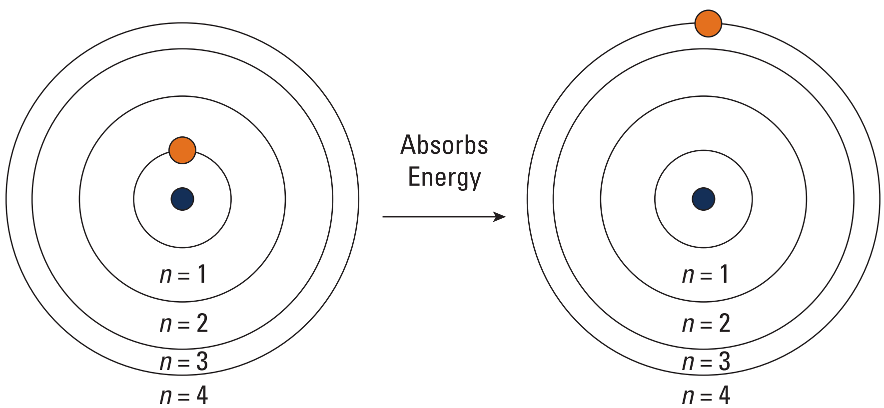


4. The electrons in the outermost energy level play a special role in chemistry. They are called **valence** electrons.
   a. How many valence electrons are in Li, N, and Ne based on your drawings above (Q9)? **1, 5, 8**
   b. What group (column) are Li, N, and Ne in? **1A, 5A, 8A**
   c. Is there a connection between the number of valence electrons and the group number? **Number of valence (e^-) = group number "A"**
   d. How many valence electrons are in P, drawn in Q10? **5**
   e. Why do you think it's similar to N? (**HINT:** Location in periodic table) **Both N and P are in group 5A.**
   f. Without drawing them out, how many valence electrons do you predict are in a sodium (Na) and an argon (Ar) atom? **Na = 1 ve$^-$ (Group IA)** **Ar = 8 ve$^-$ (Group 8A)**

PART 2: HOW DO ATOMS EMIT (RELEASE) LIGHT?

Below is a diagram of hydrogen. Notice that even though hydrogen is in the first period, there are many energy levels in the atom. *Theoretically, every atom contains an infinite number of energy levels.* Because hydrogen only has one electron, all the other energy levels are empty. **It doesn't mean they can't be used later on.** Also note that *the energy levels get closer to each other the farther they are from the nucleus.* Just to keep it simple, let's only consider the first four energy levels.

When hydrogen's electron is in the lowest possible energy level (closest to the nucleus), it is said that hydrogen is in its **"ground state"** configuration. If energy were supplied to this atom (in the form of heat or light), the electron can move to higher energy levels. This results in an **"excited state"** configuration.



Ground State　　　　Excited State

This is called an **electron transition.** If given enough energy, the electron can undergo an $n = 1 \longrightarrow n = 4$ transition, as shown above.

5. What are other possible transitions if the electron absorbs less energy than shown above?

   $n = 1 \longrightarrow$ ***$n = 2$***　　　　$n = 1 \longrightarrow$ ***$n = 3$***

6. Keeping in mind that the electron absorbed energy to "jump" from $n = 1$ to $n = 4$, do you think the electron will ***absorb*** or ***release*** energy when dropping back down from $n = 4$ to $n = 1$? Explain your choice.

   Release. Energy is *released as e^- falls* from higher energy level to a lower energy level.

   Law of conservation of energy. If energy is absorbed for $n = 1$ to $n = 4$ going to higher levels, then energy must be released for $n = 4$ to $n = 1$, the reverse process.

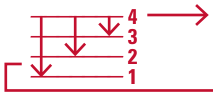
7. Only considering the four energy levels shown, what are the possible electron transitions as the atom releases energy?

   $n = 4 \longrightarrow$ ***$n = 3$***　　　　$n = 4 \longrightarrow$ ***$n = 2$***　　　　$n = 4 \longrightarrow$ ***$n = 1$***

8. Of the transitions you listed above, which one releases the MOST energy? Explain.

$n = 4$ to $n = 1$, this is the largest drop in energy possible.

9. Each transition releases energy in the form of light. If three colors are emitted (red, green, and violet), which transition would go with which color? List them below.



$n = 4 \longrightarrow n = 3$ Red (smallest E) $n = 4 \longrightarrow n = 2$ Green $n = 4 \longrightarrow n = 1$ Violet (greater energy change or largest energy gap)

10. Only considering the four energy levels shown, can you think of any other possible transitions as the atom releases energy? (**HINT:** Does the electron have to start in $n = 4$?) List all the ones your group can think of.

$n = 3$ to $n = 2$ $n = 3$ to $n = 1$ $n = 2$ to $n = 1$

EXERCISES

1. List the following types of electromagnetic radiation in order of increasing energy: gamma rays, radio waves, microwaves, visible light, x-rays, ultraviolet light, infrared radiation.

   Radiowaves < Microwaves < IR < VIS < UV < X-rays < Gamma rays

2. Which color of light comes from a greater energy transition, red or blue? **Blue**

3. Is there a relationship between the Group number of an element and the number of valence electrons in an element? What is this relationship? Write it down:

   Group number "A" = number of valence e^-

4. Based on magnesium being in Group 2A, how many valence electrons do you predict magnesium has? Using the periodic table to guide you, draw the Bohr diagram for a magnesium atom and see if you're correct. **2 ve^-**

5. An electron drops from $n = 4 \xrightarrow{\Delta E_1} 3$, and then $n = 3 \xrightarrow{\Delta E_2} 1$. Two frequencies of light are emitted. How does their combined energy compare with the energy of the single frequency that would be emitted if the electron dropped $n = 4 \xrightarrow{\Delta E_3} 1$?

   $\Delta E_3 = \Delta E_1 + \Delta E_2$ (Law of Conservation of Energy)

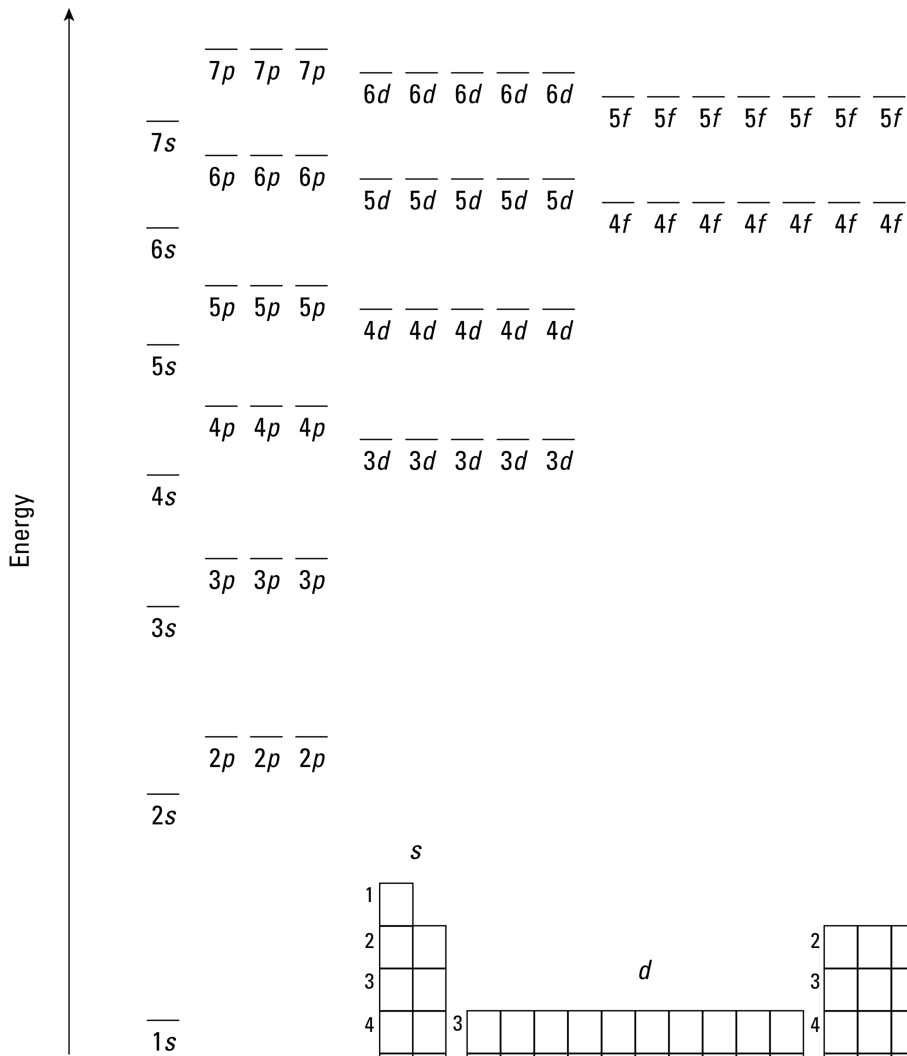
6. What is the meaning of QUANTIZED energy?

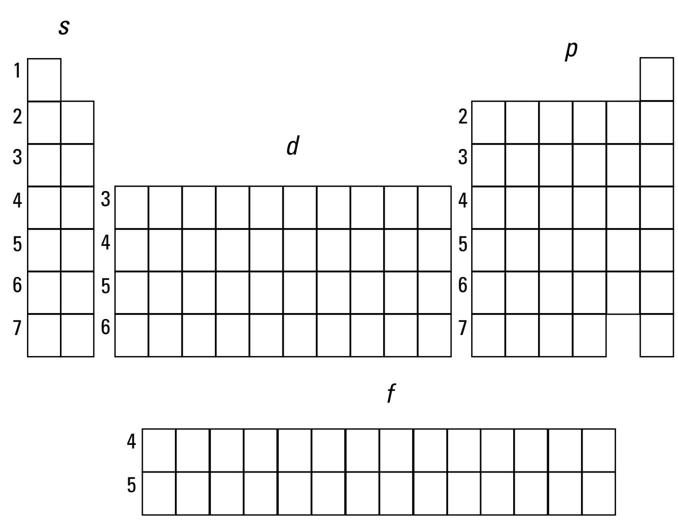
   Only certain amounts of energy can be absorbed or released by atoms. Specific energy associated with an e^- moving in a particular orbit. Electron can not exist in between orbits, 2^+ can exist in any particular orbit only.

7. What might the spectrum of an atom look like if the atom's electrons were not restricted to particular energy levels? (*if the energy were not QUANTIZED*)

   Then we will observe continuous spectrum, all colors of light rather than a line spectrum.

Atomic Orbital Energy Levels





*Energy levels are not drawn to scale.

Electron Configurations 1

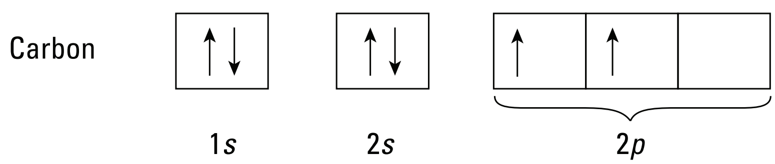
There are three common ways to represent an electron configuration.

1. **The Orbital (Box) Diagram**

   These show each orbital as a box (like the diagram on page 40). The boxes are filled with arrows (representing electrons). The electrons are paired with opposite spin (one arrow pointing up, one arrow pointing down).

   NOTE: Hund's rule says that when orbitals within the same subshell (like 2p, 3d, etc.) are available, you first fill each orbital with ONE electron.

EXAMPLE



2. **The Full Electron Configuration (a short cut of the orbital diagram)**

   This is like the exercise on the previous page.

EXAMPLE

Sodium's full electron configuration is $1s^2\ 2s^2\ 2p^6 3s^1$

3. **The Noble Gas Shorthand Configuration (a short cut of the full configuration)**

   Look at the Noble gas (element in group 8A) that comes BEFORE the element of interest.

EXAMPLE

The Noble gas before Na is Ne.

The Noble gas shortcut for sodium would be [Ne] $3s^1$

(The $1s^2\ 2s^2\ 2p^6$ is substituted for [Ne].) It's very convenient!

EXERCISES

1. Practice writing all three kinds of electron configurations (shown in #1–4 above) for these elements: Al, Xe, Ba, Pb (the last three are long, but it's an exercise!)

2. In the full electron configuration (#1) for all four elements done in Exercise 1, circle all the subshells (ex. $3s^2$) in the highest energy level (1, 2, 3, etc.). These are called valence electrons (they are those electrons in the outermost principal energy level) and are the electrons responsible for chemical reactions.

3. Make a table of each of the four elements (in Part 1), the number of valence electrons, and the group number of each element (Usually given as a Roman numeral in that column of the periodic table). Do you see a relationship between the number of valence electrons and the group #? What is the relationship?

Al = $13e^-$ = $1s^22s^22p^63s^23p^1$, 3ve$^-$ Group 3A

Xe = $54e^-$ = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^6$, 8ve$^-$ Group 8A

Ba = $56e^-$ = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^2$, 2ve$^-$ Group 2A

Pb = $82e^-$ = $1s^22s^22p^63s^23p^63d^{10}4s^24p^64d^{10}4f^{14}5s^25p^25d^{10}6s^26p^2$, 4ve$^-$ Group 4A

Order of fillings is = $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^2$

Electron Configurations 2

1. Write the unabbreviated electron configurations of the following elements:

   a. sodium **= 11e$^-$ = $1s^2 2s^2 2p^6 3s^1$**

   b. iron **= 26e$^-$ = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$**

   c. bromine **= 35e$^-$ = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$**

   d. barium **= 56e$^-$ = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$**

   e. neptunium **= 93e$^-$ = $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 7s^2 5f^5$**

2. Write the abbreviated electron configurations of the following elements:

   a. cobalt **= 27e$^-$ = [Ar] $4s^2 3d^7$**

   b. silver **= 47e$^-$ = [Kr] $5s^2 4d^9$**

   c. tellurium **= 52e$^-$ = [Kr] $5s^2 4d^{10} 5p^4$**

   d. radium **= 88e$^-$ = [Rn] $7s^2$**

   e. lawrencium **= 103e$^-$ = [Rn] $7s^2 4f^{14} 6d^1$**

3. Determine what elements are denoted by the following electron configurations:

   a. $1s^2 2s^2 2p^6 3s^2 3p^4$ **= 16e$^-$ = Sulfur, s**

   b. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^1$ **= 37e$^-$ = Rubidium, Rb**

   c. [Kr] $5s^2 4d^{10} 5p^3$ **= 51e$^-$ = Antimony, Sb**

   d. [Xe] $6s^2 4f^{14} 5d^6$ **= 76e$^-$ = Osmium, Os**

   e. [Rn] $7s^2 5f^{11}$ **= 99e$^-$ = Es = Einsteinium**

4. Determine which of the following electron configurations are not valid:

   a. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 4d^{10} 4p^5$ **$3d^{10}$ should be occupied after 4s, not valid**

   b. $1s^2 2s^2 2p^6 3s^3 3d^5$ **$3p^5$ should be there instead of $5d^5$, not valid**

   c. [Rn] $7s^2 5f^8$ **Can not use alkali earth metal here, only Noble gases are allowed, not valid**

   d. [Kr] $5s^2 4d^{10} 5p^5$ **Okay, is iodine, valid**

   e. [Xe] **54e$^-$ Okay, valid**

From chemfiesta.com.

Electron Configurations 3

The periodic table mimics the electronic structure of the atom. Electrons fill sub-levels from lowest to highest energy. The row of the periodic table that an element is located in corresponds to the number of occupied energy levels in an atom of that element.

The number of electronic sub-levels possible is the same as the electronic energy level number. For example, the first energy level, $n = 1$, has only the "s" sub-level, $n = 2$ has the s and p sublevels, $n = 3$ has s, p, and d sublevels, $n = 4$ has s, p, d, and f sublevels.

1. Write electron configurations for the following elements:

   a. O: **$1s^2 2s^2 2p^4$ or [He] $2s^2 2p^4$**

   b. P: **$1s^2 2s^2 2p^6 3s^2 3p^3$ or [Ne] $3s^2 3p^3$**

   c. Br: **$35e^-$ = [Ar] $4s^2 3d^{10} 4p^5$**

   d. Mn: **$25e^-$ = [Ar] $4s^2 3d^5$**

   e. Pb: **$82e^-$ = [Xe] $6s^2 4f^{14} 5d^{10} 6p^2$**

2. Write electron configurations for each of the following ions:

   a. Br^- **= $36e^-$ (Br gained $1e^-$) = [Ar] $4s^2 3d^{10} 4p^6$ ← $1e^-$ added or $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$**

   b. K^+ **= $18e^-$ (lost $1e^-$) = $18e^-$ = [Ar]**

   c. Ca^{2+} **= Ca lost $2e^-$ = $18e^-$ = [Ar]**

   d. P^{3-} **= P gained $3e^-$ = $18e^-$ = [Ar]**

   e. Al^{3+} **= Al lost $3e^-$ = $10e^-$ = [Ne]**

Orbital Diagrams

1. Aufbau principle: Electrons will fill the lowest energy orbitals first.
2. Pauli Exclusion Principle: Orbitals may hold no more than 2 electrons with opposing spins – symbolized by up and down arrows.
3. Hund's rule: When occupying degenerate orbitals, electrons will distribute to minimize electron-electron repulsion. For example, nitrogen has 5 valence electrons; the 3 electrons in the 2p sub-level will each occupy a different 2p orbital so that each 2p orbital has one electron with parallel spin. This minimizes electron-electron repulsion. Degenerate means "having equal energy."

N = **7e⁻** =

↑↓	↑↓	↑	↑	↑
1*s*	2*s*	2*p*	2*p*	2*p*

Ti = **22e⁻** =

↑↓	↑↓	↓↑	↓↑	↑↓	↑↓	↓↑	↓↑	↓↑	↑↓	↑	↑			
$1s^2$	**$2s^2$**		**$2p^6$**		**$3s^2$**		**$3p^6$**		**$4s^2$**			**$3d$**		

Mg = **12e⁻** =

↓↑	↓↑	↑↓	↑↓	↑↓	↓↑
$1s^2$	**$2s^2$**		**$2p^6$**		**$3s^2$**

Fe = **26e⁻** =

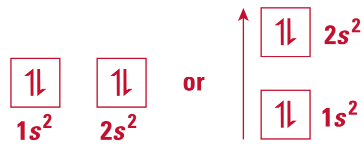
↓↑	↓↑	↑↓	↑↓	↑↓	↓↑	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑
$1s^2$	**$2s^2$**		**$2p^6$**		**$3s^2$**		**$3p^6$**		**$4s^2$**			**$3d^5$**		

Ar = **18e⁻** =

↓↑	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓
$1s^2$	**$2s^2$**		**$2p^6$**		**$3s^2$**		**$3p^6$**	

Review: Electron Configurations

1. How many electrons can a single orbital hold? **2 max**
2. How many total electrons can fit in the 2p subshell of an atom? **6 max**
3. What is wrong with the electron configuration $1s^2 1p^6 2s^2 2p^1$?
   There is no 1p sublevel available.
4. Draw an orbital box diagram for Be.



5. Write the electron configuration for O. **= $8e^- = 1s^2 2s^2 2p^4$**
6. To the right, draw an orbital energy diagram for Al.

7. Write the electron configuration for Ga. **= $31e^- = 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$**
8. Write the condensed electron configuration for K. **= [Ar] $4s^1$**
9. Write the condensed electron configuration for Cl. **= [Ne] $3s^2 3p^5$**

The electrons in the outermost shell (n) of an atom are called the **valence electrons.** Electrons that are not valence electrons are called **core electrons.**

10. How many valence electrons are in a Cl atom? **7 (Group 7A)** How many core electrons? **10**
11. How many valence electrons are in a K atom? **1 (Group 1A)** How many core electrons? **18**
12. How many valence electrons are in F? How about Br? Compare your answers for F and Br to Cl (see above). What is the trend?
    F has 7 and Br has 7. All elements in Group 7A have 7 valence e^-. Group number "A" = number of ve.

Argon, Ar, is in Group 8A—the noble gases. Argon atoms are very stable—they do not like to bond together to form molecules or make compounds with atoms of other elements.

13. How many valence electrons are in an Ar atom? **8**
14. What would a Cl atom need to do to have the same number of valence electrons as Ar? What would a K atom need to do to have the same number of valence electrons as Ar?
    Cl would need to gain $1e^-$. Cl^- = [Ar]. K would need to lose $1e^-$. K^+ = [Ar].

Electron Configurations: Puzzle

A puzzle for you …

Instructions: Read each statement and identify the element described. Write the element symbol in the first column, then identify the first letter of that element symbol and place it in the second column.

Description	Element Symbol	First Letter of Element Symbol
This is one of only two letters that does not appear anywhere on the periodic chart. It is alphabetically the second of the two.	**Q**	**Q**
The third element in the second row of the inner-transition elements.	**U**	**U**
Atoms of this element like to lose $3e^-$ to become isoelectronic with neon.	**Al**	**A**
This nonmetal has the most unpaired electrons in the second period.	**N**	**N**
Atoms of this element like to gain $2e^-$ to become isoelectronic with xenon.	**Te**	**T**
If an atom of this element were to lose $6e^-$ it would have 86 electrons.	**U**	**U**
This neutral atom has an outer shell configuration of $3s^2$.	**Mg**	**M**
This neutral atom has an electronic configuration of $1s^2 2s^2 2p^2$.	**C**	**C**
This neutral atom has an electronic configuration of $1s^2$.	**He**	**H**
The sixth element in the first row of the inner-transition elements.	**Eu**	**E**
The fifth transition metal in period 4.	**Mn**	**M**

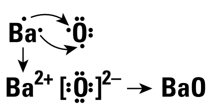
Q **U** **A** **N** **T** **U** **M** **C** **H** **E** **M** !

Adapted from Boyd, S.L. *J. Chem. Educ.* **2007** *84* 619–621.

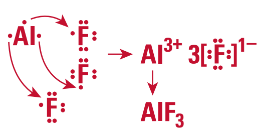
Lewis Dot Structures of Ionic Compounds

1. Draw the Lewis dot structure (drawing) for each of the elements in the table in the first blank column.
2. In the second column, determine whether each element will likely gain electrons (become an anion) or lose electrons (become a cation) and draw the ion it forms. For cations, write the element symbol and the ion charge. For anions, include Lewis dots in your drawing and square brackets around the structure and the ion charge as a superscript.
3. Write the Lewis dot structure for the ionic compound which forms between these two elements. Show the charges and the chemical formula as shown in example 3a.

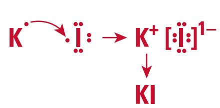
a. Barium and Oxygen



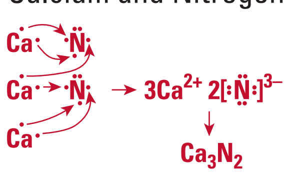
b. Aluminum and Flourine



c. Potassium and Iodine



d. Calcium and Nitrogen



Element	Lewis Dot Structure	Lewis Dot Structure after Reaction (HINT: Cation, Anion, or Neither?)
O	$:\ddot{\underset{\cdot}{O}}\cdot$	$[:\ddot{\underset{..}{O}}:]^{2-}$
Ba	$\dot{Ba}\cdot$	Ba^{2+}
Cl	$:\ddot{\underset{..}{Cl}}\cdot$	$[:\ddot{\underset{..}{Cl}}:]^{1-}$
Al	$\cdot\dot{Al}\cdot$	Al^{3+}
Ar	$:\ddot{\underset{..}{Ar}}:$	Neither
Ca	$\cdot Ca\cdot$	Ca^{2+}
N	$\cdot\ddot{\underset{\cdot}{N}}\cdot$	$[:\ddot{\underset{..}{N}}:]^{3-}$
K	$K\cdot$	K^{+}
I	$:\ddot{\underset{\cdot}{I}}\cdot$	$:\ddot{\underset{..}{I}}:$

Tutorial: Lewis Dot Structures of Covalent Compounds

STEPS FOR DRAWING A LEWIS STRUCTURE

1. Calculate the ***total*** number of valence electrons (v.e.) for the compound.

   # of v.e. = element's Group A # on the periodic table

   - For polyatomic anions, add the number of electrons equal to the charge.
   - For polyatomic cations, subtract the number of electrons equal to the charge.
2. Figure out the central atom (which atom goes in the middle?)
   - Usually it's the one that appears only once in the formula or the one that yields the most symmetric structure (but not always) or the atom that is nearest to metals. For more complicated cases, it may be underlined.
   - Sometimes there can be more than one central atom. Hydrogen and halogens like F, Cl, Br are rarely ever central atoms.
3. Make single bonds to the central atom. Add remaining electrons in pairs to surrounding atoms so each atom gets an octet of electrons. Give pairs of electrons to the outer atoms first. Do all atoms have an octet yet?
4. If there are any remaining electrons, place them in pairs on the central atom. Check if all atoms have an octet.
5. If you run out of electrons and there are atoms without an octet, then use one or more pairs of non-bonding electrons (lone pairs) to bond to the central atom to form double or triple bonds.

EXAMPLE

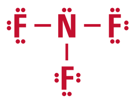
NO_2^-

Lewis Dot Structures 1

Draw the Lewis dot structure for each of these substances.

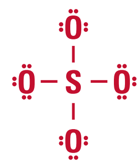
NF_3

$5 + 3(7) = 26\ ve^-$



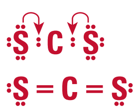
SO_4^{2-}

$6 + 4(6) + 2 = 32\ ve^-$



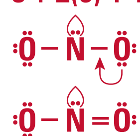
CS_2

$4 + 2(6) = 16\ ve^-$



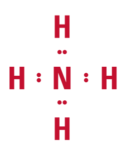
NO_2^-

$5 + 2(6) + 1 = 18\ ve^-$



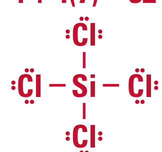
NH_4^+

$5 + 4 - 1 = 8\ ve^-$



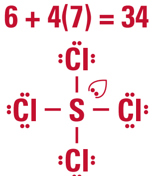
$SiCl_4$

$4 + 4(7) = 32\ ve^-$



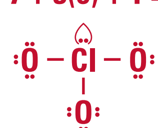
SCl_2

$6 + 4(7) = 34\ ve^-$



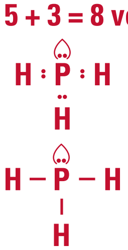
ClO_3^-

$7 + 3(6) + 1 = 26\ ve^-$



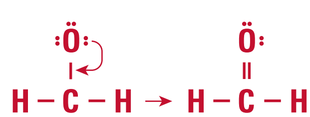
PH_3

$5 + 3 = 8\ ve^-$



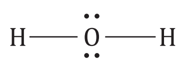
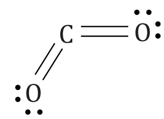
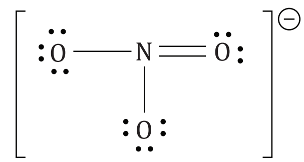
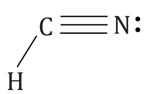
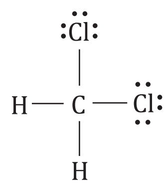
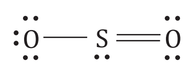
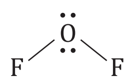
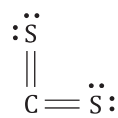
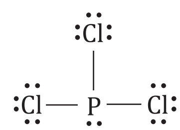
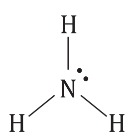
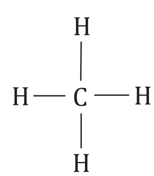
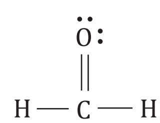
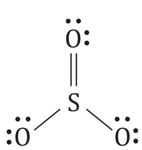
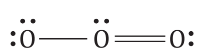
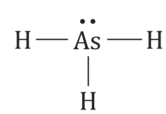
CH_2O

$4 + 2 + 6 = 12\ ve^-$

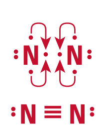


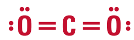
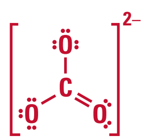
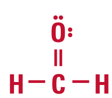
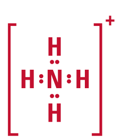
VSEPR / Molecular Shape

The 2-D Lewis structure of each molecule is provided (but not necessarily its 3-D shape). Use VSEPR to identify the electron geometry and molecular shape for each of the following. **NOTE:** Some are drawn with an incorrect shape!

Lewis Structure		
 Geometry: **Tetrahedral** Shape: **Bent**	 Geometry: **Linear** Shape: **Linear**	 Geometry: **Trigonal planar** Shape: **Trigonal planar**
 Geometry: **Linear** Shape: **Linear**	 Geometry: **Tetrahedral** Shape: **Tetrahedral**	 Geometry: **Trigonal planar** Shape: **Bent**
 Geometry: **Tetrahedral** Shape: **Bent**	 Geometry: **Linear** Shape: **Linear**	 Geometry: **Tetrahedral** Shape: **Trigonal pyramid**
 Geometry: **Tetrahedral** Shape: **Trigonal pyramid**	 Geometry: **Tetrahedral** Shape: **Tetrahedral**	 Geometry: **Trigonal planar** Shape: **Trigonal planar**
 Geometry: **Trigonal planar** Shape: **Trigonal planar**	 Geometry: **Trigonal planar** Shape: **Bent**	 Geometry: **Tetrahedral** Shape: **Trigonal planar**

Lewis Dot Structures, Geometry, and Shape

Formula of Compound	Lewis Dot Structure	Electron Geometry	Molecular Shape	Bond Angle
Cl_2 chlorine		Linear	Linear	—
H_2O water		Tetrahedral	Bent	104.5°
CH_4 methane		Tetrahedral	Tetrahedral	109.5°
NH_3 ammonia		Tetrahedral	Trigonal pyramid	109.5°
HF hydrogen fluoride		Linear	Linear	—
N_2 nitrogen		Linear	Linear	—

Formula of Compound	Lewis Structure	Electron Geometry	Molecular Shape	Bond Angle
CO_2 carbon dioxide $4 + 12 = 16$ ve$^-$		Linear	Linear	180°
CO_3^{2-} carbonate $4 + 3(6) + 2 = 24$ ve$^-$		Trigonal planar	Trigonal planar	120°
CH_2O formaldehyde $4 + 2 + 6 = 14$ ve$^-$		Trigonal planar	Trigonal planar	120°
NH_4^+ ammonium $5 + 4 - 1 = 8$ e$^-$		Tetrahedral	Tetrahedral	109.5°
SO_2 sulfur dioxide $6 + 2(6) = 18$ ve$^-$		Trigonal planar	Bent	<120°

Review: Bonding

1. Write the Lewis structure of Mg.

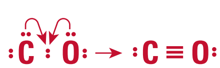
2. Write the Lewis structure of the compound NaBr.

3. Use Lewis theory to predict the formula of the compound that forms between magnesium and nitrogen.

Mg_3N_2

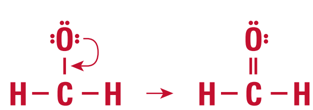
4. Write the Lewis structure for CO.

$4 + 6 = 10\ ve^-$



5. Write the Lewis structure for H_2CO.

$2 + 4 + 6 = 12\ ve^-$

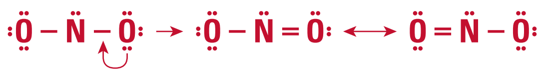


6. Write the Lewis structure for the ClO^- ion.

$7 + 6 + 1 = 14\ ve^-$

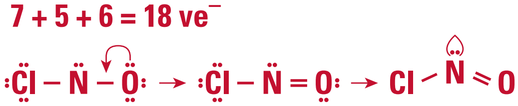
7. Write the Lewis structure for the NO_2^- ion. Include resonance structures.

$5 + 2(6) + 1 = 18\ ve^-$



8. Predict the molecular geometry of ClNO (N is the central atom).

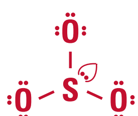
$7 + 5 + 6 = 18\ ve^-$



"Bent"

9. Predict the molecular geometry of the SO_3^{2-} ion.

$6 + 18 + 2 = 26\ ve^-$



Trigonal pyramid

10. Predict whether the atoms will combine to form an ionic or covalent compound. Draw its Lewis structure.

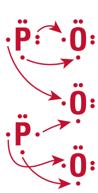
a. I and I

Covalent

b. Cs and Br

Ionic

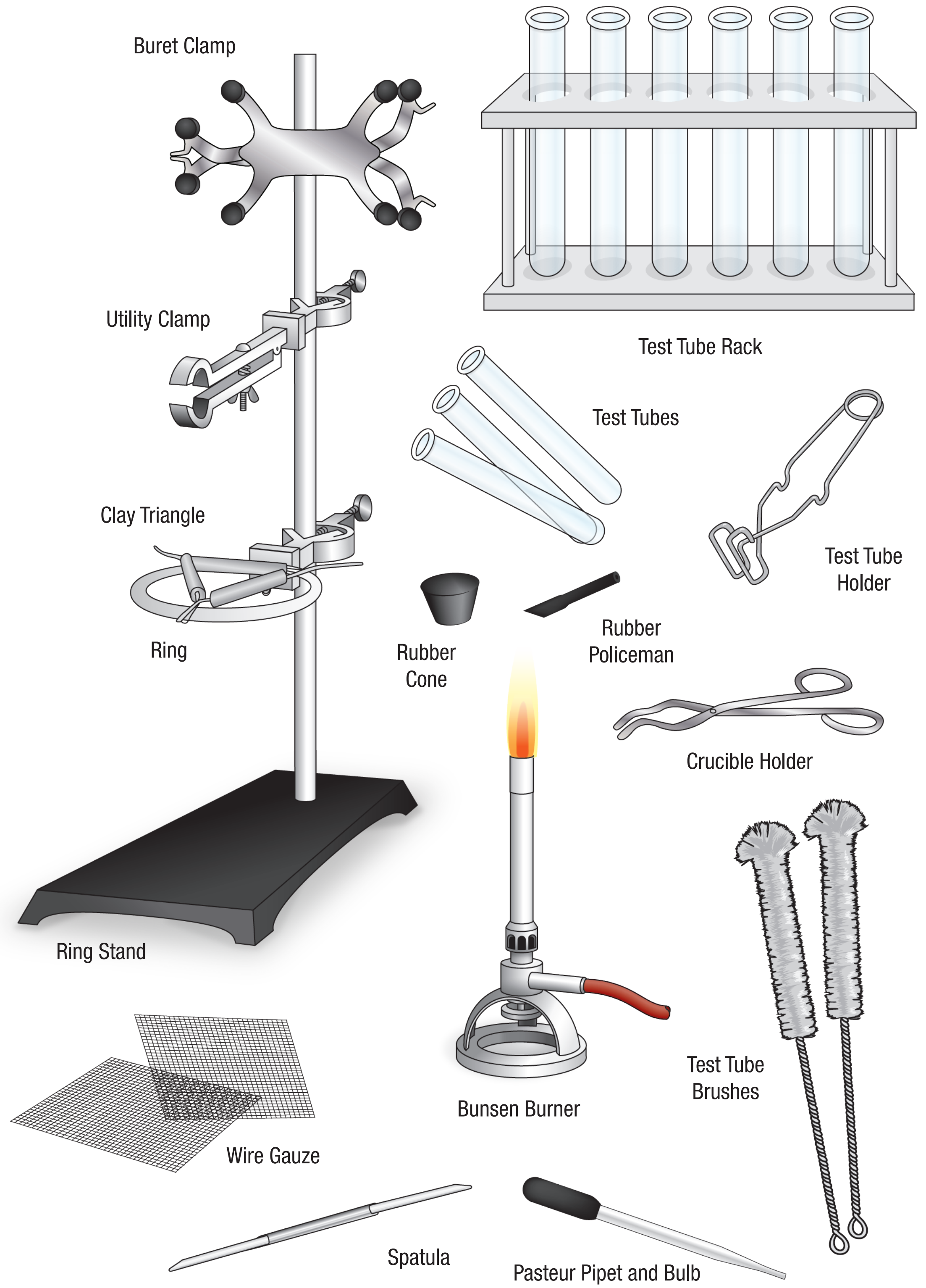
c. P and O

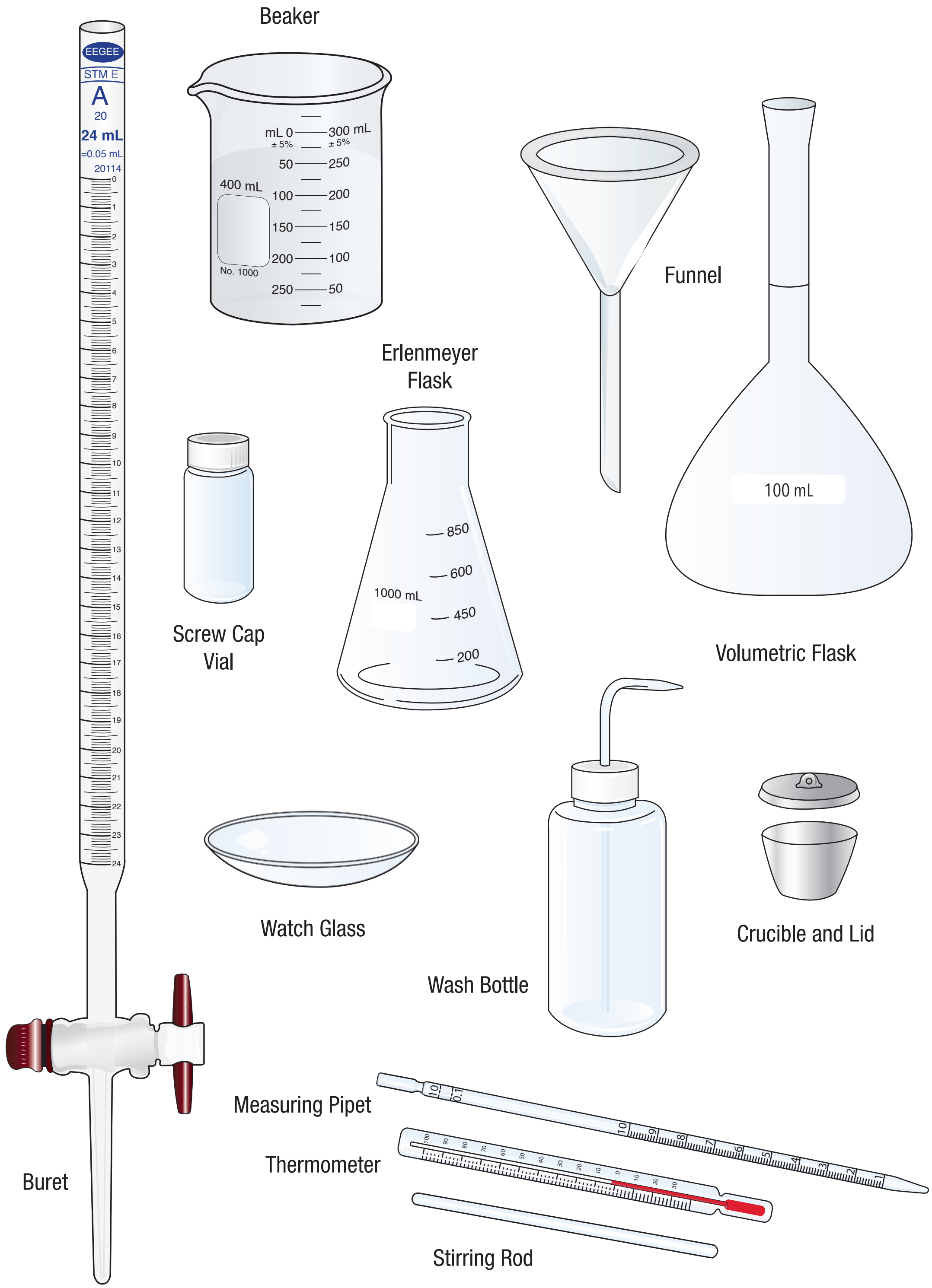


Covalent

P_2O_3

Laboratory Equipment: Names and Pictures




Beaker
400 mL
No. 1000
Funnel
Erlenmeyer Flask
1000 mL
Screw Cap Vial
Volumetric Flask
100 mL
Watch Glass
Wash Bottle
Crucible and Lid
Measuring Pipet
Thermometer
Stirring Rod
Buret
EEGEE
24 mL

EXPERIMENT 1
RUBBER BAND BLAST!

We thank Cynthia Stanich from University of Washington, Seattle, for permission to use this activity.

STUDENT LEARNING OBJECTIVES

- Apply the scientific method to a simple problem.
- Become more familiar with the vocabulary of science.
- Investigate the meanings of accuracy and precision.
- Get to know each other and have a bit of fun!

YOUR MISSION

Your mission today is to figure out how to shoot a rubber band as far as possible by placing it over the end of a finger, adjusting the angle of the finger relative to the ground, and stretching the rubber band some amount prior to releasing it. In other words, your stated purpose for this laboratory should concern the question "What factors cause a rubber band shot from the end of a finger to travel the greatest distance possible?" The following walks you through using the scientific method to answer this question. Each student should fill in their own worksheet to keep.

REPORT SHEETS
Experiment 1: Rubber Band Blast!

Name: ____ Section: __________

Lab Partner: ____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

NAMES OF GROUP MEMBERS

(Please be in a group of 3–4 students!)

__________________________________ __________________________________

__________________________________ __________________________________

1. What do you think is the **purpose** of this activity?

2. **Identify All Possible Variables**

   With your group, brainstorm possible variables which may affect the distance the rubber band travels.

EXAMPLE

Paula can take big steps or little steps when running. These are *independent variables* because she manipulates them to vary the *dependent variable,* which may be speed. A *constant variable* might be the direction she runs, the clothes she wears, or something else.

3. **Hypothesis**

   With these variables in mind, write a hypothesis about what will make a rubber band travel the farthest that you can test today.

4. **Experimental Plan**

   Fill in the choices you make.

   a. Pick one variable to be your independent (or manipulated) variable. ______________
   b. Identify the dependent (or responding) variable. ________________________________
   c. Will you have any constant (or control) variables? If so, what are they? __________
   ____
   d. How will you test your independent variable?

DEFINITIONS

Accuracy is a measure of how close a measurement is to a true value. *Precision* is the measure of how close together the measurements are.

   e. Will you do any repeat trials? Will this choice affect your precision? Your accuracy?

5. **Experimental Work**

   Show your work to your instructor before you do this step. Carry out the experiment and record your data in a table. As you began to collect data, did you find you needed to change anything in your experimental plan?

   Our Data Table:

6. **Data Analysis and Visualization**

Graph your data. Do you think you can throw any points out? Why? Remember that the independent (manipulated) variable always goes on the *x*-axis, while the dependent (responding) variable goes on the *y*-axis. Draw a best fit (or trend) line through your data points; ***DO NOT*** "connect the dots." Ask if you do not know the difference.

Title: ______________________________

RUBBER BAND BLAST-OFF

Once each group is satisfied that they have determined the scientifically proven method for the best way to shoot a rubber band, we will see which group can shoot their rubber band the farthest.

7. **Discussion**

   a. Do your experimental results uphold your hypothesis or contradict it?

   b. If your experimental results did not fully match your hypothesis, write a new hypothesis.

   c. How accurate do you feel your results are (how close to a true value)? Why?

   d. How precise do you feel your data are (how close to other measurements)? Why?

   e. Do your results allow you to answer the original question about what makes a rubber band travel farthest?

   f. Are you ready to write a scientific law or a theory, or do you feel that more experimentation is needed?

EXPERIMENT 2
MEASUREMENTS

OBJECTIVES

- Practice using scientific tools for measurement and recording proper measurements.
- Determine the number of decimal places to record based on the graduations of a measuring tool.
- Practice metric conversions.
- Practice performing calculations rounding to the correct number of significant figures.
- Use different types of glassware, each with a different level of precision.

INTRODUCTION

Measurements are essential to experimental sciences such as chemistry, physics, biology and geology. This experiment is intended to give practice in making **accurate** measurements and being aware of the level of **precision** possible with certain measuring tools.

Measurements in science are usually made using metric system ("SI") units. The precision of a measurement is limited by the calibration of the measurement tool. A balance that gives readings to only the nearest 0.001 gram cannot be used to give masses to 0.00001 gram.

Always *include* ***dimensional units*** *in your measurements.* They are required to perform your mathematical calculations. As such they act as mathematical objects that are at least as important as the numbers themselves.

MATERIALS AND EQUIPMENT

- Balance
- Metric rulers
- Meter stick
- Alcohol thermometer
- 125- or 150-mL Erlenmeyer flask
- 10-, 25-, 50-, and 100-mL graduated cylinders
- Metallic solid unknown

PROCEDURE

MASS MEASUREMENTS

Your instructor will give directions on how to use the balance. Some general rules to follow when using the balances follow:

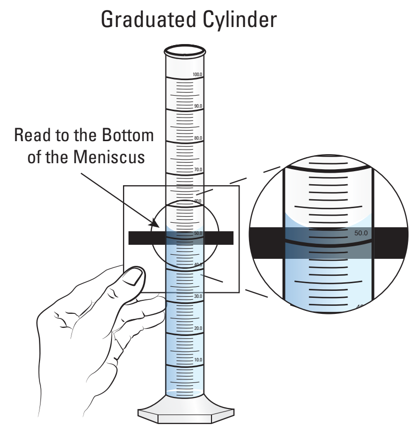
- **Use the "tare" button.** It will set the balance to zero grams, which is your starting point.
- Always **use a container or weighing paper** for weighing chemicals. *DO NOT place chemicals directly on the balance pan!*
- **Clean up** any materials on or near the balance.
- If a balance seems to be out of order, please tell your instructor. *DO NOT attempt to make adjustments to the balance yourself!*
- All measurements must be recorded using the *Rules for Significant Figures.*
- In order to gain practice with the balance,
  - measure the mass of a penny. Include dimensional units. In this case the mass is measured in grams ("g").
  - find an object in the lab that has a mass of about 10 grams.

LENGTH MEASUREMENTS

You will be using a provided ruler printed on paper and the plastic ruler in your lab bench drawer.

VOLUME MEASUREMENTS

There are 10-mL, 25-mL and 100-mL graduated cylinders available. Use the size most appropriate for your measurements. Consult your instructor if you are not sure what size to use.



1. Fill a 125- or 150-mL Erlenmeyer flask to the top mark with water. Then pour that water into graduated cylinder(s) to measure and record its volume.
2. Measure the edge length of the provided cube using the plastic ruler.

TEMPERATURE MEASUREMENTS

Temperature measurements are made using mercury thermometers, thermocouples, gas-filled thermometers, alcohol thermometers, etc. We will be using an alcohol thermometer.

Measurement errors can result from the way the thermometer is located in a liquid. We can minimize some sources of error if we observe the following practices:

- Be sure the liquid is thoroughly mixed.
- Position the thermometer probe away from the walls of the container. You may want to attach a thermometer clamp to the pole of your experiment stand to keep the thermometer in place.
- Allow the thermometer to be in contact with the liquid long enough for the temperature of the thermometer to reach equilibrium with the temperature of the liquid.
- Temperatures should be measured to the precision allowed by the thermometer. If the thermometer scale reads to ±1.0 °C, only readings to the nearest degree are possible.

IDENTIFYING TYPES OF GLASSWARE

Please ask your instructor to display a graduated pipet, a buret, and a volumetric flask, in addition to the beakers, Erlenmeyer flasks, and graduated cylinders that you can find in the shelves at your table.

Please carefully observe the scale (or graduation, increment, or division marks) on the glassware. Also, note in which direction the numbers increase.

Draw a beaker.	Draw a graduated cylinder.	Draw an Erlenmeyer flask.
Draw a pipet.	Draw a buret.	Draw a volumetric flask.

Measurements Practice/ Reference Sheet

Option #1: Set up by your instructor: Various types of chemistry lab glassware filled with volumes of water. Record a measurement for each and then fill in the other columns of information about that measurement and measuring device.

Option #2: On your own in lab: For each type of chemistry lab glassware that you encounter in lab, fill in the columns in the bold box section identifying the smallest increment, uncertainty, and decimal place to which you should record for that measuring device. **Refer back to this table when making measurements in future experiments.**

	Your Recorded Measurement	# of Certain Digits	# of Uncertain Digits	Total # of Sig Figs	Measuring Device	Smallest Increment	Uncertainty (+/–)	Always Record to This Decimal Place
1								
2								
3								
4								
5								
6								
7								
8								

Decimal places:

hundreds	tens	ones	**decimal point**	tenths	hundredths	thousandths
100	10	1	.	1/10	1/100	1/100

REPORT SHEETS
Experiment 2: Measurements

Name: ______________________________ Section: __________

Lab Partner: ______________________________

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

You will be graded based on correct number of significant figures reported for each measurement and calculated value, as well as correct unit abbreviations.

MASS

Record the mass of a penny. ______________

What is an object in the lab that has a mass of approximately 10 g? ______________

LENGTH

Record the length and width in centimeters of the **provided piece of post-it note paper** using the **provided paper ruler.** Include proper unit abbreviations in your measurements.

Length: ______________ Width: ______________

Calculate and report the area of this **provided piece of post-it note paper** determined using this ruler. Include proper unit abbreviations. Show your work.

Area: ______________

Record the length and width in centimeters of the **provided piece of post-it note paper** using the **plastic ruler in your lab bench drawer.**

Length: ______________ Width: ______________

Calculate and report the area of this **provided piece of post-it note paper** determined using this ruler. Include proper unit abbreviations. Show your work.

Area: ______________

What is an object that has a length of about 10 cm? ______________

What is an object that has a length of about 1 m? ______________

VOLUME

VOLUME OF WATER IN AN ERLENMEYER FLASK

Fill a 125- or 150-mL Erlenmeyer flask to the top mark with water. Measure this volume of water using a combination of appropriate graduated cylinder(s). Record all volume readings and calculate the total volume. All volumes must be reported to the appropriate number of significant figures and with correct unit abbreviations.

Volume readings (in mL): ________________ ________________

Total volume (using sig fig rules for addition): ________________

VOLUME OF METALLIC SOLID UNKNOWN

(Use the same metallic solid for the density measurement.)

Using the plastic ruler, measure the edge length of the provided cube in centimeters. Report it to the appropriate number of significant figures and with correct unit abbreviation.

Edge length of cube: ________________

Calculate the volume of the provided cube in **cubic centimeters.** Report it to the appropriate number of significant figures and with correct unit abbreviation.

Volume of cube: ________________

Show your work here.

Calculate the volume of the provided cube in **cubic millimeters.** Report it to the appropriate number of significant figures and with correct unit abbreviation.

Volume of cube: ________________

Show your work here.

TEMPERATURE

Observe the level of graduation of a thermometer and sketch a small portion of it below. Record the temperature of the tap water in the lab. Report it to the proper number of significant figures and with correct unit abbreviation.

Temperature of tap water: ____________

Search the internet to find the freezing point of water in °C: ____________

Search the internet to find the boiling point of water in °C: ____________

LAB EXERCISE
Experiment 2: Measurements

Name: ______________________________ Section: __________

1. What are the base units (and symbols) of measurements of the following in metric system?
   a. Mass ____________________ c. Temperature ____________________
   b. Length ____________________ d. Volume ____________________

2. Write the full names of each of the following units:
   a. mL ____________________ c. km ____________________
   b. ng ____________________ d. °C ____________________

3. Write symbols for each of the following units:
   a. Liter ____________________ c. kilogram ____________________
   b. micrometer ____________________ d. milligram ____________________

4. Convert 5.00 cm into mm. Use the proper number of significant figures.

5. Which is better to use for measuring volume: a 100 mL beaker or a 100 mL graduated cylinder? Explain.

6. In your own words, explain how you use the graduations on a piece of glassware to determine how many decimal places to record a measurement using that glassware.

EXPERIMENT 3
DENSITY

OBJECTIVES

- Practice reporting measurements to the correct number of significant figures.
- Practice using significant figures in calcualtions.
- Determine the density of a liquid.
- Determine the density of a solid by water displacement method.
- Distinguish between intensive and extensive properties.
- Use empirical data to identify an unknown sample.

INTRODUCTION

The density of a substance is an *intensive* physical property. This means that its value does not depend on the *amount* of the substance being measured. It can be used to help identify a material. The formula for calculating the density D is

$$D = \frac{\text{mass in grams}}{\text{volume in milliliters}}$$

or, more generally

$$D = \frac{\text{mass}}{\text{volume}}$$

Intensive properties do not depend on the amount of sample present. Some examples of intensive properties are luster, color, and boiling point. **Extensive properties** are dependent on the amount of sample; some examples include mass and volume. Density is an example of an intensive property.

To determine the density of a substance, you must determine the mass of the substance, a value that can be obtained from weighing the object on a balance.

You must also determine the volume of the substance. The volume of a substance can be determined in different ways.

FOR OBJECTS/SUBSTANCES WITH A DEFINED SHAPE

The first, most straight forward method is to measure the volume in a graduated cylinder. This method is especially useful for determining the volume of liquids but does not work for accurately determining the volume of solids.

If a solid has measurable dimensions, then the volume can be determined using geometry. For a rectangular solid, volume is determined by the equation: where values for the length (L), width (W), and height (H) are multiplied against each other.

$$\textbf{volume of a rectangle} = L \times W \times H$$

In the case of a cubic solid the length, width, and height at all the same value and the equation simplifies to

$$\textbf{volume of a cube} = L^3$$

FOR OBJECTS/SUBSTANCES WITH AN AMORPHOUS SHAPE

Determining the volume of an amorphous (undefined shape) solid can be more difficult than the other two methods. Direct measurement in a graduated cylinder does not account for the air that takes up some of the space, and there are no measurable dimensions to apply to a geometric formula.

For an amorphous solid the volume can be determined by the water displacement method. Fill a graduated cylinder with water. Record the initial volume. Place the substance in the water and the water level rises. Record the final volume. The difference is due to the volume of the substance added.

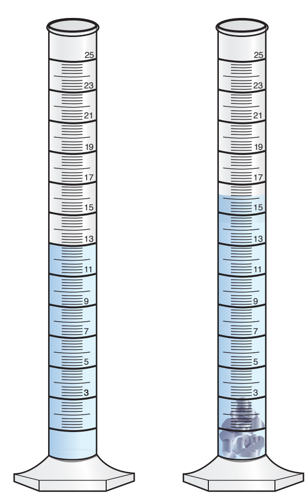


Figure 1. Displacement of water using an amorphous solid

DENSITY OF A SOLID BLOCK

1. Obtain a solid metal block from your instructor. It should have a number written on it. Make sure to record this number on your report sheet as "block # ____."
2. Weigh the solid to the correct amount of significant figures.
3. Using a ruler, measure each side of the block in units of cm. Record using the correct number of significant figures.
4. Calculate the volume in cubic centimeters (cm^3).
5. Calculate the density (in grams per cubic centimeter).
6. Determine the identity of the metal using the table of densities.
7. Check with your instructor on what the identity of the metal is.
8. Based on the correct answer, calculate the % error:

$$\textbf{\% error} = \left(\frac{\textbf{experimental value} - \textbf{true value}}{\textbf{true value}} \right) \times \textbf{100\%}$$

DENSITY OF WATER

1. Weigh a clean, dry 10-mL graduated cylinder. Record its mass.
2. Add tap water to the graduated cylinder, bringing the water level to between 8 and 9 mL. Record the volume of the water to the correct number of significant figures.
3. Weigh the graduated cylinder with the water. Calculate the mass of the water by subtracting the mass of the cylinder from the mass of the cylinder *plus* the water.
4. Use a thermometer to measure the temperature of the water to the nearest °C.
5. Calculate the density from the mass of water and the measured volume using the formula for density.

DENSITY OF AN AMORPHOUS SOLID

1. Obtain a vial containing metal shots. Record the number that is written on the vial on your report sheet.
2. Carefully empty the contents of the entire vial onto a tared piece of weigh paper. Record the mass of the metal shot.
3. Fill a 10-mL graduated cylinder with water so it is about a quarter full. Record the starting volume using the correct number of significant figures.

4. Carefully slide, without splashing, the unknown solid objects into the graduated cylinder so it becomes totally submerged. Observe that the water level rises. It should be below 10 mL in volume. If it is not, start with less water initially.
5. Read and record the final water level (final volume, V_{final}).
6. Calculate the volume of the solid object by subtracting the initial volume from the final volume. $V_{solid} = \Delta V_{water} = V_{final} - V_{initial}$
7. Calculate the density of the solid object using the formula for density and identify which metal it is likely to be.
8. Obtain the true identity of the solid from your instructor and calculate % error.
9. Dry the unknown solid and return it in the vial provided.

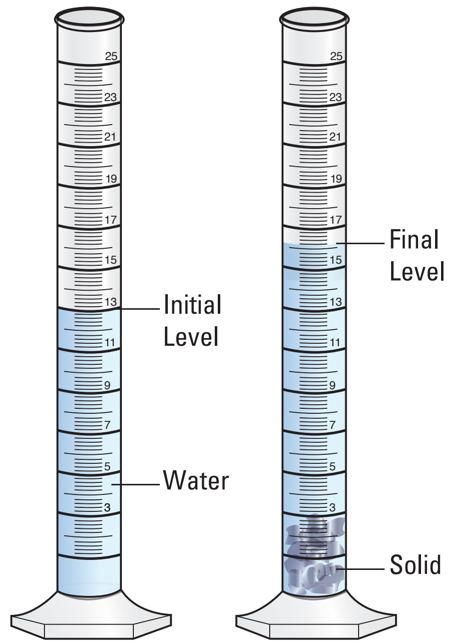


Figure 2. Graduated cylinder before and after addition of unkown solid

Density of Water as a Function of Temperature[1]		**Density of Various Metals**	
Temperature (°C)	Density of Water (g/mL)	Material	Density at 25°C (g/mL)
20	0.9982	Aluminum	2.70
21	0.9980	Zinc	7.14
22	0.9978	Iron	7.87
23	0.9975	Brass	8.55
24	0.9973	Copper	8.94
25	0.9970	Lead	11.3

1 From the *Handbook of Chemistry nand Physics,* 86th edition, 2005–2006.

REPORT SHEETS
Experiment 3: Density

Name: ____ Section: _________

Lab Partner: ____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

You will be graded based on correct number of significant figures reported for each measurement and calculated value, as well as correct unit abbreviations.

DENSITY DETERMINATION FOR THE METAL BLOCK

BLOCK # ________

Report all measurements and calculated values to the appropriate number of significant figures and with correct unit abbreviations.

1. Mass of the metal block: ___________
2. Length, width, and depth of the metal block: _________ _________ _________ cm
3. Show the calculation for the volume of the metal block (include units). Use sig figs.
4. Show the calculation for the density of your metal block (include units). Use sig figs. This will be the "experimental value."
5. Based on the table of densities, your block is made out of ______________________
6. The correct identity of your block is ___________which has a density of __________. (This will be the "true value.")
7. Calculate the % error using this equation:

$$\textbf{\% error} = \left(\frac{\textbf{experimental value} - \textbf{true value}}{\textbf{true value}} \right) \times \textbf{100\%}$$

DENSITY DETERMINATION FOR WATER

Report all measurements and calculated values to the appropriate number of significant figures and with correct unit abbreviations.

Mass of Empty 10-mL Graduated Cylinder	
Mass of Empty 10-mL Graduated Cylinder ***Plus Water***	
Calculated Mass of Water	
Volume of Water	
Calculated Density of Water	
Temperature of the Water to the Nearest Degree C	

1. Fill in the table for the mass of the cylinder empty and with water. Show your mass calculation here and report it as “mass of water” in the table above:

2. Show your density calculation here and report the density in the table above. This is your “experimental value.”

3. Use the table in this lab to find the density of water at the temperature you measured and report it here (include units): _______________. This is your “true value” for the density.

4. Calculate your percent error for the density of water at this temperature. Show your work here and use sig figs.

$$\textbf{\% error} = \left(\frac{\textbf{experimental value} - \textbf{true value}}{\textbf{true value}}\right) \times \textbf{100\%}$$

Your % error: __________

DENSITY DETERMINATION OF A METALLIC SOLID

VIAL # _______

Report all measurements and calculated values to the appropriate number of significant figures and include correct unit abbreviations.

Mass of solid object	
Initial volume of Water in _____-mL Cylinder *(Record the size of the cylinder you used.)*	
Final Volume of Water and Solid Object	
Calculated Volume of Solid Object	
Calculated Density of Solid Object	

1. Fill in the mass, initial volume, and final volume into the table above. Show your volume calculation and report the volume of the solid in the table above. Use sig figs.

2. Show your density calculation and report it in the table above. Use sig figs.

3. Based on the table of densities, your solid is made out of ______________________

4. The correct identity of your solid is _____________ which has a density of ____________. (This will be the "true value.")
5. Calculate your percent error for the density of the solid object. Show your work here:

$$\textbf{\% error} = \left(\frac{\textbf{experimental value} - \textbf{true value}}{\textbf{true value}}\right) \times \textbf{100\%}$$

LAB EXERCISE
Experiment 3: Density

Name: ____ Section: _________

1. a. Write the equation used to calculate density.

   b. If the density and volume of a substance are given, write an equation to calculate the *mass* of this substance.

   c. If the density and mass of a substance are given, write an equation to calculate the *volume* of this substance.

2. Read the procedure. In your own words, summarize how you will determine the density of an amorphous solid.

3. What are the commonly used units of density (in the metric system) for each of the following:

   a. Solids: ____________ b. Liquids or solutions __________ c. Gases ____________

4. a. Describe in your own words the meaning of the term physical property.

   b. Is density a physical property or a chemical property? ______________________

5. Refer to the Introduction section for this experiment.
   a. Describe in your own words the difference between the terms **extensive property** and **intensive property.**

   b. Is the denstiy an extensive property or an intensive property? ________________
6. Search the internet to find:
   a. Density of ice at 0°C: _________
   b. Density of water at 20°C: _________
   c. Density of steam at 100°C: _________
7. What is unusual about density values of water in its different states?

8. Why do density values change with a change in temperature? (**HINT:** What part of the equation for density might change with temperature? What do atoms do as heat is added to a sample?

EXPERIMENT 4
CHEMICAL CHANGES

INTRODUCTION

We can characterize chemical substances by their properties. Some of these are chemical properties (i.e., how they react with other substances), and some of these are physical properties. Most chemical changes are accompanied by changes in physical properties. Since many physical properties can be observed by an experimenter's senses, changes in physical properties are very often used to detect and to provide evidence that chemical changes have taken place.

Simple examples of evidence of chemical change include: a temperature change away from room temperature, changes in phase (a gas, a liquid, or a solid), change in color, solubility or precipitation (forming a new solid), how clear a solution is, and anything new and different or unexpected.

OBJECTIVES

In this experiment you will

- Observe physical changes as evidence that chemical changes have taken place.
- Use vocabulary terms relating to physical and chemical changes.
- Reinforce the practices of chemical safety and waste disposal.

To perform this laboratory, you do not need any prior chemical experience. You will need to be familiar with our laboratory's safety procedures and follow them carefully. You will also need to read labels with care and to follow instructions exactly.

During this experiment you will need to make observations. Some changes are obvious, and some are subtle; some observations are relevant to what you are studying, and some are not. To make a good observation, you need to make a comparison between an initial and a final condition in order to note a change, or better yet, make a comparison between two samples, one of which differs only by the variable you are interested in. The best observations are those that can readily be verified by another experimenter.

Read the section in your textbook that discusses physical and chemical changes. Read the sections in the laboratory manual that discuss safety and waste disposal practices. Then, using your book's index/glossary, a dictionary, or another reference, complete the Prelab assignment for this laboratory and be prepared to have it checked before beginning work.

HAZARDS

The acids and bases used in this experiment can damage eyes and harm skin. Safety goggles must be worn by everyone if any chemical work, including cleanup, is in progress in the laboratory. Wash your hands before leaving lab. You will be rotating to different experimental stations: do not carry out any procedure until you have read through the entire procedure for that section!

PROCEDURE

GENERAL INSTRUCTIONS

- Work with your lab partner to perform the experiments at each of the six stations. Each station will require between two and ten minutes. You may complete the experiments in any order you wish.
- Make sure the station and test tubes are clean for the next group. Since the test tubes will be filled with mixtures containing water, it is not necessary to use dry test tubes.
- Fill in your observations on the first page of the report. As you proceed through the stations, read (but do not fill in) the questions on each reaction that are on the second page of the report; complete these questions after you have completed all of the experiments.

REACTION STATIONS

1. **Calcium Chloride Solution with Sodium Carbonate Solution**
   a. Place about 20 drops of 5% sodium carbonate (Na_2CO_3) solution in a test tube.
   b. Add about 5 to 10 drops of 5% calcium chloride ($CaCl_2$) solution.
   c. Record your observations.
   d. Pour the solution into the waste container and rinse the test tube with water. You do not need to dry the test tube. Leave the station clean for the next group.
2. **Formation of a Potassium Nitrate Solution**
   a. Put about a teaspoon of postassium nitrate (KNO_3, a solid) in a test tube.
   b. Touch the bottom of the test tube to the inside of your arm to get a sense of its temperature.
   c. Add deionized water to the test tube (about 1/3 full) and stir with a glass stirring rod.
   d. Touch the bottom of the test tube to the inside of your arm again. If you don't feel a difference, add more potassium nitrate. Record your observations.
   e. When finished, place the contents of the test tube in the waste container. Rinse the test tube with water. You do not need to dry the test tube. Leave the station clean for the next group.
3. **Limestone with Acidic Water**
   a. Place a limestone chip (use smooth white chips for best results) on a watch glass and add about 5 drops of deionized water. Did anything happen?
   b. Add 5 drops of 3 M hydrochloric acid (HCl). Did anything happen? What is different?
   c. When finished observing, rinse the limestone chips with water and dry for reuse by others. Rinse the watch glass with water into the waste container. You do not need to dry the watch glass. Leave the station clean for the next group.
4. **Vegetable Dye with Acidic and Basic Water**
   a. Obtain 3 test tubes and fill each one with half a pipet full (0.5–1 mL) of red cabbage juice.
   b. To the first test tube, add an equal amount of deionized water. To the second test tube, add an equal amount of 0.1 M HCl (hydrochloric acid). To the third test tube, add an equal amount of 0.1 M $NaHCO_3$ (baking soda).
   c. Note any color changes and record them on your data sheet.
   d. Pour the liquids into a waste container. Thoroughly rinse the test tubes with water. Leave the station clean for the next group.

5. **Bleaching of Paper**
   a. Obtain three test tubes and label them 1, 2, 3. To the first one, add approximately 5 mL of deionized water. To the second one, add approx. 5 mL of chlorox bleach. To the third one, add approx. 5 mL of regular chlorine bleach.
   b. Using scissors, cut three strips of brown/colored paper or bright colored cloth (enough to fit inside the test tube with some of it submerged in the liquid and some of the paper above the liquid). Place one strip into each test tube.
   c. Let it sit for 1 min at room temperature. Record your observations.
   d. Heat the above test tubes in the hot water bath for 5 minutes. Record observations.
   e. Remove the paper from the test tubes. Place the used paper in the garbage. Discard your bleach solutions in the waste container. Rinse out the test tubes with plenty of water, and leave them for the next group (they do not need to be dry). Leave the station clean for the next group.
6. **Aluminum with Copper Ions**
   a. Obtain a test tube and add about 2–3 mL of 5% (*m*/*v*) $CuCl_2 \bullet 2H_2O$ (copper (II) chloride dihydrate) solution.
   b. Record your observations. (Include color, temperature, states of matter, etc.)
   c. To this test tube, add an approximately 1-inch piece of aluminum wire. Touch the test tube to the inside of your arm. Wait for a minute and record your observations. (Include color, temperature, states of matter, etc.)
   d. When finished, place the aluminum wire and copper chloride solution in a labeled waste container. Rinse the test tube with water (it does not need to be dry). Make sure the station is clean for the next group. Leave the station clean for the next group.

REPORT SHEETS
Experiment 4: Chemical Changes

Name: ____ Section: __________

Lab Partner: ____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

DATA

For each reaction, write down your observation(s) before and after the reaction that you believe provide evidence that a chemical change took place. Be clear and concise; use only enough detail to communicate your findings.

Reactions	Observations	
	Before	After
1. Calcium Chloride + Sodium Carbonate		
2. Formation of Potassium Nitrate Solution		
3. Limestone + Acidic Water		
4. Vegetable Dye + Acid/Base Solutions		
5. Bleaching Paper		
6. Al Wire + Copper (II) Chloride Solution		

POST-LAB QUESTIONS

1. **Reaction of Calcium Chloride Solution with Sodium Carbonate Solution:** What term from the pre-laboratory assignment (besides "chemical change") describes this result? Explain.

2. **Formation of a Potassium Nitrate Solution:** It could be ***correctly*** argued that this is a physical change ***or*** that this is a chemical change.
   a. What did you observe that gives evidence that it is a ***chemical*** change?

   b. For what practical purposes do you think that this reaction can be used?

3. **Reaction of Limestone with Acidic Water:**
   a. Were the results for the deionized water similar or different than for the acid?

   b. What do you think acid rain would do to a limestone statue or gravestone?

4. **Reaction of a Vegetable Dye with Acidic and Basic Solutions:** What was the purpose of adding deionized water to one of test tubes containing cabbage juice?

5. **Bleaching of Paper:**
   a. What did you observe that indicates that this is a chemical change and not just a physical change?

   b. Did the reaction occur instantaneously? What do you think would happen if you waited longer?

6. **Aluminum with Copper Ions:**
   a. Based on your observation and chemicals used, what substance do you think deposited on the aluminum wire? What observations explain your choice?

   b. What practical application is there for what you observed?

PRE-LAB
Experiment 4: Chemical Changes

Name: ____________________ Section: ________

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

1. Read through the experimental procedure. Most of the experiments require use of clean test tubes, but why is it not necessary to use dry test tubes?

2. What should you do with any waste materials from this laboratory? If you take excess chemicals that you don't use, what should you do with it?

3. Use your textbook's glossary and index to define the following terms:
   a. physical property
   b. chemical change
   c. reactants
   d. products
   e. precipitate
   f. solution

4. List three types of evidence that a chemical change has taken place. (Refer to Section 6.1 in your Timberlake textbook.)

EXPERIMENT 5
EMPIRICAL FORMULA

OBJECTIVE

Determine the empirical formula of magnesium chloride.

BACKGROUND

The word equation for the reaction we will examine is

> Solid magnesium reacts with hydrochloric acid (an aqueous solution of hydrogen chloride) to produce hydrogen gas and an aqueous solution of magnesium chloride.

$$\mathbf{Mg}\ (s) + \mathbf{HCl}\ (aq) \longrightarrow \mathbf{H_2}\ (g) + ?$$

Hydrochloric acid will be added to a measured amount of magnesium. The amount of hydrochloric acid solution used is in excess to ensure that all of the magnesium reacts. The reaction is quite vigorous and a substantial amount of heat will be released. Reactions that release heat are called exothermic. When the reaction is complete, the hydrogen will have escaped into the room. The beaker will contain magnesium chloride dissolved in water and excess hydrochloric acid. In order to determine the mass of magnesium chloride, the water and excess hydrochloric acid must be removed. This can be accomplished by heating the sample to dryness. The remaining powder is the magnesium chloride.

HAZARD

HCl is corrosive. Avoid contact with skin and clothing. Avoid inhalation. Perform all heating steps in the hood. **Goggles must be worn at all times.**

PROCEDURE

1. Weigh a clean dry and empty 250 ml beaker and record its mass.
2. Weigh approximately 1.00 gram of magnesium in this beaker and record mass.
3. While working ***in the hood, slowly*** add approximately 50 mL of 3 M HCl (hydrochloric acid solution) to the magnesium. Record observations as the reaction proceeds. You may carefully touch the beaker. Do not inhale any of the vapors.
4. After the metal has all reacted, heat the reaction mixture using a hot plate ***in the hood*** to boil off the water and drive off the excess HCl. Avoid splattering of the hot liquid when the volume is low by heating gently.
5. When the material in the beaker seems dry, remove the beaker from hot plate and allow it to cool to room temperature. Weigh the beaker and the magnesium chloride. Record the mass.
6. After weighing, reheat the beaker and solid for 5 minutes on the hot place ***in the hood.*** If you have water droplets condensing towards the top of the beaker, use a cool flame of a Bunsen burner to remove them. Allow the beaker to cool and reweigh. Record the mass. Use the lowest mass reading for calculations. Do not discard the dried magnesium chloride until you complete your calculations and get these checked by your instructor.
7. Scrape off the solid magnesium chloride in beaker and place it in a labeled waste container. You may find it helpful to squirt a little water into the beaker during scraping.
8. Please rinse the beaker with plenty of water in the sink before putting it away.

REPORT SHEETS
Experiment 5: Empirical Formula

Name: ______________________ Section: __________

Lab Partner: ______________________

Show all calculations. Record all data and report all calculated results to the appropriate number of significant figures. Include units.

Mass of beaker and magnesium __________ Mass of magnesium __________

Mass of empty beaker __________ Moles of magnesium __________

Mass of beaker and magnesium chloride after the first heating __________

after the second heating __________

Use the mass after the second heating to complete your calculations.

Mass of magnesium chloride __________

Mass of chloride __________

Moles of chloride __________

Calculate the ratio of moles of the Mg to the moles of Cl ____mole Mg : ____mole Cl
(Do not round to nearest whole number yet.)

Empirical formula __________
(The metal is written first. Round your ratio to the nearest whole number.)

POST-LAB QUESTIONS

1. Why do you think the solid sample of magnesium chloride was re-heated?

2. What changes occurred during the entire experiment? Record your observations here.

3. Write a balanced chemical equation for the reaction that occurs as solid magnesium reacts with hydrochloric acid solution to produce hydrogen gas and an aqueous solution of magnesium chloride. **Include physical states of matter.** (If your experiment did not result in the correct formula for magnesium chloride, do not use your formula in this equation; instead, use the correct formula.)

LAB EXERCISE
Experiment 5: Empirical Formula

Name: ______________________________ Section: __________

Lab Partner: ______________________________

Show your work. Use appropriate number of significant figures and units.

1. A compound containing nitrogen and oxygen is decomposed (separated into nitrogen and oxygen) in the laboratory and produced 1.78 g of nitrogen and 4.05 g of oxygen. Calculate the empirical and molecular formula of this compound by following the steps below.
   a. Calculate the moles of nitrogen formed during the decomposition:

   b. Calculate the moles of oxygen formed during the decomposition:

   c. Calculate the simplest ratio of moles of nitrogen and oxygen:

   d. Write the empirical formula of this compound containing nitrogen and oxygen.

   e. Calculate the empirical formula mass of this compound.

   f. The molar mass of this compound is 92.0 g/mol. Calculate the ratio of the molar mass to the empirical formula mass. Based on this ratio, what is the molecular formula?

   g. Identify this compound as an ionic or molecular compound. Circle your answer.

   h. Write the systematic name of this compound.

EXPERIMENT 6
CHEMICAL REACTIONS

OBJECTIVES

Become more familiar with different types of chemical reactions by

1. Conducting several reactions.
2. Discussing your observations and recording the results.
3. Writing correct formulas and balanced chemical equations.

INTRODUCTION

Chemical reactions take place all around us and even inside us. Chemical reactions can be relatively simple, like the combination of oxygen and hydrogen to form water. They can also be complex, like the synthesis of large protein molecules from thousands of much simpler molecules.

Chemical reactions are changes in the way atoms and molecules combine with each other to form different kinds of matter. Because it is not generally possible to actually see the atoms and molecules in a chemical reaction directly, we need other means to identify when and how a chemical reaction has taken place. Fortunately, many chemical reactions produce observable changes that are easily detected. Color changes, the formation of gases, the formation of solids, heat absorption or emission, and light emission can all be evidence of a chemical reaction. However, they are not definitive evidence that a chemical reaction has taken place. Only chemical analysis can show that the substances we started with initially have been changed into other substances.

In a chemical reaction, the substances present before the chemical changes are called *reactants*, and the substances present after the change are called *products*. Chemical reactions convert reactants into products.

We represent a chemical reaction by a balanced chemical equation, where reactants are stated on the left and products on the right. Instead of the "=" we use in a mathematical equation, chemical equations use "→" to indicate the conversion of reactants into products.

In writing the chemical formula of different substances (reactants and products), we also include the states of matter of the reactants and products using the following convention:

(***g***) – *gas*   (***l***) – *liquid*   (***s***) – *solid*   (***aq***) – *aqueous*

PROCEDURE

Caution: You must wear safety goggles at all times during this experiment.

Follow all directions carefully!

- You will be rotating to different experimental stations.
- Use **fume hoods** for reactions **6, 7, and 8.**

For each of the following reactions answer the following questions:

- How did you know a chemical reaction took place?
- Formula of reactant(s), refer to your text book on writing correct formulas.
- Formula of product(s)
- Write the balanced equation for the reaction (including all physical states).
- Type of chemical reaction

REACTION 1

Place 10 drops of 0.2 M silver nitrate into a small test tube. Add 10 drops of 0.2 M sodium chloride to the same test tube.

> ***Equation:*** Aqueous silver nitrate reacts with aqueous sodium chloride to produce solid silver chloride (also called **precipitate**) and aqueous sodium nitrate.

REACTION 2

Place a small piece of zinc into a small test tube. Add enough 3.0 M hydrochloric acid (HCl) to just cover the piece of zinc.

> ***Equation:*** Solid zinc reacts with aqueous hydrochloric acid to produce aqueous zinc chloride and hydrogen gas.

REACTION 3

Place a small piece of magnesium ribbon (about 10 mm long) into a small test tube.

Add enough 0.1 M copper (II) sulfate to just cover the ribbon. Set aside for 15 minutes. Examine the contents of the test tube.

> ***Equation:*** Aqueous copper (II) sulfate reacts with magnesium to produce aqueous magnesium sulfate and solid copper.

REACTION 4

Place 10 drops of 0.2 M barium chloride into a small test tube. Add 10 drops of 0.2 M sodium sulfate to it.

> ***Equation:*** Aqueous barium chloride reacts with aqueous sodium sulfate to make solid barium sulfate and aqueous sodium chloride.

REACTION 5

Place 20 drops of hydrogen peroxide, H_2O_2, into a small test tube. Add a small amount of manganese (IV) oxide to the tube (just enough to cover the bottom of the test tube).

NOTE: Manganese (IV) oxide is a *catalyst*. It is neither a reactant nor a product: It speeds up the rate of the reaction. A catalyst is indicated in a chemical reaction by *writing the formula of the catalyst over the arrow*.

> ***Equation:*** Aqueous hydrogen peroxide, H_2O_2, produces liquid water and oxygen gas.

REACTION 6 (USE FUME HOOD)

Place a small amount of sulfur into a deflagration spoon. Heat over a Bunsen burner until the sulfur begins to burn.

> ***Equation:*** Solid sulfur (S_8) reacts with oxygen gas to yield sulfur dioxide gas.

REACTION 7 (USE FUME HOOD)

Place a small amount (small scoop) of hydrated copper (II) sulfate into a small test tube. Heat the hydrate over a Bunsen burner until the chemical reaction is complete (the reaction should be obvious).

NOTE: A *hydrate* is a compound that contains water. The formula of a hydrate is written a little differently than formulas for other compounds. The formula for the hydrate of copper (II) sulfate is written $CuSO_4 \cdot 5\ H_2O$. This means there are five water molecules bound to one $CuSO_4$ formula unit. (The correct name for this compound is copper (II) sulfate pentahydrate).

> ***Equation:*** When solid copper (II) sulfate pentahydrate is heated, solid anhydrous copper (II) sulfate and water vapor are produced. (*Anhydrous* means "no water.")

REACTION 8 (USE FUME HOOD)

Use the matches/lighter to ignite the liquid methanol, CH_3OH, in the lamp by lighting the wick. Extinguish the flame by putting the cap over the wick.

> ***Equation:*** Liquid methanol, CH_3OH, reacts with oxygen to produce carbon dioxide gas and water vapor.

REPORT SHEETS
Experiment 6: Chemical Reactions

Name: ____ Section: __________

Lab Partner: ____

For each reaction, write observations, type of reaction, and a balanced chemical equation that includes physical states.

REACTION #1

Observations:

Type of reaction:

Equation:

REACTION #2

Observations:

Type of reaction:

Equation:

REACTION #3

Observations:

Type of reaction:

Equation:

REACTION #4

Observations:

Type of reaction:

Equation:

REACTION #5

Observations:

Type of reaction:

Equation:

REACTION #6

Observations:

Type of reaction:

Equation:

REACTION #7

Observations:

Type of reaction:

Equation:

REACTION #8

Observations:

Type of reaction:

Equation:

POST-LAB QUESTIONS

1. Summarize the major types of visual evidence you observed indicating that a chemical reaction has occurred in your experiment reactions 1 to 8?

2. Why can the formation of a precipitate be evidence of a chemical reaction?

3. When a hydrate compound is heated, does the water get broken down into its component elements, hydrogen and oxygen?

4. When any hydrocarbon compound (C_xH_y) is combusted, what products are formed?

LAB EXERCISE
Experiment 6: Chemical Reactions

Name: ____ Section: __________

1. List at least three possible types of evidence of a chemical reaction.

2. What is a reactant in a chemical reaction?

3. What is a product in a chemical reaction?

4. Give the four symbols for physical states of reactants and products and explain what each means.

5. Why must chemical equations be balanced?

6. Which of the following is a chemical reaction? If it *is* a chemical reaction, give your reasons.

   Liquid water freezes to form solid ice __________

   Copper turn greens on exposure to air __________

   When sodium bicarbonate is combined with hydrochloric acid, bubbling is observed __________

   Butane evaporating out of a butane lighter __________

7. Classify each of the following reactions as a *synthesis*, *decomposition*, *single-displacement*, *double-displacement, or combustion.*

   $2\ Al\ (s) + 2\ H_3PO_4\ (aq) \rightarrow 2\ AlPO_4 + 3\ H_2\ (g)$ ______________________

   $2\ K\ (s) + Br_2\ (l) \rightarrow 2\ KBr\ (s)$ ______________________

   $CuSO_4\ (aq) + 2\ KOH\ (aq) \rightarrow Cu(OH)_2\ (s) + K_2SO_4\ (aq)$ ______________________

   $CuCl_2\ (aq) \rightarrow Cu\ (s) + Cl_2\ (g)$ ______________________

8. Balance the following equations:

   ______ $Sb\ (s)$ + ______ $Cl_2\ (g)$ → ______ $SbCl_3\ (s)$

   ______ $C_2H_6\ (g)$ + ______ $O_2\ (g)$ → ______ $CO_2\ (g)$ + ______ $H_2O\ (g)$

   ______ $Fe\ (s)$ + ______ $HCl\ (aq)$ → ______ $FeCl_3\ (aq)$ + ______ $H_2\ (g)$

EXPERIMENT 7
SYNTHESIS OF ALUM

OBJECTIVES

- Apply the principles of synthetic chemistry, the process by which one pure substance is transformed into another.
- Practice stoichiometric calculations including theoretical and percent yield.

INTRODUCTION

In this experiment we will perform a *synthesis*—taking one or more compounds to form another. During the synthesis we will transform the aluminum in aluminum foil into a substance known as alum, which has the chemical formula $KAl(SO_4)_2 \cdot 12\ H_2O$. (This chemical formula is described on the next page.)

Alum is a versatile substance used in pickling, leather tanning, dyeing fabrics, treatment to fireproof textiles, and water purification in municipal water treatment plants. Alum is also an ingredient in baking powder and aftershave products. In medicine, it has applications as an astringent (constriction agent) and a styptic (to coagulate the blood and stop the bleeding).

Aluminum foil is made of pure aluminum (Al).

STEP 1

The first step in this experiment is to react the aluminum foil with potassium hydroxide (KOH) solution. The KOH helps to dissolve the thin oxide coating on the aluminum that protects it from further oxidation. This is an oxidation-reduction reaction, in which the aluminum is oxidized and the water is reduced. Transfer of electrons from the aluminum to the water produces hydrogen gas and hydroxide ions. The hydroxide ion then combines with the dissolved aluminum cation to form a complex ion $Al(OH)_4^-$, the tetrahydroxyaluminate ion.

The balanced chemical equation for this oxidation-reduction reaction is

$$2\ Al\ (s) + 2\ KOH\ (aq) + 6\ H_2O_{(l)} \rightarrow 2\ KAl(OH)_4\ (aq) + 3\ H_2\ (g)$$

STEP 2

The second step of the procedure is to convert the $KAl(OH)_4$ to alum by the addition of sulfuric acid (H_2SO_4). Addition of a slight excess of sulfuric acid to the products neutralizes the excess KOH and 4 OH- groups on the $Al(OH)_4^-$ ion. This is an acid-base neutralization reaction in which acid and base combine to form water. The solution then contains Al^{3+}, K^+ and SO_4^{2-} ions in the proportions in which they are found in alum. The formula of crystalline alum is $KAl(SO_4)_2 \cdot 12\ H_2O$. There are 12 water molecules associated with each formula unit of the salt.

The alum has a limited solubility in water, so it precipitates from the solution. The balanced chemical reaction that occurs in this step is

$$KAl(OH)_4\ (aq) + 2\ H_2SO_4\ (aq) + 8\ H_2O\ (l) \rightarrow KAl(SO_4)_2 \cdot 12\ H_2O\ (s)$$

STEP 3

Next, the solution containing the dissolved alum is cooled on ice. The solubility of alum decreases as the temperature is lowered. Cooling the alum solution increases the amount of solid alum that can be recovered.

To get pure alum, it is necessary to filter out the crystals and wash them with ethanol to remove the excess H_2SO_4. The ethanol will evaporate quickly and does not dissolve the alum. The crystals of alum can be dried at room temperature.

SAFETY

- *You must wear safety goggles at all times* during this experiment.
- Gloves recommended for strong acids & bases.
- 1.5 M KOH **(caution: caustic/strong base)**
- 3 M H_2SO_4 **(caution: corrosive/strong acid)**

PROCEDURE

1. Cut a piece of aluminum foil 10 × 13 cm (this yields between 0.5–0.6 g Al).
2. Cut the aluminum into small squares, about 0.5 × 0.5 cm.
3. Weigh an empty 150-mL beaker and record its mass to the nearest milligram.
4. Add the 0.5–0.6 g of small pieces of aluminum (from Steps 1 & 2) and record the mass of the beaker containing the aluminum to the nearest milligram.
5. Measure 25 mL of 1.5 M KOH solution in a graduated cylinder.

Work under the fume hoods!

1. Add the 25 mL of 1.5 M KOH to the 150 mL beaker that contains the small pieces of aluminum.
2. Place the beaker on a hot plate set at low heat. Warm the solution: do not allow it to boil. The aluminum should react and be consumed in about 20 minutes.
3. Set up a glass funnel[1] using a ring stand and iron ring to support the funnel.
4. Place the stem of the funnel into a second, clean 150-mL beaker.
5. Fold a large circular filter paper into quarters and place it into the funnel[2]. Pour a little water onto the filter paper so that it will cling to the inner surface of the funnel and stay in place.
6. Remove the beaker from the hot plate. Be sure that all aluminum has reacted.
7. Obtain 15 mL of 3 M H_2SO_4 and carefully and rapidly pour all of the H_2SO_4 into the beaker, stirring steadily to avoid formation of insoluble $Al(OH)_3$ precipitate. If white crystals are visible, return the beaker to the hot plate and stir until all of the material is dissolved. You may need to add up to 5 mL more of the 3 M H_2SO_4 if the white crystals do not disappear adequately.

1 This is a glass gravity funnel not a vacuum Büchner funnel.

2 If you haven't installed filter paper into a funnel before, please consult the instructor.

8. Pour the hot solution into the previously prepared glass filter funnel. *Caution*: Do not overfill the filter funnel *and* do not touch the filter paper with the stirring rod. The filtrate (the liquid that has passed through the filter) should be clear and colorless. Save the filtrate for next step. Discard any residual crystals on the filter paper in a labeled waste container and the filter paper in the trash.
9. Allow the 150-mL beaker containing the filtrate to cool.
10. Take a 600-mL beaker and fill it halfway with crushed ice. Add water to this to create a 0 °C ice bath.
11. Place the 150-mL beaker, containing the filtrate, in the ice bath for about 20 minutes, allowing it to cool. Stir the filtrate mixture frequently as the crystals of alum form.
12. Insert a Büchner funnel into a vacuum filter flask and attach the flask's nozzle to a vacuum hose. Attach the other end of the hose to the vacuum nozzle in the hood.
13. Select a 7.5 cm filter paper. Using a pencil, label the filter paper with your name and record its mass. Place the filter paper evenly around the inside of the funnel so it can catch all of the alum you've produced.
14. Moisten the filter paper and turn on the vacuum to the Büchner funnel.
15. Remove the 150-mL beaker from the ice bath. Stir the contents and transfer all of the crystals to the Büchner funnel. You may use a plastic spatula to make sure you transfer all of the crystals produced.
16. Add two 10-mL portion of ethanol to wash the excess H_2SO_4 out of the crystals.
17. Continue to draw air through the Büchner filter for 15 minutes to dry the alum crystals.
18. Using the forceps (tweezers), remove the filter paper containing the alum crystals from the Büchner funnel and allow alum crystals to air dry till next week.
19. Weigh 10 g of sodium bicarbonate. Add it to the remaining solution in the filter flask. Swirl the flask until the bubbling ceases. Add more $NaHCO_3$ until you do not observe any more fizzing. This is a neutralization reaction of the sulfuric acid H_2SO_4 by sodium bicarbonate $NaHCO_3$:

$$H_2SO_4\,(aq) + 2\,NaHCO_3\,(s) \longrightarrow Na_2SO_4\,(aq) + 2\,CO_2\,(g) + 2\,H_2O\,(l)$$

20. Pour the neutralized solution (filtrate) in the filtration flask down the sink[1] and rinse the flask.
21. After the crystals have dried, weigh the filter paper with alum crystals. Record the mass and calculate the mass of alum, $KAl(SO_4)_2 \cdot 12\,H_2O$ (s), crystals obtained. This is the **actual yield** of alum.
22. Calculate the **theoretical yield** of alum based on the mass of aluminum foil used as aluminum is the limiting reactant in this reaction. Calculate the **percent yield** of alum after you recorded the mass of dried alum crystals.

1 The neutralized solution will not damage the pipes.

REPORT SHEETS
Experiment 7: Synthesis of Alum

Name: ______________________________ Section: __________

Lab Partner: ______________________________

Data	
Mass of Empty 150-mL Beaker	
Mass of Empty 150-mL Beaker + Aluminum	
Mass of Aluminum Metal Used (calculated)	
Mass of Filter Paper	
Mass of Filter Paper + Alum Crystals	
Mass of Alum Crystals Obtained (calculated)	

Show your work and express your answer in correct number of significant figures based on your data. Use appropriate units.

1. Calculate the number of moles of aluminum that you used. __________
   Show your calculations.

2. Calculate the molar mass of alum $KAl(SO_4)_2 \cdot 12\ H_2O$ and round it to two decimal places. Include units. (**HINT:** The dot means each crystal of $KAl(SO_4)_2$ contains 12 H_2O molecules. How does this affect its molar mass?) __________
   Show your calculations.

3. Calculate the maximum number of grams of alum that you could produce (**theoretical yield**), assuming that aluminum is the limiting reactant. __________

4. Refer to your data table. Which piece of data is your **actual yield** of alum? Rewrite your actual yield here. __________

5. Calculate the percentage yield of alum. __________
   Show your calculations.

$$\textbf{\% yield} \ \frac{\textbf{actual yield}}{\textbf{theoretical yield}} \times \textbf{100}$$

LAB EXERCISE
Experiment 7: Synthesis of Alum

Name: ____ Section: __________

Show your work and report your answer in correct number of significant figures.

1. For the reaction, ammonia reacts with oxygen to produce nitrogen monoxide and water, answer the following questions:

   ______NH_3 (*g*) + ______O_2 (*g*) $\rightarrow$ ______NO (*g*) + ______H_2O (*g*)

   a. Balance the equation to whole number ratios.

   b. How many molecules of water can be produced from 60. molecules of O_2?

   c. How many moles of O_2 would be required to react with 0.330 moles of NH_3?

   d. If 2.00 g of NH_3 are used, how many moles of NO can be produced?

   e. If 1.11 moles of NH_3 are consumed in a reaction, how many grams of H_2O are produced?

   f. How many grams of O_2 would be required to react with 2.00 g of NH_3?

2. For the reaction of nitrogen gas with hydrogen gas to produce ammonia gas, NH_3, answer the following questions:
   a. Write a balanced chemical equation for this reaction, include physical states of matter.

   b. Calculate the **theoretical yield** of ammonia, in grams, when 10.00 g of N_2 is reacted.

   c. Calculate the **percent yield** of this reaction when 10.00 g of N_2 is reacted and 9.10 g of ammonia is produced.

      HINT: See the previous part in which you calculated the theoretical yield of ammonia based on 10.00 g of N_2 reacting. The actual yield is the actual amount produced in the reaction, not the theoretical (expected or calculated) yield or the starting reactant amount.

$$\textbf{\% yield} \ \frac{\textbf{actual yield}}{\textbf{theoretical yield}} \times \textbf{100}$$

EXPERIMENT 8
STOICHIOMETRIC ANALYSIS OF AN ANTACID

INTRODUCTION

In this lab, you will use the concept of stoichiometry to solve two sequential problems. First, you will try to determine the products of a certain reaction (below), choosing between three possibilities. Then, you'll use your results of this first part to determine the amount of sodium bicarbonate in a common household substance.

$$CH_3COOH\ (aq) + NaHCO_3\ (s) \rightarrow CO_2\ (g) + ???$$

You've probably seen this reaction in elementary school—add a few drops of red food coloring, and you have the classic "volcano reaction." Or, you can perform it easily in your kitchen by mixing vinegar (dilute acetic acid) and baking soda (sodium bicarbonate). The most noticeable sign of the reaction is vigorous bubbling, a result of very rapid carbon dioxide generation.

Gaseous carbon dioxide is one of the products, as you can see with your own eyes. (You can prove the gas to be carbon dioxide by collecting it in a flask, and inserting a burning match into the flask. The flame will be immediately extinguished.) Aside from carbon dioxide, what else is produced by the reaction?

A chemist approaching this problem would most likely form some hypotheses about the other products and then design experiments to evaluate which hypothesis is best supported by experimental evidence. For this experiment, we'll supply three possible reactions, shown below. Notice that they are all balanced equations.

A. $CH_3COOH\ (aq) + NaHCO_3\ (s) \rightarrow 2\ CO_2\ (g) + CH_2O\ (aq) + Na^+\ (aq) + 3\ H^+\ (aq)$

B. $CH_3COOH\ (aq) + NaHCO_3\ (s) \rightarrow CO_2\ (g) + H_2O\ (l) + CH_3COO^-Na^+\ (aq)$

C. $CH_3COOH\ (aq) + 2\ NaHCO_3\ (s) \rightarrow CO_2\ (g) + Na_2CO_3\ (aq) + H_2O\ (l) + 2\ CH_2O\ (aq)$

Your job is to determine which of these three possibilities is correct, using some simple laboratory measurements. Focus on the differences between the three proposals: Reactions A and C produce formaldehyde (CH_2O), but Reaction B doesn't. The products of Reaction A are acidic (H^+ is produced); those of B and C are basic ($CH_3COO^-Na^+$ and Na_2CO_3 are produced). These things could be tested, but an even simpler method would be take advantage of the different amounts of carbon dioxide produced, relative to sodium bicarbonate:

In Reaction A, 1 mole $NaHCO_3$ produces 2 moles CO_2

In Reaction B, 1 mole $NaHCO_3$ produces 1 mole CO_2

In Reaction C, 2 mole $NaHCO_3$ produces 1 mole CO_2

You will measure the ratio of moles $NaHCO_3$ used to moles of CO_2 produced, and if it is approximately 1:2, you may conclude that Reaction A is correct; if the ratio is around 1:1, you can bet that Reaction B is correct, and if it's about 2:1, you should choose Reaction C.

Keep in mind that your results may not give you exact whole number mole-to-mole ratios because of basic experimental errors. Your results may be off by as much as 20% for this experiment, but you will still be able to choose between the three reactions (A, B, or C) with a fair amount of confidence if you work carefully and collect good data.

Determining the moles of $NaHCO_3$ is easy: Use the measured mass you scoop out of the container to use. (The other reactant, acetic acid, will be used in excess, so its exact amount will have no relationship to the amount of carbon dioxide generated.)

Determining the moles of CO_2 is less straightforward; it's not so simple to collect and measure the mass of a gaseous substance, as you can imagine. In each of the three reactions above, carbon dioxide is the only gas, and all other reactants and product are liquids, solids, or aqueous. As the reaction occurs, carbon dioxide will bubble out of the reaction solution and escape into the laboratory.

Therefore, the mass of your reaction mixture after the reaction will be lighter due to the loss of carbon dioxide, and a simple subtraction tells you how much carbon dioxide was produced:

$$\textbf{Mass of } CO_2 = \left(\begin{array}{c}\textbf{mass of reaction mixture}\\ \textbf{before reaction}\end{array}\right) - \left(\begin{array}{c}\textbf{mass of reaction mixture}\\ \textbf{after reaction}\end{array}\right)$$

One small complication is that some of the CO_2 produced will remain dissolved in the reaction mixture because carbon dioxide is somewhat soluble in water. This means that the mass you calculate by subtraction in the above equation is somewhat too low—i.e., you have not accounted for the carbon dioxide that goes into the water. You will account for this with a correction factor in your calculations.

Once you have chosen the correct reaction between acetic acid and sodium bicarbonate, you can use it to measure the amount of sodium bicarbonate in Alka-Seltzer tablets using a similar methodology. In this case, the mass of sodium bicarbonate will be an unknown. You can measure the amount of CO_2 produced as you did before, and use the mole-to-mole ratio of the chosen reaction to calculate the number of moles and the mass of sodium bicarbonate which reacted. Finally, you will determine your experimental error by comparing your experimentally determined mass of sodium bicarbonate present with what the manufacturer reported on the package of Alka-Seltzer.

OBJECTIVES

In this experiment, you will

- Determine the stoichiometry of a reaction experimentally.
- Weigh by difference a reaction mixture before and after the reaction in order to find the mass of a gas produced.
- Practice molar mass and mole ratio calculations.
- Calculate a percent error and determine how an inaccuracy in a specific measurement affects the outcome.

HAZARDS

Even though the reagents in this lab are fairly safe, please wear safety goggles and dispose of waste in the labeled waste container.

PROCEDURE

PART A: REACTION STOICHIOMETRY

Pour a small amount of sodium bicarbonate into a small beaker to use as your personal supply for the experiment. You can always get more if needed. Clean up any sodium bicarbonate spills with paper towels and water, and dispose of excess down the drain with water.

Run the reactions in a clean and dry 250-mL beaker, with a large watch glass as a cover. Tare a watch glass, measure about 2 g of sodium bicarbonate ($NaHCO_3$) and record the mass in the data table. Measure 50 mL of acetic acid into a graduated cylinder and pour this into a 250-mL beaker.Stack the watch glass on top of the beaker (with the sodium bicarbonate still in the watch glass). Re-zero a balance and carefully measure the mass of the whole stack altogether: beaker with acetic acid, plus watch glass and sodium bicarbonate on top. Record the mass in the data table.

Without spilling any of the sodium bicarbonate or reaction mixture, return to your work area, and initiate the reaction by carefully transferring all of the sodium bicarbonate from the watch glass into the acetic acid. (Note that the solution may overflow if you dump the sodium bicarbonate too quickly.) Put the watch glass back on top of the beaker and swirl the reaction mixture gently. After a few minutes, when the reaction is complete and no more bubbling is observed, weigh and record the total mass of the beaker, reaction mixture, and empty watch glass.

Pour the reaction mixture down the sink with plenty of water. Rinse the glassware thoroughly with water (no soap required) and dry completely. Never put a wet beaker onto a balance!

Repeat for at least one more trial. Do all the calculations for Part A. **Get approval from your instructor before you proceed.**

PART B: ANALYSIS OF AN ANTACID

Using your conclusion about which of the three reactions are correct from Part A, you can now measure the amount of sodium bicarbonate contained in Alka-Seltzer tablets.

Obtain one pack of two tablets from the lab staff. This will allow you to conduct two trials of the reaction.

Use the same procedure as you used in Part A, replacing the sodium bicarbonate with the Alka-Seltzer tablet each time. Give the reaction at least 5 minutes to take place fully, and don't worry if there are small amounts of white solid in the reaction mixture after the reaction is complete.

REPORT SHEETS
Experiment 8: Stoichiometric Analysis of an Antacid

Name: ______________________________ Section: __________

Lab Partner: ______________________________

PART A: REACTION STOICHIOMETRY

	Trial 1	Trial 2	Average
Data			
Mass of $NaHCO_3$			—
Mass of Beaker + Acetic Acid + $NaHCO_3$ + Watch Glass before Reaction			—
Mass of Beaker + Reaction Mixture + Watch Glass after Reaction			—
Calculated Results (see below for guidance)			
Mass of CO_2 Gas Released			—
Moles of CO_2 Gas Released			—
Moles of CO_2, Corrected for Amount Dissolved* (see below for note)			
Moles of $NaHCO_3$ Used			

NOTE: *Calculate by adding 0.0040 moles to the "Moles of CO_2 Gas Released." This correction accounts for the amount of CO_2 that dissolves in 50 mL of aqueous solution.

Calculate (and show your work for at least Trial #1) for the following:

1. Calculate the mass of CO_2 released and then convert it to moles of CO_2.

2. Calculate the number of moles of CO_2, corrected, by following the special note.

3. Using the mass of $NaHCO_3$, calculate the number of moles of $NaHCO_3$.

4. Compare the calculated number of average moles of CO_2 and $NaHCO_3$ in the table above by calculating a ratio. Which of the three possible reactions, A, B, or C (refer to Introduction section for the options), is best supported by these results? Write out the complete balanced chemical equation you chose, and **explain** your reasoning.

PART B: ANALYSIS OF AN ANTACID

	Tablet 1	Tablet 2
Data		
Mass of Beaker + Acetic Acid + Tablet + Watch Glass before Reaction		
Mass of Beaker + Reaction Mixture + Watch Glass after Reaction		
Calculated Results		
Mass of CO_2 Gas Released		
Moles of CO_2 Gas Released		
Moles of CO_2, Corrected for Amount Dissolved* (see note below)		
Moles of $NaHCO_3$** (see note below)		
Mass of $NaHCO_3$		
Average Mass of $NaHCO_3$ for 2 Tablets		

*Calculate by adding 0.0040 moles to the "Moles of CO_2 Gas Released."

**Use the balanced chemical equation from Part A, Question 4, for the mole-to-mole ratio to convert moles of CO_2 (corrected) to moles of $NaHCO_3$.

Calculate the following (and show your work for at least Trial #1):

1. Calculate the mass of CO_2 gas released by the tablet and convert it to moles of CO_2.

2. Calculate the number of moles of $CO_{2,}$ corrected, released by the tablet by following the special note*.

3. Using the mole ratio of CO_2 to $NaHCO_3$ in the balanced chemical equation that you determined in Part A, calculate the moles of $NaHCO_3$ present in the tablet.

4. Convert the moles of $NaHCO_3$ in the tablet to mass of $NaHCO_3$ in the tablet.

5. Calculate for the average mass of $NaHCO_3$ for the two tablets.

6. The mass of sodium bicarbonate in each tablet of Alka-Seltzer is reported as 1916 mg on the package (which we will call the "actual value").

Using your experimental value for the average mass of $NaHCO_3$, calculate the percent error. **Show your work.**

$$\textbf{\% error} = \frac{\textbf{(actual value – experimental value)}}{\textbf{actual value}} \times 100\ \%$$

7. In Part B, suppose the tablet was mostly dissolved and had mostly reacted when some of your solution splashed out of the beaker.
   a. How would this affect the perceived mass of CO_2? Would it be artificially high or artificially low? **Explain**.

   b. Would your final calculated mass of sodium bicarbonate in the tablet be ***artificially high*** or ***artificially low*** as a result of this splashing? **Explain.**

LAB EXERCISE
Experiment 8: Stoichiometric Analysis of an Antacid

Name: ____ Section: __________

1. When acetic acid and sodium bicarbonate react, why does total mass decrease?

2. Suppose you are running a trial of the reaction in Part A. You use 2.80 g of $NaHCO_3$, and determine that 2.83 g of carbon dioxide are produced. Using these data and a periodic table, show calculations to:

   a. Convert the mass of $NaHCO_3$ to moles using the molar mass of $NaHCO_3$.

   b. Convert the mass of CO_2 to moles using the molar mass of CO_2.

   c. Find the simplest ratio between the number of moles of $NaHCO_3$ and CO_2. **(HINT: Divide both numbers by the smaller of the two to get a ratio.)**

   The mole ratio of $NaHCO_3$: CO_2 is ____:____.

   d. Compare this ratio to the ratio of the coefficients for $NaHCO_3$ and CO_2 in the three balanced chemical equations (given on page 1). Which of the three possible reactions discussed is consistent with these results?

   Write the equation here.

   NOTE: The result in part C is not necessarily the correct answer. It is a hypothetical situation for practicing the calculations required in this lab. Do not use this actual result in the lab!

EXPERIMENT 9
TITRATION OF VINEGAR

OBJECTIVES

- Determine the concentration of acetic acid in vinegar (mass percent and molarity).
- Understand how to perform a titration.

INTRODUCTION

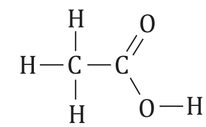
Acid base neutralization is a reaction between an acid and a base to produce a salt and water. An example of this type of reaction is the reaction of the strong acid HCl with the strong base NaOH to produce the salt sodium chloride and water:

$$HCl\ (aq) + NaOH\ (aq) \longrightarrow NaCl\ (aq) + H_2O\ (l)$$

The same result occurs when a strong acid reacts with a weak base, or when a weak acid reacts with a strong base. In each case the acid and base react to form a salt and water, neutralizing the acid and base. The neutralization reaction provides us with one method for determining the amount of either the acid or the base in a solution.

The general method employed for determining the amount of an unknown substance in a solution by using another solution with a known concentration and the reaction stoichiometry from the balanced chemical reaction is known as titration. Titration is a procedure for determining the amount of substance A by adding a carefully measured volume of a solution B with a known concentration until the reaction between them is just complete. Titrations are not used exclusively with acids and bases but anytime you need to determine the concentration of a species in a solution.

In this lab, we will titrate a vinegar solution with a solution of sodium hydroxide, NaOH. The active ingredient in vinegar is the weak acid, acetic acid, CH_3COOH or $HC_2H_3O_2$. The structure is shown here (lone pairs of electrons are not shown), the hydrogen that is attached to oxygen is the acidic hydrogen, meaning it will be the only hydrogen involved is this acid-base reaction. The abbreviation for acetic acid is HAc.

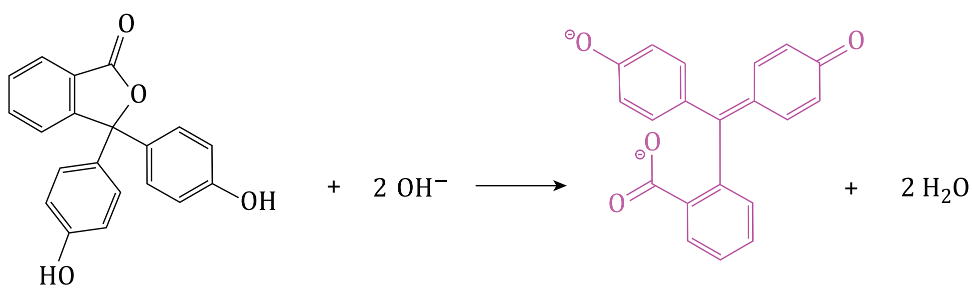


The sodium hydroxide will neutralize the acetic acid, HAc, in the vinegar. The Ac actually represents the acetate anion, CH_3COO^- or $C_2H_3O_2^-$.

$$HAc\ (aq) + NaOH\ (aq) \rightarrow NaAc\ (aq) + H_2O\ (l)$$

The point at which all the acetic acid has reacted with the sodium hydroxide is called the **equivalence point.** Once the neutralization reaction is complete, any additional base added will produce a basic solution and the pH will greatly increase. We can determine when the titration is complete by employing an acid-base **indicator** that will change color in a basic solution.

Phenolphthalein is an acid-base indicator that changes color with changes in pH. This particular indicator **is colorless in acidic solutions but turns pink in basic solutions.** Phenolphthalein changes color at pH = 8.2. If one extra drop of base is added after the acid has been neutralized, the pH will jump significantly as the solution becomes basic and it will change to a pink color.



Colorless   Pink

As soon as very faint pink color persists for 30 seconds, you have reached the **endpoint.** In this case the equivalence point is very close to the endpoint; therefore, the endpoint can be used to approximate the equivalence point.

At the equivalence point, the number of moles acid will have exactly reacted with the number of moles of base according to the mole ratio in the balanced equation.

$$\text{Number of moles of } OH^- = \text{Number of moles of } H^+$$

We can determine the molarity of the acid by titrating a known volume of the acid solution with a known concentration and volume of a base.

The concentration of the known solution is determined through a process called **standardization.** For this laboratory the standardization of NaOH has already been performed for you.

INFORMATION ABOUT STANDARDIZATION (FOR YOUR FUTURE INFORMATION)

Hydroxides are very hydroscopic (absorb water from the atmosphere) making them difficult to accurately weigh out on a balance. Hydroxide solutions can absorb carbon dioxide from the atmosphere, which can lead to changes to the concentration of the solution via acid-base reactions. And, although simple measuring is a good approximation of concentration, it is not precise enough for use to determine the concentration of a solution for use in a titration. Instead, the concentration of the known solution (in this case the NaOH solution) can be determined through reaction with an appropriate primary standard.

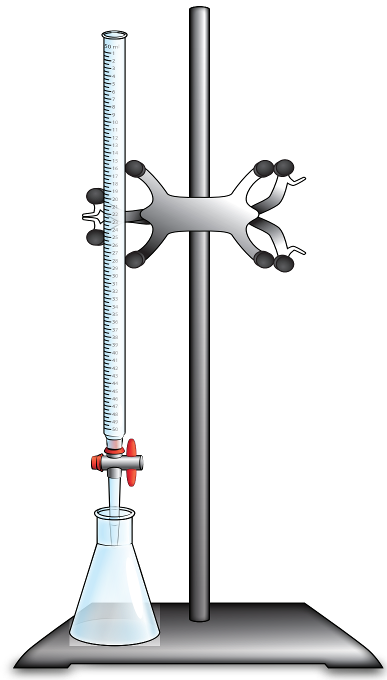
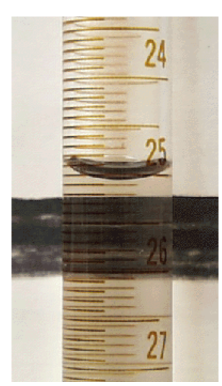


Figure 1. Equipment set-up for titration

LABORATORY TECHNIQUE FOR BURETS

Burets are used to deliver a recorded amount of liquid or solution to another container. A buret is marked in milliliters like a graduated cylinder, but buret markings show **0 mL** at the top, and the numbers increase as you go down the buret. The stopcock controls the liquid flow. It is ***open*** when ***parallel*** to the length of the buret and ***closed*** when ***perpendicular*** to the length of the buret.

- **Washing and rinsing the buret:** To clean a buret, wash its interior with soap and tap water. Next, rinse the buret with 5–10 mL portions of DI water. With the buret over the sink and the stopcock open, pour the water into the buret and let it drain out the tip. Use a beaker to pour solutions into the buret—most breakage occurs during washing, and burets do ***NOT*** fit under the faucet.



- **Conditioning the buret:** After the buret is well-drained, close the stopcock and add about 5 mL of the ***titrant*** (the solution to be used into the buret). Tilt the buret sideways and roll the barrel to completely rinse the inner walls of the buret. Drain the solution through the buret tip to insure the tip is also conditioned. Repeat this step at least twice to be sure all interior surfaces are rinsed with titrant.

- **Filling the buret:** Close the stopcock. Use a clean funnel to fill the buret with titrant just above the "0" mark. Place a container under the buret tip, and open the stopcock briefly to fill the buret tip with solution, leaving no air bubbles, and to get the level of meniscus to fall within the markings of the buret. If the tip does not fill with solution when the stopcock is in the open position, there may be an air bubble in the stopcock. Consult your instructor. **NOTE: The initial level of titrant need not be exactly at 0.00 mL** as the initial level of liquid will be recorded and subtracted from the final volume to determine the volume delivered after a titration is complete.
- Reading the buret: Always remove the funnel used to fill the buret before taking any measurements. Record the volume of titrant by noting the bottom of the meniscus. On the buret shown below, numbers marked for every 1 mL, and the ten lines between each number represent every 0.1 mL. Thus, the level of titrant in the buret can be estimated to one more decimal place than the markings or to the nearest 0.01 mL. Thus, in the figure to the right, the meniscus is about halfway between 25.0 and 25.1 mL, so the level of titrant can be recorded as ***25.04 mL, 25.05 mL, or 25.06 mL*** depending on whether the bottom of the meniscus appears to be just above, just at, or just below halfway, respectively.
- **Cleaning the buret:** Afterwards, empty the buret, disposing of the titrant according to the waste disposal instructions for each experiments. Wash the buret with soap and tap water, then rinse with several portions of tap water, allowing some tap water to run through the tip. Do a final rinse with small portions of DI water, allowing the DI water to run through the tip, then return the buret to the stockroom.

SAFETY PRECAUTIONS

CAUTION: Harmful Chemicals

NaOH is caustic. Handle with care. In case of contact with skin, rinse the area with large amounts of water and notify your instructor. Wear goggles at all times in the chemistry laboratory. You may wear gloves for extra protection.

PROCEDURE

1. Record the molarity of NaOH and obtain about 50 mL of this NaOH solution. (Make sure your goggles are on!) You will use a small amount of NaOH (about 15 mL) for the next step, and the rest for the titration.
2. Prepare a buret for titration as described under "Laboratory Techniques for Burets" on the previous page. Fill the buret with the sodium hydroxide solution and follow the usual procedures for eliminating air bubbles and setting the initial level.
3. Use a graduated pipet to measure out 2.00 mL of vinegar. Record the actual volume used. Transfer the vinegar to a clean Erlenmeyer flask (the 250-mL size should be large enough). Add about 25 mL of distilled water. (*Does adding water change the number of moles of acid in your sample? No, it just dilutes it to lower concentration. You recorded the initial volume of vinegar used and you will use that volume to calculate the concentration.*)
4. Add 2–4 drops of phenolphthalein indicator to the flask.
5. Record your starting buret reading. This is your initial volume reading, V_{initial}. (*Have your lab partner verify your measurement—use the proper number of significant figures! A buret has an uncertainty of ± 0.01 mL.*)
6. Using the buret, add NaOH dropwise to the vinegar sample with constant swirling of the solution (analyte flask). The endpoint may be easier to see if you have a white sheet of paper under the flask. Stop when a last drop (or partial drop) of hydroxide solution results in a faint pink color that persists for at least 30 seconds.
7. Record your ending buret reading. This is your final volume reading, V_{final}. Calculate the volume of NaOH delivered (added to flask) by finding the difference as $V_{\text{final}} - V_{\text{initial}}$.
8. Pour the contents of the flask into the chemical waste container. Rinse the flask with tap water, then a couple of times with distilled water. Use the volume of NaOH delivered for the first trial to estimate if you have enough NaOH in your buret for two more trials; add more NaOH to the buret if necessary. Repeat the titration twice more with two new samples of vinegar by repeating Steps 3 through 7. The volumes delivered for each of the trials in a titration need to be within **± 0.20 mL.**

CHEMICAL WASTE

Everything should go into the appropriate labeled waste container in the fume hood.

MATERIALS AND EQUIPMENT

Ceramic stand, buret, buret clamp, 250-mL Erlenmeyer flask, 5- or 10 mL-graduated pipet, green pipet pump, 25-mL graduated cylinder, a large beaker to hold waste and rinse solutions, two small beakers to hold NaOH and vinegar.

REPORT SHEETS
Experiment 9: Titration of Vinegar

Name: ______________________________ Section: __________

Lab Partner: ______________________________

DATA

NOTE: The volumes delivered for each of the trials in a titration should be within **± 0.20 mL.**

Record your data using appropriate number of significant figures and units.

Table 1. Acid-Base Titration

	Trial 1	Trial 2	Trial 3
Volume of Vinegar (mL)			
Molarity of NaOH			
Initial Volume (mL) of NaOH in Buret			
Final Volume (mL) of NaOH in Buret			
Volume of NaOH Delivered (mL)			

CALCULATIONS

Show all of your work using proper significant figures and units to calculate the results.

(For all of the calculations that follow, show a sample calculation for **at least one trial.**)

1. Calculate the volume of NaOH (in liters) used to neutralize the acetic acid.

2. Using the molarity and volume of NaOH solution (in liters) used, calculate the moles of NaOH used to neutralize the acetic acid.

3. Using the balanced chemical equation for the reaction of NaOH with CH_3COOH, calculate the moles of acetic acid in vinegar.

4. Using the moles of CH_3COOH present and initial volume of vinegar pipetted, calculate the molarity of acetic acid in vinegar.

5. Using the molarity of acetic acid in vinegar for each trial, calculate the average molarity of acetic acid in vinegar.

6. Using the moles of acetic acid and molar mass of acetic acid, calculate the mass of acetic acid used in each titration.

7. Using the initial volume of vinegar pipetted and the known density of vinegar 1.005 g/mL, calculate the mass of vinegar solution.

8. Using the mass of acetic acid and the mass of vinegar solution, calculate the (*m*/*m*) % of acetic acid in vinegar.

9. Calculate the average (*m*/*m*) % of acetic acid in vinegar.

Tabulate your results here:

Table 2. Acid-Base Titration

	Trial 1	Trial 2	Trial 3	Average
Volume of 0.10 M NaOH Delivered (L)				—
Moles of 0.10 M NaOH Used				—
Moles of Acetic Acid in Vinegar				—
Molarity of Acetic Acid				
(*m*/*m*) % of Acetic Acid				

POST-LAB QUESTIONS

1. Most commercial vinegar solutions are 5.0 % (*m*/*m*) acetic acid.
   a. Calculate the percent error for your average % (*m*/*m*) acetic acid compared to a commercial solution of 5.0 % (*m*/*m*) acetic acid.

   b. For this lab, percent errors greater than 10% indicate that the experiment was not very accurate. How accurate were your results? Include your calculated percent error to support your answer.

LAB EXERCISE
Experiment 9: Titration of Vinegar

Name: ______________________________ Section: __________

1. What is the formula for calculating density?

2. What is the formula for calculating the % (*m*/*m*) of a solution?

3. What is the formula for calculating the molarity of a solution?

4. If you were given 250.0 mL of 0.500 M NaCl, which has a density of 1.003 g/mL, what would be the % (*m*/*m*) of the NaCl solution. Show your work for credit. Report your answer to an appropriate number of significant figures and include units.
    a. Using volume and molarity of NaCl, calculate the moles of NaCl present.

    b. Convert moles of NaCl to grams of NaCl.

    c. Calculate the mass of the 250.0 mL of solution using its density 1.003 g/mL.

    d. Calculate the % (*m*/*m*) of the NaCl solution.

EXPERIMENT 10
SPECTROSCOPY

OBJECTIVES

- To identify the unknown element by comparing its flame color to that of known elements.
- To learn about light and the relationship between energy, frequency, and wavelength.
- Observe continuous vs. line spectra resulting from emission from gas discharge tubes and identify the element in fluorescent light bulbs.

INTRODUCTION

Light is a form of energy called electromagnetic radiation. Light travels at a speed of 2.998×10^8 m/s, a constant value in a vacuum, and it is denoted by the lowercase letter “c.” A chart of the electromagnetic spectrum is shown below.

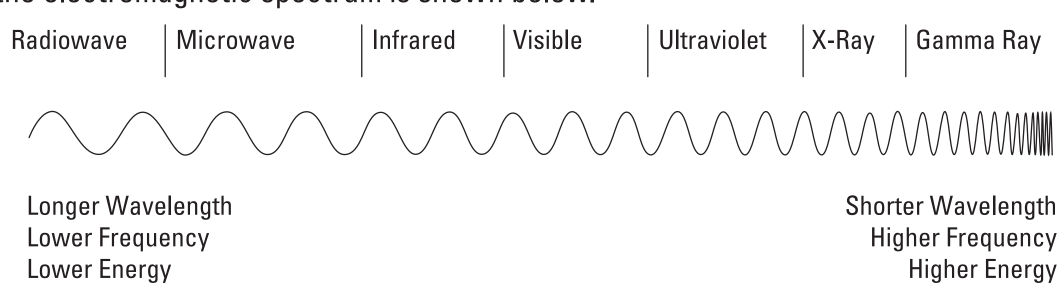


Figure 1. Types of electromagnetic radiation

Which type of electromagnetic radiation is more dangerous to humans, radiowaves, or x-rays? ___________________

Which type of electromagnetic radiation has the highest energy? ___________________

Which type of electromagnetic radiation is associated with the sun? ___________________

What is the relationship of the frequency of a wave to the energy of the electromagnetic radiation? ___________________

One model used to describe electromagnetic radiation is called the **wave model.** This model uses several variables to describe this radiation: wavelength, frequency, and energy.

Wavelength is denoted by the symbol lambda (λ) and is the distance between the crests (peaks) or troughs or one complete wave cycle of a wave. Wavelength can be measured in any length unit. For example, radio waves are often measured in meters. Gamma rays are often measured in nanometers.

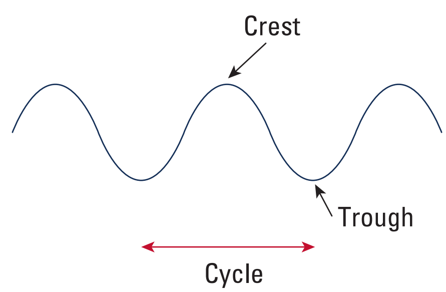


Figure 2. Terminology of a wave.

Another unit for wavelength is the micrometer (10^{-6} meters, or called microns). It's often used in biology since it is the approximate size of cells and bacteria. You will encounter microns when using most microscopes.

The second variable used to describe waves is called **frequency (ν)**, which is the number of cycles which pass per unit time (see Figure 2 above for "cycle"). The unit of frequency is the hertz (Hz). Once per second is 1 hertz, twice per second is 2 hertz, etc. One thousand hertz is expressed as a kilohertz (kHz). One million hertz is a megahertz (MHz).

Frequency is denoted by the symbol nu (ν). Frequency and wavelength of light are related to the speed of light by this equation:

$$c = \lambda\nu$$

Shorter ↓ wavelength = higher ↑ frequency

This is an inverse relationship. (See the chart above to confirm this.)

The **energy** carried by light is related to both wavelength and frequency. X-rays and ultraviolet rays are high energy radiation and can be damaging to your cells. On the other hand, radio waves are lower in energy and you need not use "blocking" agents to protect yourself!

The equation that relates energy and frequency is:

$E = h\nu$

where h is Planck's constant with a value of 6.626×10^{-34} J · s

Higher ↑ frequency = Higher ↑ energy

This is a direct relationship. (See the chart above to confirm this.)

A second model is often used to describe the energy carried by light. This model describes light as having particle characteristics. These "particles" are bundles of energy called photons. Each photon carries a small bundle of energy. A single photon associated with ultraviolet light has greater energy than a single photon associated with radio waves. Using this model, we can think of light as a stream of photons.

Let's compare radiowaves (a) with ultraviolet radiation (b):

Which type of radiation is higher energy?

Looking at **radiowaves**, when you tune your radio, you are receiving radio waves with frequencies between 530 and 1700 kHz for AM and 87.5 to 108 MHz for FM. (This corresponds to the "numbers" on your dial.) A radio wave corresponding to 1700 kilohertz on your AM radio has a wavelength of approximately 200 meters, approximately twice the length of a football field. These waves pass through us without harm.

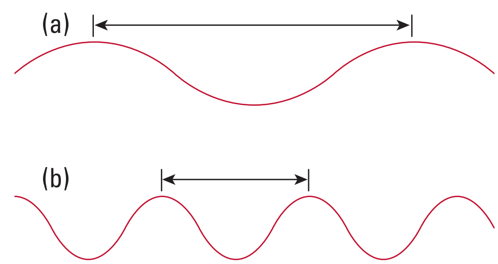


Figure 3. Comparing two waves corresponding to different types of electromagnetic radiation.

Sunscreens protect your skin from **UV radiation** by absorbing wavelengths between 290 and 400 nanometers. This is a short wavelength compared to radio waves. The frequency of UV light compared to radio waves is extremely high. UV light have frequencies in excess of 700 trillion (7×10^{14}) hertz. Thus, photons of UV light have a higher energy than photons of radiowaves.

Both radio waves and ultraviolet light are invisible to the human eye. Colors are due to visible light, which have wavelengths between radiowaves and UV, with a range between 400–700 nanometers. Each color has a corresponding wavelength. For example, red light has a wavelength in the 700 nm range, while violet light has a wavelength near 400 nanometers. The visible spectrum (the "rainbow") can be remembered with a mnemonic ROY G BIV (***R***ed, ***O***range, ***Y***ellow, ***G***reen, ***B***lue, ***I***ndigo, ***V***iolet). An interesting bit of trivia—the British use the following mnemonic—***R***ichard ***O***f ***Y***ork ***G***ave ***B***attle ***I***n ***V***ain.

ORIGIN OF SPECTRAL LINES

The origin of the spectral lines can be explained using the planetary model of the atom, in which the nucleus is in the center and surrounded by orbits (n) on which electrons travel (**NOTE:** This is not completely accurate—this is an oversimplified model). Atoms are usually in a low energy state, which means their electrons are in the lowest energy orbits such as level 1 ($n = 1$). This is called the ground state for an atom. If energy is added to an atom, the electrons can absorb energy and move to a higher energy state, called an excited state (in Figure 4a, the electron jumps from level 1 to the second level, $n = 2$). This energy can be added to atoms by an electric discharge or by heat, an incoming photon. This added energy is emitted when the atom gives off a photon and the excited electron returns to the ground state (see Figure 4b). Figure 5 shows electron transitions focusing on energy levels of electrons.

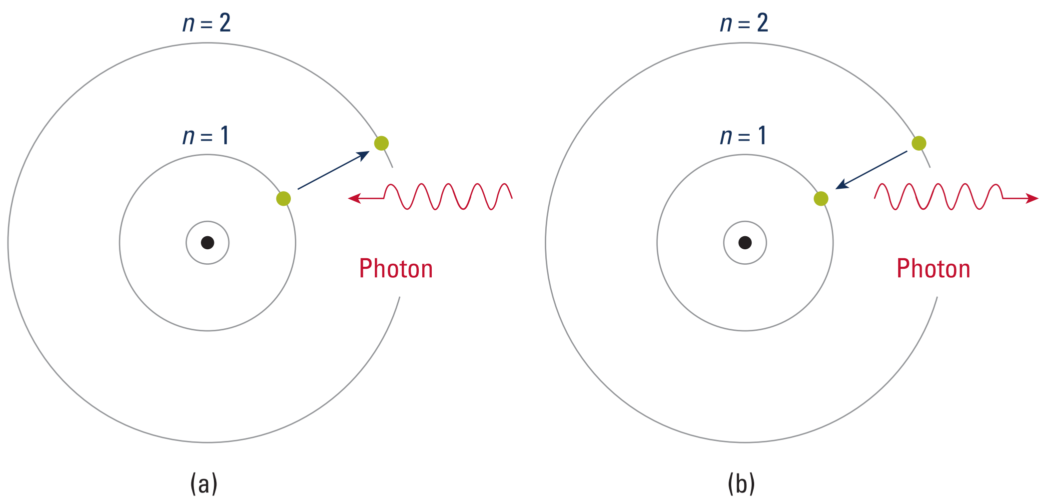


Figure 4. (a) An atom in its ground state, with the electron (green dot) moving from level 1 to level 2 by absorbing energy from an incoming photon. (b) An atom in an excited state emitting a photon as its electron moves to a lower energy level. Only two levels ($n = 1$, $n = 2$) are shown for simplicity.

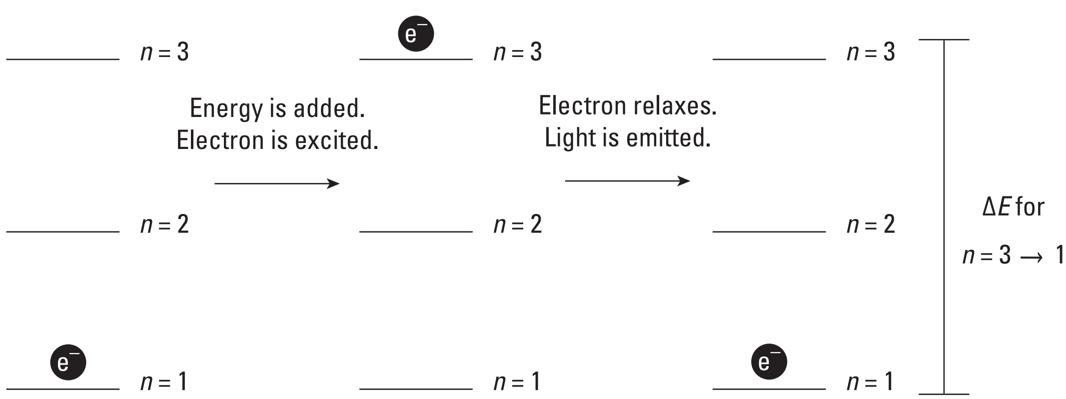


Figure 5. Electron transitions in an atom from the perspective of energy levels.

The light emitted by atoms has a definite wavelength and if visible, a color that depends on the amount of energy originally absorbed. Each excited electron will emit one photon of light as the electron moves to a lower energy state. Billions of excitations and emissions are possible because a typical sample is made up of billions of atoms. There are also many possible energy jumps for each electron. Each of these "jumps," electron transitions (ΔE), and matches a different color of light. These colors constitute the emission spectrum of the excited atoms. A large jump corresponds to a color associated with high energy photons. A small jump corresponds to a color associated with a low energy photon.

Each element has a unique set of spectral lines. Since each element has a unique set of spectral lines, spectral studies can help us identify an element. The line spectrum of elements is much like a "fingerprint" for the element.

SAFETY PRECAUTIONS

Use caution with the Bunsen burners and wooden splints. Wear goggles and tie hair/loose clothing back to prevent fires. Do not touch the connections when the power supply is plugged into an outlet. A serious electric shock can result. Also, these bulbs get very hot. Allow them to cool before touching. Turn off the power supply when you are not using it. Place chemical waste in the labeled waste container.

PROCEDURE

PART 1: FLAME TESTS (ALL OF THE SAMPLES ARE FINELY GROUND POWDERS)

1. Light a Bunsen burner. Adjust the air vent to obtain a blue flame.
2. Obtain a vial containing a known sample (note the element symbol on the vial) and a piece of weigh paper. Using a spatula, take a scoop of solid from the vial onto the weigh paper. Dip a wooden splint it into one of the beakers of water to moisten it. Gently roll the splint onto the weigh paper with the solid so the tip of the splint is covered with sample.
3. Place the splint into the clear blue flame and observe/record the color of the flame. Dip the splint into the other beaker of water to douse the flame.
4. Repeat this procedure for each known.

PART 2: IDENTIFICATION OF AN UNKNOWN SOLID

Repeat the above procedure for two unknowns. Check your results with your instructor before you proceed.

PART 3: CONTINUOUS AND LINE SPECTRA

Using a spectroscope, examine the spectrum given off by an incandescent light bulb and a fluorescent light bulb. Sketch the spectrum for each labeling of your diagram with colors and wavelengths.

PART 4: OBSERVING LINE SPECTRA WITH THE SPECTROSCOPE

1. Your instructor has set up several gas discharge tubes in power supplies. Use your spectroscope to examine the spectrum for each of the tubes. **(See Hazards note.)**
2. Sketch the results on your data sheet, noting the color and wavelength of the observed lines. Look at tubes containing He, Ne, Hg and H_2 (and others).
3. Use your spectroscope to re-examine the fluorescent light bulbs in the lab room.

REPORT SHEETS
Experiment 10: Spectroscopy

Name: ______________________________ Section: __________

Lab Partner: ______________________________

PART 1: FLAME TESTS FOR KNOWN ELEMENTS

Write the color of the flame next to each element.

Lithium ____________________ Calcium ____________________

Sodium ____________________ Strontium ____________________

Copper ____________________ Potassium ____________________

Iron ____________________ Barium ____________________

PART 2: IDENTIFICATION OF SOLID UNKNOWNS

Solid unknown # ________ Color of flame ________ Element ____________

Solid unknown # ________ Color of flame ________ Element ____________

PART 3: CONTINUOUS AND LINE SPECTRA

(Draw what you see through the spectroscope).

Incandescent light bulb:

Fluorescent light bulb:

PART 4: GAS-DISCHARGE TUBES

Draw what you see through the spectroscope when you examine each of the following gas discharge tubes. Include colors and wavelengths. Also, describe the color of the glowing discharge tube as seen by naked eye (without the spectroscope).

Helium

Color of the glowing discharge tube: ____________________

Helium spectrum

Neon

Color of the glowing discharge tube: ____________________

Neon spectrum

Mercury

Color of the glowing discharge tube: ____________________

Mercury spectrum

Hydrogen

Color of the glowing discharge tube: ____________________

Hydrogen spectrum

POST-LAB QUESTIONS

1. What are some differences in the emission spectra of the different elements?

2. In which discharge tubes are electrons undergoing the largest and smallest energy changes (from what we can tell by only observing gas tubes with the naked eye)?

3. a. Which discharge tube emits *visible* light with the highest frequency, as seen with the naked eye?

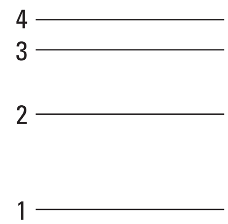
   b. Which discharge tube emits *visible* light with the longest wavelength, as seen with the naked eye?

4. What are some differences in emission between the incandescent and fluorescent light bulb?

5. Identify the light sources that produce a line spectrum. Identify those that produce a continuous spectrum.

6. Which element in the glowing discharge tubes is most likely to be the element in the fluorescent lamp? Explain your reasoning.

7. To answer a–d, consider the energy level diagram to the right.



   a. Draw arrows on the diagram showing all of the transitions an electron could make that would result in the emission of a photon.

   b. How many emission lines could we observe from this atom?

   c. Which transition would emit light with the highest frequency (shortest wavelength)?

      From n = ______ to n = ______

   d. Which transition would emit light with the lowest frequency (longest wavelength)?

      From n = ______ to n = ______

LAB EXERCISE
Experiment 10: Spectroscopy

Name: ____ Section: __________

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

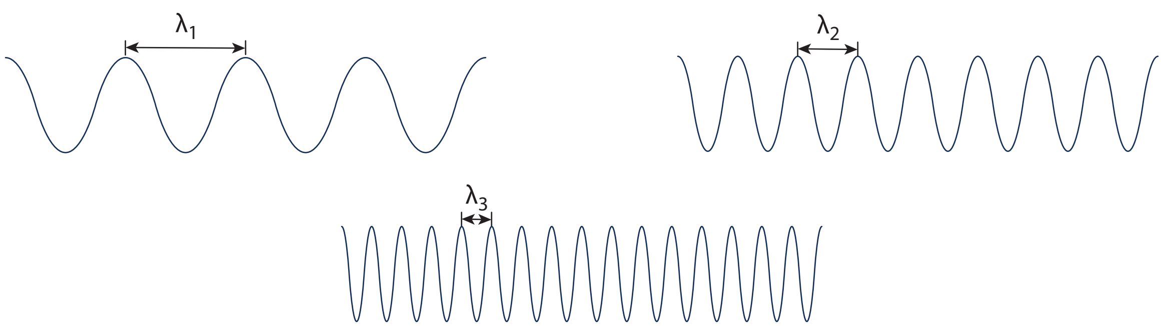
1. a. List the colors in the visible spectrum, from highest to the lowest frequency.

   b. Which color of visible light has the lowest energy? ______________________

   The longest wavelength? ______________________________________

   The lowest frequency? ______________________________________

2. The following diagrams of waves are labeled λ_1, λ_2, and λ_3. Fill in each blank with the best choice.



   a. Which wave has the longest wavelength? _____ The shortest wavelength? _____

   b. Which wave has the greatest frequency? _____ The smallest frequency? _____

   c. Which wave would be a good model for UV light? _____ Radio waves? _____ Visible light? _____

   d. Which wave has photons with the greatest energy? _____
   The smallest energy? _____

3. What is the difference between a continuous spectrum versus a line spectrum?

EXPERIMENT 11
MOLECULAR MODEL BUILDING

PART A

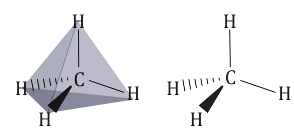
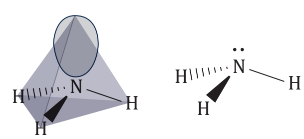
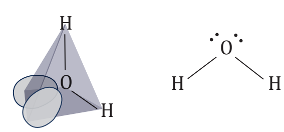
The first set of molecules we will examine contains only two atoms. These are considered to have a linear molecular shape.

PART B

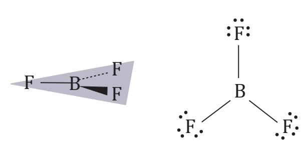
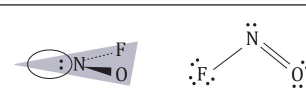
1. The molecules in this part contain more than 2 atoms. Draw the Lewis structure for each of them.
2. Make a model of the molecule. Use the short gray plastic connecting rods for single bonds and lone pairs of electrons.
   - Use the longer flexible connecting rods for double and triple bonds (2 for a double bond; 3 for a triple bond).
   - Use the following colored spheres:
     - C, Si = black
     - N, P, As = blue
     - H = white
     - O, S = red
     - Cl, Br, I = purple or orange
   - If the "red" sphere is the central atom, you won't be able to see the lone pairs of electrons; use a black sphere in these cases.

3. The electron geometry is determined by the number of electron groups on the central atom. (**HINT:** Look at the arrangement of the connecting rods.)
   - 4 groups result in a tetrahedral arrangement of the electron groups.
   - 3 groups result in a trigonal planar arrangement of the electron groups.
   - 2 groups result in a linear arrangement of the electron groups.
4. Determine the molecular shape of the molecule. While the lone pairs of electrons influence the molecular shape, the arrangement of the ***atoms*** is used to describe the molecular shape. ***Do not use any abbreviations in your answers!***

POSSIBLE SHAPES FOR MOLECULES WITH *4 ELECTRON GROUPS* AROUND THE CENTRAL ATOM (BOND ANGLES NEAR 109.5°)

Molecular Shape	3-D Sketch and Lewis Structure
Tetrahedral A four sided figure. Each side is an equilateral triangle.	
Trigonal pyramidal A four sided figure. Three sides are isosceles triangles. They are sitting on an equilateral triangular base.	
Bent A nonlinear arrangement of atoms.	

POSSIBLE SHAPES FOR MOLECULES WITH *3 ELECTRON GROUPS* AROUND THE CENTRAL ATOM (BOND ANGLES NEAR 120°)

Molecular Shape	3-D Sketch and Lewis Structure
Trigonal Planar A flat molecule. The atoms bonded to the central atom form an equilateral triangle. The central atom sits in the center of the triangle.	
Bent A nonlinear arrangement of atoms.	

REPORT SHEETS
Experiment 11: Molecular Model Building

Name: ______________________________ Section: __________

Lab Partner: ______________________________

PART A

Species/ Total v.e.	Draw the Lewis Dot Structure	Name of Molecular Shape	Name of Element or Compound
H_2 2 v.e.			
Br_2			
HBr			
N_2			
CO			

CONCLUSIONS

If only two atoms are bonded, the molecular shape will always be ____________________.

PART II

Species/ Total v.e.	Lewis Dot Structure	# of Electron Groups	Bond Angles	3-D Sketch	Name of Electron Geometry	Name of Molecular Shape	Name of Compound
CF_4 32 v.e.							
CHF_3							
OF_2							
H_2S							

Species/ Total v.e.	Lewis Dot Structure	# of Electron Groups	Bond Angles	3-D Sketch	Name of Electron Geometry	Name of Molecular Shape	Name of polyatomic ion
NO_2^-							
CO_3^{2-}							
NH_4^+							

This page contains polyatomic ions. They have gained or lost valence electrons to become charged.

*Nitrite (NO_2^-) and carbonate (CO_3^{2-}) have resonance structures—different but equivalent Lewis structures.

EXPERIMENT 12
SEPARATION OF A TERNARY MIXTURE

SUGGESTED PRE-LAB ACTIVITY

This lab has many parts to it. It is advised to sketch out the procedure to this lab as a pre-lab activity to see how the components of this mixture are separated. A flow chart or a schematic diagram showing the steps in picture format may be helpful.

OBJECTIVES

- Understand the differences between physical and chemical separations.
- Identify pure substances and mixtures.
- Distinguish homogenous mixtures from heterogeneous mixtures.

INTRODUCTION

Mixtures are physical combinations of two or more different substances where each of the substances are mixed but do not combine with each other chemically. For example, in muddy water (a mixture of dirt and water) both the dirt and the water maintain their chemical identity. Mixtures can be classified as either *homogeneous* or *heterogeneous*. Homogeneous mixtures have the same composition throughout the mixture and appear uniform. Homogeneous mixtures are often *solutions*, a situation where one substance (the solute) is dissolved in another (the solvent). Dissolving is a physical process and does not change the substance chemically. Vinegar and household ammonia are examples of homogeneous mixtures. In contrast, heterogeneous mixtures have varying composition throughout the mixture and do not appear uniform. Chocolate chip cookie dough, seashells on a beach and chicken noodle soup are all examples of heterogeneous mixtures.

Chemists often need to separate a mixture into its pure components. Chemical and physical properties can be used to separate some mixtures. **Solubility** is an example of a physical property used to separate mixtures. When water is added to a mixture of salt and sand, the salt dissolves but the sand does not. The mixture can then be separated by **filtration.** The dissolved salt passes through the filter paper with the water (the liquid or soluble mixture that passes through the filter paper is called **filtrate**) but the insoluble sand remains behind on the filter paper (the solid or insoluble substance left on filter paper is called **residue**). Neither the salt nor the sand has been chemically changed; the separation is a physical process.

In this laboratory, you will be using both physical and chemical properties to separate the components of a ternary (three substances) mixture. You will also determine the mass percentage of each substance in the mixture.

When distilled water is added to a mixture of NaCl (sodium chloride), $CaCO_3$ (calcium carbonate) and SiO_2 (silicon dioxide = sand), the water dissolves the NaCl (a physical change), leaving the insoluble $CaCO_3$ and SiO_2 behind. Filtration separates the solids (residue) from the NaCl aqueous solution (filtrate).

$$\textbf{Beaker, NaCl, } \mathbf{CaCO_{3,}\ SiO_2} \xrightarrow{\text{water}} \textbf{Beaker, NaCl, } \mathbf{CaCO_{3,}\ SiO_2}$$

Water is removed from the NaCl aqueous solution (filtrate) by **evaporation.** The mass lost after the filtrate is dried is the mass of NaCl in the mixture. From the residue (containing $CaCO_3$ and SiO_2) left on the filter paper, $CaCO_3$ is removed by adding HCl (hydrochloric acid) through a chemical change).

$$\textbf{Evaporating Dish, } \mathbf{CaCO_{3,}\ SiO_2} \xrightarrow{\text{acid}} \textbf{Evaporating Dish, } \mathbf{SiO_2}$$

The HCl reacts with the $CaCO_3$ to produce water, carbon dioxide (a gas), and $CaCl_2$ (which is dissolved in the water). The $CaCO_3$ is chemically converted into two new substances CO_2 and $CaCl_2$ according to this chemical equation[1]:

$$\mathbf{CaCO_3\,(\textit{s}) + 2\,HCl\,(\textit{aq}) \rightarrow CO_2\,(\textit{g}) + CaCl_2\,(\textit{aq}) + H_2O\,(\textit{l})}$$

The $CaCl_2$ solution is removed by **decantation**, which means removing the aqueous layer by carefully pouring it off. The mass lost after the residue has been dried is the mass of the $CaCO_3$ that was in the mixture before its chemical transformation. Since the sand does not dissolve in water or react with the acid, its mass can be determined by weighing the material left in the evaporating dish. The mass of the sand can then be calculated by subtracting

1 The chemical reaction states that 1 unit of solid $CaCO_3$ combines with 2 units of HCl acid in water solution to produce 1 unit of CO_2 gas, 1 unit of $CaCl_2$ in solution and 1 unit of water. Note that this representation conserves the number and kind of atoms found on either side of this *chemical equation.*

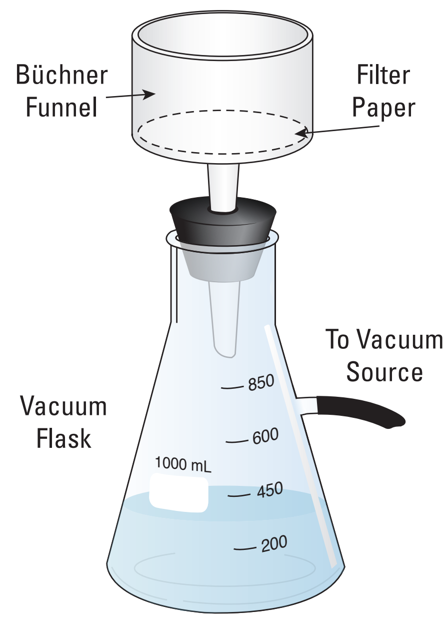
the mass of the evaporating dish from the mass of the evaporating dish and sand combined. Because the unknown contained only $CaCO_3$, SiO_2, and NaCl, the mass of $CaCO_3$ can be determined by subtracting the isolated masses of both the sand and NaCl from the mass of the total unknown.

The mass percent of each component is determined by dividing each of the isolated components (SiO_2, NaCl, and $CaCO_3$) by the mass of the initial mixture and multiplying by 100%:

$$\textit{\% component} = \frac{\text{mass of the component}}{\text{mass of the mixture}} \times 100\%$$

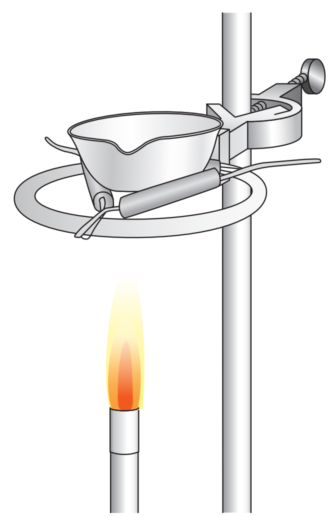
PROCEDURE

1. Obtain a 150-mL beaker.
2. Obtain an unknown mixture and accurately weigh (±0.001 g) out between 2.5 g and 3.5 g into the beaker. Record the actual mass used in your data table.
3. Add 10 mL distilled water to the mixture and stir. This should dissolve the NaCl but leave the $CaCO_3$ and SiO_2 intact.
4. Prepare a vacuum flask with a Büchner funnel (see diagram). Obtain a 7.5-cm circular filter paper that fits inside. Connect the flask to the vacuum line using a rubber hose. Clamp the flask to a ring stand so it does not fall over.
5. Carefully filter the mixture by slowly pouring through the filter paper with the vacuum. Be careful not to lose any of the filtrate (NaCl (*aq*), the salt water solution).



6. Rinse the original beaker with approx. 5 mL of water and pour over the filter paper (repeat two times). This should ensure that all of the remaining residue ($CaCO_3$ and SiO_2) is on the filter paper and that all of the NaCl is in the liquid portion (filtrate). Retain the residue for use in Step 11.
7. Weigh a clean, dry 250-mL beaker and a clean, dry watch glass and record the actual masses on your Report Sheet.
8. Pour the filtrate (the liquid in your suction flask) into the beaker. Be careful not to lose any solution. Use a small amount of water from your distilled water bottle to rinse the flask contents into the beaker.

9. Place the beaker containing the solution of NaCl on a hot plate and gently boil to remove the water. Once the volume has been reduced, you will want to place a watch glass over the beaker to prevent splattering (convex side down). Once NaCl is fully dried, wait for the beaker and watch glass to cool down to room temperature and then record their mass.
10. Weigh (±0.001 g) a clean, dry evaporating dish and record the mass on your data sheet.
11. Using forceps (tweezers), gently transfer the filter paper with residue($CaCO_3$ and SiO_2) from step 6 to the evaporating dish. Be careful not to lose any residue.
12. Obtain approximately 15 mL of 2.0 M HCl in a graduated cylinder. **Be careful: HCl is a strong acid.** Add HCl dropwise to the residue and filter paper until you no longer observe the evolution of gas bubbles. The HCl will react with the $CaCO_3$ but not the SiO_2. You will probably not need to use all of the HCl. (Excess goes in the waste beaker.) When the reaction is completed, you will have $CaCl_2$ solution and SiO_2 in the evaporating dish.
13. Use a 250-mL beaker as a waste beaker. Carefully decant the solution in to your waste beaker, making sure not to lose any of the solid SiO_2.
14. Using forceps (tweezers), pick up the filter paper and gently rinse it with distilled water from the wash bottle retaining the liquid in the evaporating dish.
15. Discard the filter paper in the appropriate waste container.
16. **Decant** (carefully pour off) the rinse water from the SiO_2 in your waste beaker.
17. Rinse the SiO_2 a third time and decant the solution into the waste beaker making sure not to lose any of the SiO_2.
18. Place the evaporating dish onto a clay triangle and/or iron ring clamped on your ring stand and gently heat the dish carefully over a cool (low) flame for 10 minutes until the SiO_2 is dry. If the SiO_2 does not appear dry, keep heating. It is very important to use a cool flame to ensure that the SiO_2 does not spatter.



19. Allow the evaporating dish to cool, then weigh the evaporating dish and SiO_2.
20. Heat the evaporating dish and SiO_2 again for two minutes, let it cool, and reweigh it. If the masses differ by more than 0.002 g, heat the dish for another two minutes followed by cooling, and weighing. Remember, the last weighing is the number you will use in your calculations.

WASTE DISPOSAL

Place all waste for this experiment in a labeled waste container.

SAFETY

You must wear safety goggles at all times during this experiment.

It is recommended you **wear gloves** when handling acids: 2.0 M HCl can cause chemical burns. If you spill any on you, wash it immediately with plenty of water and notify your instructor.

Several heating steps are involved. Handle the hot glassware and evaporating dishes with tongs or oven mitts. Make sure to let the hot beaker and hot evaporating dish cool to room temperature before weighing.

MATERIALS

- Beakers (150 mL, 250 mL)
- Watch glass
- Ring stand
- Büchner funnel
- Filtration flask
- Vacuum hose
- Filter paper (7.5 cm)
- Clamps
- Evaporating dish
- Hot plate
- Tweezers
- Pasteur pipets with rubber bulb
- Bunsen burner and striker
- Gloves

REPORT SHEETS
Experiment 12: Separation of a Ternary Mixture

Name: ______________________________ Section: __________

Lab Partner: ______________________________

Show calculations in the provided space. Always include units and use the proper number of sig figs!

1. Mass of unknown mixture ____________
2. Mass of a 150-mL beaker and watch glass ____________
3. Mass of dried NaCl + beaker + watch glass ____________
4. Mass of dried NaCl ____________
5. Percent sodium chloride in unknown mixture ____________
6. Mass of clean dry evaporating dish ____________
7. Mass of evaporating dish plus dried SiO_2 (after HCl reaction—after each heating)

   1st mass: ____________ 2nd mass: ____________ 3rd mass (see Step 20): ____________
8. Mass of SiO_2 ____________
9. Percent silicon dioxide in unknown mixture ____________
10. Mass of $CaCO_3$ ____________

    (**Remember:** You must subtract the mass of the SiO_2 and NaCl from the mass of the mixture)
11. Percent calcium carbonate in unknown mixture ____________
12. Fill in the table below using the data that you have obtained

% NaCl	% SiO_2	%$CaCO_3$	Total

(Make sure total % adds up to 100%, or nearly so.)

POST-LAB QUESTIONS

1. When water was added to the unknown in Step 3, did you observe a physical or a chemical change? Explain.

2. When decanting and rinsing with water in Steps 13–17, a student accidentally poured off some of the fine particles of SiO_2. How would the student's percents of NaCl, $CaCO_3$, and SiO_2 be affected (i.e., too high, too low, or unchanged)? For each substance, circle your answer and explain.

   % NaCl: increase / decrease / unchanged

   % $CaCO_3$: increase / decrease / unchanged

   % SiO_2: increase / decrease / unchanged

3. In an experiment, a student weighed out 2.510 grams of unknown but recorded the mass as 5.210 grams in the student's laboratory notebook. How would the percents of NaCl, $CaCO_3$, and SiO_2 be affected (i.e., too high, too low, or unchanged)? For each substance, circle your answer and explain.

   % NaCl: increase / decrease / unchanged

   % $CaCO_3$: increase / decrease / unchanged

   % SiO_2: increase / decrease / unchanged.

4. The mass of an evaporating dish with a mixture of sand, NaCl, and $CaCO_3$ is 26.817 g. The mass of the evaporating dish is 22.437 g. Water is added to the mixture, dissolving the salt, and then the liquid is poured off leaving a solid. After the solid is dried, the mass of the remaining solid mixture and the dish is 25.332 g. When 50 mL of 1.1 M HCl is added to the mixture, it fizzes, chemically changing $CaCO_3$ to $CaCl_2$ and CO_2 gas. After this residue is dried, the mass of the residue and the dish is 24.007 g. Calculate the % of sand, % of NaCl, and the % of $CaCO_3$ in the mixture. Show all calculations and report your answer to an appropriate number of sig. figs.

LAB EXERCISE
Experiment 12: Separation of a Ternary Mixture

Name: ______________________________ Section: __________

1. Identify the following changes as chemical or physical:
   a. The steak dinner you ate is being digested and used for energy. ________
   b. Iron rusts when left in a moist environment. ________
   c. Dry ice (solid carbon dioxide) sublimes to form carbon dioxide gas. ________
   d. A student drops a 150-mL beaker that breaks into 17 pieces. ________
   e. A battery terminal corrodes when battery acid is spilled on the terminal. ________

2. Name the physical methods that are used in this experiment to separate the mixture. Identify the procedure steps where each is used. (Example: dissolving in Step 3.)

3. Identify the step of this experiment that is a chemical method used to separate the mixture. Write the chemical equation for the reaction that represents this chemical change.

Notes

Notes

Notes

Notes

Notes

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Notes

Notes

POLYATOMIC IONS

ammonium	NH_4^+	sulfite	SO_3^{2-}
hydronium	H_3O^+	hydrogen sulfate (bisulfate)	HSO_4^-
hydroxide	OH^-	hydrogen sulfite (bisulfite)	HSO_3^-
carbonate	CO_3^{2-}	perchlorate	ClO_4^-
hydrogen carbonate (bicarbonate)	HCO_3^-	chlorate	ClO_3^-
nitrate	NO_3^-	chlorite	ClO_2^-
nitrite	NO_2^-	hypochlorite	ClO^-
phosphate	PO_4^{3-}	cyanide	CN^-
hydrogen phosphate	HPO_4^{2-}	permanganate	MnO_4^-
dihydrogen phosphate	$H_2PO_4^-$	thiosulfate	$S_2O_3^{2-}$
chromate	CrO_4^{2-}	oxalate	$C_2O_4^{2-}$
dichromate	$Cr_2O_7^{2-}$	borate	BO_3^{3-}
acetate	$C_2H_3O_2^-$ or CH_3COO^-	citrate	$C_6H_5O_7^{3-}$ or $C_3H_5O(COO^-)_3$
sulfate	SO_4^{2-}		

NUMBERS AND FORMULAS TO KNOW

Avogadro's Number: 6.022×10^{23}
mass percent (% *m*/*m*) $= \dfrac{\text{mass of solute}}{\text{mass of solution}} \times 100$
volume percent (% *v*/*v*) $= \dfrac{\text{volume of solute}}{\text{volume of solution}} \times 100$
mass volume percent (% *m*/*v*) $= \dfrac{\text{mass of solute (g)}}{\text{volume of solution (mL)}} \times 100$
molarity (M) $= \dfrac{\text{moles of solute}}{\text{liters of solution}}$
$T_K = T_{°C} + 273$
$T_{°F} = (1.8 \times T_{°C}) + 32$
$C_1V_1 = C_2V_2$