

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



CHEM& 121 Workbook

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

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President: Dreis Van Landuyt Project Manager: Janelle Lange Customer Care Lead: Lauren Houseworth

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What Is CHEM& 121?



CHEM& 121 is primarily designed as an introductory level chemistry course for allied health majors (nursing, ultrasound, dental hygiene, nutrition, etc.) and for students who need an intro level chem course for natural science distribution credits. It is intended for students who have never had any chemistry before, although some previous chemistry background, such as high school chemistry, may be helpful.

NOTE: Students intending to take CHEM& 161 (General Chemistry) are advised to take CHEM& 140 Preparatory Chemistry instead of CHEM& 121 as it will better prepare you for science/engineering careers. If you are unsure which course to take, talk to your instructor or an academic advisor. A science advisor is available in the Science Division office. Call 425-564-2321 to make an appointment.

HOW TO PREPARE FOR CHEM& 121

The prerequisite for this course is intermediate algebra (MATH 098) with a C or better. The course requires students to solve mathematical problems with proportions, percentages, and algebra. Using scientific notation and the metric system is also very important. You can prepare by reviewing these topics. The Basic Math Skills Check worksheet in this manual can be very useful to assess your current math skills.

TIPS ON HOW TO SUCCEED

Every student is unique and responds differently to various study techniques, but these are most common for students who are successful. Think back to habits which helped you succeed in past courses or goals in your life (music, exercise, etc.). Try multiple approaches to find the best method for you.

- Visit your instructor's office hours or make an appointment to go over course material you find challenging.
- Do the required homework or recommended exercises to get enough practice and stay current with the material (don't get behind!).
- Find people in class you can study with—working in teams can help you build on each other's strengths, test each other, and motivate you to study.
- Schedule time to study just like you would for classes, work, etc. Keep a log of studying time so you hold yourself accountable and know what you worked on.
- If there are certain topics that are challenging for you, in addition to reading the section from the book, try a YouTube or Google search to find websites and videos which may help you. Do problems and keep hitting those tough areas—in chemistry, concepts build on each other so don't assume the topic will just go away!

CAMPUS RESOURCES

- Many instructors have office hours. These are drop-in hours to get help from your instructor. If you're shy, go with a classmate!
- Check out the Science Study Center (S114). They have a schedule of available tutors in each discipline. http://scidiv.bellevuecollege.edu/SSC/
- Drop-in group tutoring is available at the Academic Success Center (D204). Check for their schedule at http://www.bellevuecollege.edu/asc/.
- If you are struggling with a C or below, you can get free 2-hour one-on-one tutoring. It takes some time to set up so go as soon as you can! Visit http://www.bellevuecollege.edu/asc/ for details. For more information and practice worksheets, visit the BC Chemistry Website and go to the CHEM& 121 webpage: http://bellevuecollege.edu/ chemistry
- Many of you are taking this course to fulfill prerequisites for a career in health sciences or natural sciences. It's a great idea to talk to an advisor to get more information and support. The science advisor is here to help. Call 425-564-2321 or stop by the Science Division office (L200) to make an appointment.

Basic Math Skills Check for CHEM& 121



Complete this assessment without the aid of a calculator. Circle clearly the correct answer.

Calculate the following expressions:

1.	(41 – 5 × 3)/2 + 5 =						
	a. 59	b.	26/7	C.	15	d.	18
2.	61 + (35 – 5) × 3 – 1 =						
	a. 121	b.	150	C.	80	d.	182
3.	$\frac{2}{5} \times \frac{5}{14}$						
	a. 7/19	b.	10/19	C.	1/7	d.	28/25
4.	$32 \div \frac{1}{8}$						
	a. 4	b.	16	C.	256	d.	128
5.	$2^3 \times 2^2 =$						
	a. 64	b.	32	C.	256	d.	16
6.	$4^{5}/4^{2} =$						
	a. 16	b.	16384	C.	256	d.	64
7.	$6^{-3} \times 6^5 =$						
	a. 36	b.	1/36	C.	144	d.	1/144
8.	$7^2/7^{-3} =$						
	a. 343	b.	7 ⁵	C.	49	d.	0.143
9.	$6.0\times10^{23}\times2.0\times10^{-5}$	=					
	a. 1.2 × 10 ²⁹	b.	8.0 × 10 ¹⁹	C.	8.0×10^{29}	d.	1.2 × 10 ¹⁹
10.	$(4.0 \times 10^8)/(2.0 \times 10^5) =$	=					
	a. 2.0×10^{13}	b.	8.0×10^{13}	C.	2.0×10^{3}	d.	$8.0 imes 10^3$

Express the following numbers in correct scientific notation:

11.	0.0276 =						
	a. 2.76×10^{-2}	b.	2.76×10^{-3}	C.	2.76×10^2	d.	$2.76 imes 10^3$
12.	3780 =						
	a. 37.8	b.	3.780×10^{3}	C.	3.780×10^{-3}	d.	0.378
13.	1,694,000,000 =						
	a. 1.694×10^{-9}	b.	1.694 × 10 ⁸	C.	1.694×10^{7}	d.	1.694 × 10 ⁹
14.	$0.0043 \times 10^{-3} =$						
	a. 4.3×10^{-6}	b.	4.3	C.	$4.3 imes 10^{-3}$	d.	4.3×10^3
Ex	press the following nu	mbe	ers in decimal nota	tion	:		
15.	67.5 × 10 ³ =						
	a. 675	b.	6,750	C.	67,500	d.	675,000
16.	$7.3 \times 10^{-3} =$						
	a. 7,300	b.	0.073	C.	730	d.	0.0073
So	lve the following equa	tion	s:				
17.	<i>y</i> + 5 = 24						
	a. 5	b.	120	C.	29	d.	19
18.	<i>z</i> – 17 = 4						
	a. 20	b.	68	C.	21	d.	-13
19.	3 <i>y</i> = 99						
	a. 102	b.	33	C.	96	d.	11
20.	$(7t - 19) \times 3 = 6$						
	a. 3	b.	22/7	C.	-22/7	d.	15
21.	Solve for T : $PV = nR$	Т					
	a. $T = PV - nR$	b.	T = nR/PV	C.	T = PV/nR	d.	T = PV + nR
22		,					
۲۲.	Solve for V: $D = M/V$	/					

23.	So	lve for <i>x</i> : -	$\frac{4}{5} = \frac{3}{1!}$	(5					
	a.	9		b.	10	C.	12	d.	13
24.	So	lve for <i>x</i> : -	$\frac{2}{9} = \frac{6}{2}$	6 K					
	a.	24		b.	27	C.	28	d.	26
25.	Со	nsider the	expressi	on x	<i>x</i> = <i>y</i> / <i>z</i> . If <i>x</i> = 10 and	Z =	5, <i>y</i> =		
	a.	<i>y</i> = 100		b.	<i>y</i> = 5	C.	<i>y</i> = 1/2	d.	<i>y</i> = 50
26.	A 1 bot	2-ounce c ttle of the	an of sod same kin	a co d of	ontains 36 grams of soda?	sug	ar. How many gram	is ar	re in a 20-ounce
	a.	60		b.	58	C.	30	d.	29
27.	Th sel	e futon tha I it to you f	at you wa for an 8%	nt t dis	o buy costs \$450, k count. How much v	out † vou	the salesman says Id the futon cost yo	that u?	t he is willing to
	a.	414		b.	45	C.	4.5	d.	420
28.	Ho	w many ki	lograms	are i	n 218 grams of sug	ar?	(1 kg = 1000g)		
	a.	21.8 kg		b.	2.18 kg	C.	0.218 kg	d.	218,000 kg
29.	Ho	w many m	illimeters	s are	e in 2 meters? (1 m =	= 10	00 mm)		
	a.	20 mm		b.	2,000 mm	C.	0.2 mm	d.	0.002 mm
30.	Ho	w many fe	et are in	15 y	ards? (1 yard = 3 fe	et)			
	a.	18 ft		b.	5 ft	C.	45 ft	d.	12 ft

When you are done, go to the next page and circle your correct answers on the answer key. Count the number of correct answers as indicated. Write your score on each of the four sections.

SELF-GRADING

Question #	Answer	Question #	Answer
1	D	16	D
2	В	17	D
3	С	18	С
4	С	19	В
5	В	20	А
6	D	21	С
7	А	22	D
8	В	23	С
9	D	24	В
10	С	25	D
11	А	26	А
12	В	27	А
13	D	28	С
14	А	29	В
15	С	30	С

QUESTIONS 1–10

A score of less than 7 correct answers means that you should review "number crunching."

QUESTIONS 11–16

A score of less than 5 correct answers means that you should review scientific notation.

QUESTIONS 17–24

A score of less than 6 correct answers means that you should review basic algebra.

QUESTIONS 25–30

A score of less than 5 correct answers means that you should review proportionality and unit conversions.

/6

/10

/8

/6

CHEM& 121 Student Learning Objectives



Learning Objectives	Section(s) in Timberlake 5 th Edition
UNIT 1	
Matter	
1. Define chemistry and provide examples of matter	1.1, 3.1
2. List and describe the key elements of the scientific method	1.2
3. Explain the differences between elements, compounds, and homogeneous and heterogeneous mixtures	3.1
4. Represent elements and compounds using chemical symbolism	4.1
5. Define the states of matter on a molecular level	3.2
6. Describe the difference between chemical and physical properties and changes	3.2
UNIT 2	L
Measurements	
7. Use prefixes in metric measurements (memorize micro, centi, milli, deci, deca, kilo, mega)	2.1, 2.4
8. Solve unit conversion problems using dimensional analysis, proportional reasoning, and/or algebraic expressions for health applications, including mass, volume, length, density, time, dosages, body mass index (BMI), temperature, energy units, and pressures	2.1, 2.5, 2.6, 2.7, 3.3, 3.4, 3.5
 Apply the relationship between significant figures and uncertainty in laboratory measurement(s) (no tracking sig figs in calculations) 	2.2, 2.3
UNIT 3	
Atoms and Isotopes	
10. Describe atomic structure in terms of the subatomic particles which make up the atom	4.3
 Extract information about an element from the periodic table or a given isotope symbol (for example: the element symbol, atomic number, atomic mass, and the number of protons, neutrons, and electrons) 	4.2, 4.3, 4.4
12. Apply the concept of a weighted average to the natural abundance of various isotopes (no calculations)	4.5
UNIT 4	
Electromagnetic Radiation and Bohr Model	
13. List the types of electromagnetic radiation in order of increasing energy and identify the types that are ionizing	4.6
14. Draw a Bohr model diagram for elements #1–20 and determine the number of core and valence electrons in an atom (no electron configurations)	4.6, 4.8
15. Explain the relationship between electron transitions and emission spectra of elements	4.6

Learning Objectives	Section(s) in Timberlake 5 th Edition
16. Explain why the elements in a periodic group exhibit similar chemical properties using the octet rule	4.6, 4.8
17. Identify the location of groups and periods, metals and nonmetals in the Periodic Table	4.2
UNIT 5	
Nuclear Chemistry	1
18. Describe what nuclear radiation is, list natural sources, and explain how it is detected	5.1, 5.3
19. Identify types (alpha, beta, gamma, positron) and properties (charge, mass, penetrating ability, ionizing power) of radioactive decay	5.1, 5.2
20. Solve half-life problems, including carbon dating and common isotopes used in the medical field	5.4
21. Balance nuclear decay reactions	5.2
22. Discuss health effects and medical applications of nuclear and electromagnetic radiation (cancer treatment, tracers, x-rays, PET scan, CT scan)	5.5, 5.6
UNIT 6	
Covalent Bonds and Structure of Molecules	1
23. Define covalent bonding	6.5
24. Write Lewis structures for atoms and simple molecules	6.6
25. Write the formulas and names for binary molecular compounds	6.5
26. Draw simple organic molecules using shorthand notation and identify organic functional groups present in compounds, including biomolecules (fatty acids, monosaccharides, nucleic acids, amino acids), using a provided table of organic functional groups	12.1, 12.2, 12.3, 12.4
27. Predict the geometry of molecules using VSEPR Theory	6.8
28. Use electronegativity and symmetry to predict the polarity of bonds and molecules	6.7, 6.8
UNIT 7	
Solids, Liquids, Gases, and Intermolecular Forces	
29. Determine the type of intermolecular force that can occur for a given substance	6.9
30. Describe how hydrogen bonding works in interactions of biological molecules and affects the properties of water (e.g., surface tension, boiling point)	6.9

Learning Objectives	Section(s) in Timberlake 5 th Edition				
31. Explain the roles that intermolecular forces and the kinetic-molecular theory play in determining the properties and states of matter	6.9, 3.7, 8.1				
32. Name phase changes (freezing, melting, vaporization, condensation, sublimation, deposition) and describe changes in particle spacing, energy, and intermolecular forces	3.7				
33. Interpret a heating or cooling curve (no calculations)	3.7				
UNIT 8					
Ionic Compounds					
34. Predict the charge of common cations and anions using the octet rule and the periodic table	6.1				
35. Illustrate the transfer of electrons in the formation of ions and ionic compounds using Lewis structures	6.1				
36. Write the formulas and names for binary ionic compounds using the Periodic Table and a list of polyatomic ions (optional: use Latin-based names of ions such as ferric vs. ferrous)	6.2, 6.3, 6.4				
37. Summarize major differences between molecular and ionic compounds	6.1, 6.5				
Reaction Rates and Energies					
38. Describe the collision theory of reactions	10.1				
39. Draw potential energy diagrams for exothermic and endothermic reactions and label the reactants, products, transition state, activation energy, and heat of reaction	7.9, 10.1				
40. Describe the ways to alter the rate of a reaction; temperature, concentration, and catalyst	10.1				
UNIT 9					
The Mole and Molar Masses					
 Explain the concept of the mole as a chemical counting unit and in terms of Avogadro's number 	7.4				
42. Compute molar masses of elements and compounds	7.5				
43. Use molar masses and Avogadro's number in calculations involving masses, moles, and number of particles of chemical substances	7.5				
UNIT 10					
Chemical Reactions	I				
44. Write and balance chemical equations and relate this to the Law of Conservation of Mass	7.1				
45. Interpret the information conveyed by the coefficients of a balanced chemical equation on the molecular level	7.1				

Learning Objectives	Section(s) in Timberlake 5 th Edition		
Redox			
46. Define oxidation and reduction as a transfer of electrons	7.3		
47. Recognize simple redox reactions based on reactants and products in chemical equations and without use of oxidation numbers	7.3		
48. Complete and balance combustion reactions	7.2		
Stoichiometry			
49. Solve stoichiometric problems using a chemical equation and amounts of reactants or products in particles, moles, or grams	7.6, 7.7		
50. Identify a limiting reactant using a chemical equation and amounts of reactants in moles (not from mass)	7.8		
51. Dynamic equilibrium			
52. Define dynamic equilibrium (no equilibrium constant calculations)	10.2		
UNIT 11			
Solutions			
53. Describe solutions as homogeneous mixtures and in terms of solutes and solvents	9.1, 9.6		
54. Predict whether two substances will mix by comparing molecular polarity	9.1		
55. Explain using words or molecular-scale drawings to explain the dissolving process of ionic and polar molecular substances in water (i.e., solvation, dissociation, hydration, etc.)	9.1, 9.2		
56. Distinguish the electrolytic properties of soluble ionic compounds and molecular compounds	9.2		
57. Describe a saturated solution in terms of dynamic equilibrium	9.3, 10.2		
58. Perform calculations and describe how to prepare solutions using concentration units of percent (<i>m/m</i> , <i>m/v</i> , <i>v/v</i> , also use of <i>w/v</i> , <i>w/w</i>) and molarity	9.4		
59. Perform calculations and describe how to prepare solutions using the concept of dilution	9.5		
60. Describe the process of osmosis and predict direction of solvent flow between two solutions in health applications, including osmosis in biological cells and dialysis	9.6		
UNIT 12			
Apply Le Chatelier's Principle	10.5		
61. Write the equilibrium constant expression for a reaction	10.3		
62. Use the magnitude of the equilibrium constant to predict the extent of reaction (favors reactants or products)	10.4		

Learning Objectives	Section(s) in Timberlake 5 th Edition		
Acids and Bases			
63. Define and recognize Brønsted-Lowry acids and bases	11.2		
64. Identify conjugate acid-base pairs	11.1		
65. Write reactions of molecular acids and bases with water	11.2		
66. Describe the differences between strong and weak acids and bases in terms of conductivity and in terms of reversible chemical reactions and equilibrium	9.2, 11.3, 11.4		
67. Recognize the structures of simple organic acids and bases			
Neutralization Reactions			
68. Write acid-base neutralization reactions, including carbonates/bicarbonates	11.7		
69. Solve simple neutralization problems	11.7, 11.8		
pH Scale			
70. Use the pH scale to determine whether a solution is acidic, neutral, or basic	11.6		
71. Calculate the pH of a solution given the hydronium ion concentration	11.6		
Buffers			
72. Recognize the components of a buffer solution	11.9		
73. Explain using chemical reactions how a buffer solution resists changes in pH	11.9		
74. Determine whether changes in concentration of hydronium ion or carbon dioxide will cause acidosis or alkalosis	11.9		

CHEM& 121 Pre- and Post-Exam Reflection



It can be very disappointing to not spend enough time studying to do well, or to spend a lot of time studying and still not perform well on a test. This activity is designed to help you prepare for an exam and assess your exam performance for future improvement. Answer the questions sincerely and to the best of your knowledge.

PRE-EXAM PREPARATION

- 1. Schedule time every day to study, even if it's in small chunks (10–15 minutes several times a day). Treat it like exercise!
- 2. Every instructor will want you to review your notes from class as a starting point. Here is a suggested list of other things to look at:
 - a. If your notes are messy, try re-writing them neatly. Or re-write and organize them as a way of reviewing the material.
 - b. Summarize your notes from each lecture or topic onto one sheet of paper. This will make it easier to review and provide a resource for studying for final exams, future standardized exams, etc.
 - c. Make sure you practice, practice, practice! Do problems from each topic. Do them multiple times. Check answers to make sure you are on the right track, or get some assistance (book, instructor, friends) as needed.
 - d. Use the CHEM& 121 course student learning objectives listed in this book (with the guidance of your instructor) as a checklist of what you should be able to do on your own on an exam (without notes or help from others).

POST-EXAM REFLECTION

- 1. After you get your exam back, identify your trouble spots. Estimate the percentage of points you lost due to each of the following:
 - a. Did not understand what the problem was asking and didn't get clarification _____
 - b. Skipped a problem—didn't see it or forgot to go back to it
 - c. Didn't make my answer clear enough
 - d. Trouble with calculations
 - e. Trouble with multiple choice questions
 - f. Lack of understanding of the concept
 - g. Missed points on rounding using sig figs or forgetting to write the unit
 - h. Did not study something that was asked but knew it was something that was covered
 - i. Did not know something was going to be covered
 - j. Did not use the CHEM& 121 Student Learning Objectives as a study guide.
 - k. Not knowing how to approach the problem or what to do
 - I. Did not practice using my calculator or forgot to take my calculator to the exam
 - m. Careless mistakes, things you should have gotten right but didn't for unknown reasons
 - n. Other
- 2. Based on your responses to the questions above, name at least three things you plan to do differently in preparing for the next exam. Be specific!
- 3. What support services will you seek out to support your learning and preparation for the next exam? (Go to instructor office hours, Science Study Center, Academic Success Center, Multicultural Services, or Trio.)



You will be using symbols for many different elements, but the most common that you will use are the first 20 elements. It is not necessary for you to memorize the element number or order; just focus on recognizing the name and element symbol. The more you practice this, the less time you will spend looking at the Periodic Table of the Elements to find the element that you need!

Write the names and symbols for elements #1-20.

Element Name	Symbol	Element Name	Symbol
hydrogen	Н	sodium	Na
helium	He	magnesium	Mg
lithium	Li	aluminum	AI
beryllium	Be	silicon	Si
boron	В	phosphorus	Р
carbon	C	sulfur	S
nitrogen	N	chlorine	CI
oxygen	0	argon	Ar
flourine	F	potassium	K
neon	Ne	calcium	Ca

Make sure to capitalize the first letter and use lower case for the second letter in the symbol.

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	d.	copper <mark>Cu</mark>	g.	gold <mark>Au</mark>	i.	iron <mark>Fe</mark>
b.	mercury <mark>Hg</mark>	e.	potassium <mark>K</mark>	h.	silver <mark>Ag</mark>	j.	sodium <mark>Na</mark>
C.	tin <mark>Sn</mark>	f.	tungsten W				

CHEM& 121 WORKBOOK



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!). Chemical elements may be named from various sources; sometimes based on the person who discovered it, or the place it was discovered. Some elements have Latin or Greek roots (e.g., Ag – silver derived from its Latin name "Argentum").

The first 36 elements (plus a few others) in the periodic table are fairly common. Consider memorizing them so that you can 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 a laminated CHEM& 121 periodic table in the classroom so 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.	0 Oxygen	C.	Na Sodium
b.	Ca Calcium	d.	Ag <mark>Silver</mark>

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

a.	neon <mark>Ne</mark>	f.	tungsten
b.	chlorine Cl	q.	gold

- c. potassium K h. radon**NOTE:** There is often confusion between radium and radon.
- 16 Unit 1

e. manganese

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, a compound of carbon and oxygen.

Checkpoint 6: What is the difference between an element and a compound?

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 7: What is the chemical formula of carbon dioxide? CO₂
- Checkpoint 8: How many atoms in an ozone molecule, O₃? 3
- Checkpoint 9: How many atoms are in CuSO₄? 6
- Checkpoint 10: How many atoms are in Ba(OH)₂? HINT: The ()₂ means that everything inside the () is multiplied by two. 5

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 11: Which boxes show molecules? List all that apply.

B and C

Checkpoint 12: 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 13: Which *one* of these is considered a mixture, A or **B**? Explain.



Only one substance (pure).



Two or more substances.

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 14: Which of the following is a pure element (no mixtures)? B
- Checkpoint 15: Which of the following is a pure compound (no mixtures)? D
- Checkpoint 16: Which of the following are mixtures? A and C
- Checkpoint 17: Which of the following is a homogeneous mixture? C
- Checkpoint 18: Which of the following is a heterogeneous mixture? A



IMPORTANT!

Notice in Checkpoint 13 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 , CI_2 , H_2 , O_2 , and F_2 (Mnemonic: Brinchhof or Hofbrinch)

Another note: H atom is an element; H₂ 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 19: 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
- b. In YouTube, search for **"ammonium dichromate volcano"** video (length is 0:59). Or here's the URL: http://www.youtube.com/watch?v=Ula2NWi3Q34

Checkpoint 20: Are each of the following statements *chemical* or *physical* descriptions?

- a. alcohol evaporates easily Physical (liquid to gas-it's still alcohol)
- b. iron rusts over time Chemical (iron reacts with air)
- c. natural gas is burned for heat Chemical (gas reacts with oxygen)
- d. salt is crushed into a fine powder **Physical (appearance changes, not chemi**cal composition)
- e. adding food color to icing boiling an egg
- Checkpoint 21: Lightning converts O₂ (oxygen we breathe) to O₃ (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 The simplest form of matter
 - b. compound A substance composed of two or more elements

- c. pure substance Only one substance is present
- d. mixture More than one substance is present
- 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 (BrINCIHOF)!

These are diatomic! (0₂ not 0, N₂ not N)

$\infty N_2 \qquad \infty N_2$	∞0 ₂	∞ 0 ₂	
	∞ N ₂	∞N_2	
© N ₂			

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

Only one substance, H₂O, is present.

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

It depends ... but probbly a mixture! Tap water is a mixutre. Distilled water is a pure substance.

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

No, hydrogen and oxygen are separate substances unitil a chemical change turns them into water (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 vs. liquid)
 - c. soda with ice cubes Heterogeneous mixture (solid (ice)/liquid (soda)/gas (bubbles))
 - d. fresh garden salad Heterogeneous mixture (lettuce, carrtos, etc.)tiramisu
 - e. toothpaste vs. Na₂HPO₄
- 7. 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
 - d. Draw a diatomic molecule







Liquid

Solid (atoms uniformly mixed) Two separate phases, all are molecules

Elements, Compounds, and 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 **element or compound**.
- An element is always uniform all the way through (homogeneous).
- An element **cannot** be separated into simpler materials (except during nuclear reactions).
- Over 100 existing elements are listed and classified in 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 cannot 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 element (E), compound (C) or heterogeneous mixture (HM). Write the letter X if it is none of these.

E Diamond (C)	C Sugar	C Baking Soda	E Titanium (Ti)
M Air	$(C_6H_{12}O_6)$	(NaHCO ₃)	E Iron (Fe)
E Krypton (K)	C Sulfuric Acid	M Milk	X Electricity
C Water (H ₋ O)	(H ₂ SU ₄)	M Gasoline	, M Poncorn
	E Bismuth (Bi)	E Uranium (U)	
(NH ₃)		M Pail of Garbage	IVI A dog
M Wood	(CH ₃ UH)	¥ Eporav	E Gold (Au)
	C Salt (NaCl)	A Lifergy	M Pizza
C Dry Ice (CO ₂)	M Bronze (Cu + other meta	M Ink Is)	M Concrete

Why is air a heterogeneous mixture?

PART 3

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



Tutorial: Chemical vs. Physical Changes

CHEMICAL CHANGES

During a chemical change reactants will chemically react to form products, which are different chemical substances than the original reactants. There is a change in the chemical identity of the substances. You can recognize chemical change in a provided equation by looking for different chemical substances before and after the reaction takes place (on left and right side of the reaction arrow). In a molecular-scale drawing, look for changes to how the atoms are bonded together by noticing differences in how different shaded spheres are connected together.

Some common examples of chemical changes and representative chemical equations:

- Rust formation
- Silver tarnishing
- Combustion of a fuel

- Metabolism of glucose/digestion
- Baking
- Bleaching

PHYSICAL CHANGES

During a physical change the chemical substance changes physical state (solid, liquid, gas) or form (shape, size) but there is no change in the chemical identity of the substance. In a provided equation, the reactant and product will have the same chemical formula. In a molecular-scale drawing, the groups of shaded spheres will not change but their spacing will change.

Some common examples of physical changes and representative chemical equations:

- Phase changes: freezing, melting, subliming, condensing, vaporizing, deposition
- Shape/size changes: cutting, bending, tearing, slicing

PRACTICE

- 1. Label each of the following as physical (P) or chemical (C) changes.
 - a. Tearing aluminum foil
 - b. Rusting of iron
 - c. Slicing a banana
 - d. Bleaching hair

- P e. Grinding coffee beans
- **C** f. Carving plastic
- P g. Burning paper C
- C

Ρ

Ρ

Notes: Physical States



PHYSICAL STATES OF MATTER

Gases, Liquids, and Solids on the Molecular Level						
		8.000				
Particle spacing	Farthest apart	Close	Closest together			
Volume and shape	Indefinite volume Indefinite shape	Definite volume Indefinite shape	Definite volume Definite shape			
Compressibile?	Most compressible	No (mostly no)	Least compressible			
Flow?	Can flow	Can flow	Cannot flow			
Relative kinetic energy	Highest KE	Intermediate KE to flow	Lowest KE			
Attractive forces	Least attractive forces	Strong IMFs	Greatest attractive forces			

Why? It's a balancing act! The balance between the kinetic energy (KE) of the molecules and the attractive forces (intermolecular forces) between the molecules determines the physical state of a substance.

PHYSICAL STATE CHANGES

ENDOTHERMIC PROCESSES

IMFs are **broken** as heat (energy) is **absorbed/added**.



E Gas

60

B and D are phase

changes (no ΔT)

50

D

Liquid and gas

HEATING CURVES

A **heating curve** is a graph of the change in temperature vs. heat added to a system. (For cooling curves, heat is removed.) Here is a heating curve for 18 grams (1 mole) of water.

 Match each of the following statements to positions A–E on the graph.



100

50

0

-50

Solid &

liquid

В

A Solid

10

C Liquid

20

30

Heat Added (kJ)

40

Temperature

e. Heating water at 100 °C results in no temperature change.
CHEM& 121 WORKBOOK

UNIT 2 MEASUREMENTS



Measurements and Significant Figures



- 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 Most decimal places
- 3. Measure the arrow using the ruler provided below it using the number of digits allowed by this ruler. **5.3 cm (answers may vary slightly)**

1 2 3 4 5 6 7 8 9 10 cm											
	1	2	3	4	5	6	7	8	9	10	cm

 Provide a measurement for the temperature displayed on this Celsius thermometer (a portion of the thermometer is shown, the dashed line indicates the temperature). 20.32 °C (estimated digit, may vary ... 20.31, 20.32, or 20.33 °C okay)



- 5. Beaker
 - Report a measurement for the volume in this beaker.
 360 mL (estimated digit)
 - b. Would this beaker be a good choice for measuring a volume of 5.0 mL? No, because the tens place is estimated, so measuring to the ones place is not possible.
- 6. Scientific Law, Theory, and Hypothesis Examples
 - a. If I give a plant an unlimited amount of sunlight, then the plant will grow to its largest possible size.
 - b. To every action, there is an equal and opposite reaction.



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 s.f.** 2) 3.05 mL **3 s.f.** 3) 3.00 mL **3 s.f.** 4) 3.000 mL **4 s.f.**

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 s.f.** 6) 0.300 mL **3 s.f.**

7) 0.00300 mL **3 s.f.**

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 s.f .	10) 301 mL <mark>3 s.f.</mark>	12)	30000. mL <mark>5 s.f.</mark>	14) 30000.00	mL 7 s.f.
9)	31 mL <mark>2 s.f.</mark>	11) 310 mL <mark>2 s.f</mark> .	13)	Decimal denotes signed and the signe	zeros Landa	Trailing zeros after decimal (measured!)

List of Metric Prefixes



Prefix	Symbol	Mathematical Meaning	In words
Yotta-	Y	10 ²⁴	
Zetta-	Z	10 ²¹	
Exa-	E	10 ¹⁸	
Peta-	Р	10 ¹⁵	
Tera-	Т	10 ¹²	trillion
Giga-	iga- G 10 ⁹		billion
Mega-	Μ	10 ⁶	million
Kilo-	k	10 ³	thousand
Hecto-	h	10 ²	hundred
Deca-	da	10 ¹	ten
		1 = 10 ⁰	one
Deci-	d	10 ⁻¹	tenth
Centi-	C	10 ⁻²	hundredth
Milli-	m	10 ⁻³	thousandth
Micro-	μ	10 ⁻⁶	millionth
Nano-	n	10 ⁻⁹	billionth
Pico-	р	10 ⁻¹²	trillionth
Femto-	f	10 ⁻¹⁵	
Atto-	а	10 ⁻¹⁸	
Zepto-	Z	10 ⁻²¹	
Yocto-	У	10 ⁻²⁴	

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

COMMONLY USED

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

Metric System Units and Definitions



1.	Fill in the following blanks for units of length.										
	a.	1 megameter (Mm) =	10 ⁶ m		or	1 m =	10 ⁻⁶ Mm				
	b.	1 kilometer (km) =	10 ³ or 1	000 m	or	1 m =	10 ⁻³ or 0.001 km				
	C.	1 decameter (dam) =	10 m		or	1 m =	10⁻¹ or 0.1 dam				
	d.	1 centimeter (cm) =	10 ⁻² or	0.01 m	or	1 m =	10 ² or 100 cm				
	e.	1 decimeter (dm) =	10 ⁻¹ or	0.1 m	or	1 m =	10 dm				
	f.	1 millimeter (mm) =	$10^{-3} \mathrm{or}$	0.001 m	or	1 m =	10 ³ or 1000 mm				
	g.	1 micrometer (µm) =	10 ⁻⁶ m		or	1 m =	10⁶ μ m				
2.	Give the full names of each of the following units:										
	a.	mL milliliter		с.	ma mi l	ligram					
	b.	das decasecond		d.	cL cen	tiliter					
•											
3.	VVr	ite symbols for each of	t the fol	lowing units:							
	a.	liter L	C.	kilosecond <mark>ks</mark>		е.	microliter µL or mcL				
	b.	decimeter <mark>dm</mark>	d.	megameter M	m	f.	milligram <mark>mg</mark>				
4.	Express each of the following in terms of their base units (e.g., 1 mL = 0.001 L).										
	a.	1 cm 10 ⁻² m or 0.01 m b.	1 mg 1	0 ⁻³ g or 0.001 g C.	1 km 10)00 m or 1	0 ³ m d. 1 μL 10 ⁻⁶ L or 0.000001 g				
5.	Wr	ite the correct prefixes	s for the	e following mea	sureme	nts:					
	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.	Wł	nich 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 = 10³ or 1000 mL
 j. 1 m = 10² or 100 cm
 b. 1 km = 10³ or 1000 m
 k. 1 mL = 10⁻³ or 0.001 L
 - c. $1 \text{ mg} = 10^{-3} \text{ or } 0.001 \text{ g}$ I. $1 \text{ L} = 10^{6} \mu \text{ L}$
 - d. $1 \mu L = 10^{-6} L$ m. $1 \text{ mm} = 10^{-3} \text{ or } 0.001 \text{ m}$
 - e. $1 \text{ cm} = 10^{-2} \text{ or } 0.01 \text{ m}$ n. $1 \text{ ML} = 10^{6} \text{ L}$
 - f. $1 g = 10^6 \mu g$ o. $1 L = 10^6 \mu L$
 - g. $1 dL = 10^{-1} L$ p. $1 m = 10^{-1} or 0.1 Dm$
 - h. $1 \mu m = 10^{-6} m$ q. $1 mL = 10^{-3} or 0.001 L$
 - i. 1 Dg = 10 g r. $1 \text{ L} = 10^{-6} \text{ ML}$

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 μ m = 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 = 10⁶ μ m if you think in terms of 1 of the larger unit equals a lot of the smaller unit.

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

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:

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

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

METHOD #2: MOVING THE DECIMAL POINT

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

Μ			k			Base	d	С	m			μ	
10 ⁶			10 ³			Unit	10 ⁻¹	10 ⁻²	10 ⁻³			10 ⁻⁶	
1	I.	1	I.	1	1	I.	1	I.	I.	1	1	1	1

Find the base unit (marked as –) then count spaces over to the μ symbol. There are six spaces, so there is a difference of 10⁶ 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.)

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.

- с. М 10⁶or 1,000,000 e. µ 10⁻⁶or 0.000001 a. m **10⁻³or 0.001** d. c **10⁻²or 0.01** f. d **10⁻¹or 0.1** b. k **10³or 1000**
- 2. Complete each conversion as a fraction with the correct number.

2	1 cm	or	0.01 m	٥	1 mL	or	0.001 L
α.	0.01 m	U	1 cm	6.	0.001 L	U	1 mL
b.	1 m	or	100 cm	f	1 L	or	1000 mL
	100 cm	U	1 m	1.	1000 mL	U	1 L
•	1 kg	or	1000 g	n	1 µs	or	10 ⁻⁶ s
0.	1000 g	U	1 kg	y.	10⁻⁶ s	U	1 µs
d.	1000 g	or	1 kg	h	10⁶ µs	or	1 s
	1 kg		1000 g	11.	1s	U	10 ⁶ µs

- 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.
$$30 \text{ cm} \times \left(\frac{1 \text{ m}}{100 \text{ cm}}\right) = 0.3 \text{ m}$$
 c. $150 \text{ g} \times \left(\frac{1 \text{ kg}}{1,000 \text{ g}}\right) = 0.15 \text{ kg}$
b. $45 \text{ m} \times \left(\frac{100 \text{ cm}}{1 \text{ m}}\right) = 4500 \text{ cm}$ d. $10 \text{ mL} \times \left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.01 \text{ L}$
e. $0.005 \text{ ks} \times \left(\frac{10^3 \text{ s}}{1 \text{ ks}}\right) \times \left(\frac{10^6 \text{ \mu s}}{1 \text{ s}}\right) = 5 \times 10^6 \text{ \mu s}$

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

- - -

a. 900 g to kg
900 g
$$\left(\frac{1 \text{ kg}}{1000 \text{ g}}\right) = 0.9 \text{ kg}$$

b. 3.4 cm to m
3.4 cm $\left(\frac{1 \text{ m}}{10^2 \text{ cm}}\right) = 0.034 \text{ m}$
(. 800 mm $\left(\frac{\text{m}}{1000 \text{ mm}}\right) = 0.8 \text{ m}$
(. 0.0450 kg to cg
(. 2.87 kL to μ L
0.0450 kg $\left(\frac{10^3 \text{ g}}{1 \text{ kg}}\right) \left(\frac{10^{-2} \text{ cg}}{1 \text{ g}}\right)$
2.87 kL $\left(\frac{10^3 \text{ L}}{1 \text{ kL}}\right) \left(\frac{10^6 \mu \text{ L}}{1 \text{ L}}\right) = 2.87 \times 10^9 \mu \text{ L}$
= (4.50 × 10⁻²)(10³)(10²) = 4.50 × 10³ cg

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.



	Ab	breviation	Base Unit		Ab	breviation	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	dag	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	f.	480 cm = 48.0 dm
b.	5 L = 5000 mL	g.	8 mm = 0.8 cm
C.	16 cm = 160 mm	h.	65 daL = 650,000 mL
d.	104 km = 104,000 m	i.	198 g = 0.198 kg
e.	2500 s = 2,500 × 10⁹ or 2.5 × 10¹² n	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	f.	3.6 n > 360 p
b.	5 g > 508 mg	g.	$4 \times 10^{-7} \mathrm{m} = 400 \mathrm{nm}$
C.	536 cm > 53.4 dm	h.	3.5×10^{-3} mm = 3.5μ
d.	1500 mL = 1.5 L	i.	100 p < 10 ⁻⁶ mL
e.	500 mg < 5 g	j.	500 g < 5 T

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

 $2.14 \text{ kg}\left(\frac{1000 \text{ g}}{1 \text{ kg}}\right) = 2.14 \times 10^3 \text{ g or } 2140 \text{ g}$

2. 6172 mm to meters

 $6172 \text{ mm} \left(\frac{1 \text{ m}}{10^3 \text{ mm}}\right) = 6.172 \text{ m}$

3. 0.0256 L to milliliters

 $0.0256 L \left(\frac{1000 mL}{1 L} \right) = 25.6 mL$

- 4. 3.14×10^{-5} g to micrograms 3.14×10^{-5} g $\left(\frac{10^{6} \mu g}{1 \text{ g}}\right) = 31.4 \mu g$
- 5. 1.01×10^{-8} s to deciseconds

$$1.01 \times 10^{-8} \text{ s} \left(\frac{10 \text{ ds}}{1 \text{ s}} \right) = 1.01 \times 10^{-7} \text{ ds}$$

6. How many milliliters are in 5 liters?

 $5 L \left(\frac{1000 mL}{1 L} \right) = 5000 mL or 5 \times 10^3 mL$

7. How many kilograms are in 598 grams?

$$\frac{598 \text{ g}}{\left(\frac{1 \text{ kg}}{1000 \text{ g}}\right)} = 0.598 \text{ kg}$$

8. How many centiseconds are in 12 seconds?

$$\frac{12 \text{ s}}{1 \text{ s}} \left(\frac{100 \text{ cs}}{1 \text{ s}}\right) = 1200 \text{ cs or } 1.2 \times 10^3 \text{ cs}$$

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?

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

1. How many millimeters are in 29 micrometers?

29 μ m $\left(\frac{10^{-6} \text{ m}}{1 \ \mu\text{m}}\right)\left(\frac{10^{3} \text{ mm}}{1 \text{ m}}\right) = 0.029 \text{ mm or } 2.9 \times 10^{-2} \text{ mm}$

2. How many kilograms are in 100 gigagrams?

$$\frac{100 \text{ Gg}}{1 \text{ Gg}} \left(\frac{10^9 \text{ g}}{1 \text{ Gg}} \right) \left(\frac{10^{-3} \text{ kg}}{1 \text{ g}} \right) = 1 \times 10^8 \text{ kg}$$

3. Convert 1316 mg to kilograms.

$$1316 \text{ mg}\left(\frac{10^{-3} \text{ g}}{1 \text{ mg}}\right) \left(\frac{10^{-3} \text{ kg}}{1 \text{ g}}\right) = 1.316 \times 10^{-3} \text{ kg}$$

4. Convert 5.88 mL to deciliters.

$$5.88 \text{ mL} \left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right) \left(\frac{10 \text{ dL}}{1 \text{ L}}\right) = 0.0588 \text{ dL or } 5.88 \times 10^{-2} \text{ dL}$$

5. Convert 2.19 cm to micrometers.

2.19 cm
$$\left(\frac{10^{-2} \text{ m}}{1 \text{ cm}}\right) \left(\frac{10^{6} \mu \text{m}}{1 \text{ m}}\right) = 2.19 \times 10^{4} \mu \text{m}$$

Two-Step Conversions



Some conversions require more than one step. Here are some two-step problems that will require you to use two conversion factors.

EXAMPLE

How many seconds are in 2 hours?

 $2h \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} 2 \times 60 \times 60 \text{ s} = 7200 \text{ s}$

Use these relationships:

- 12 inch (in) = 1 foot
- 1 inch (in) = 2.54 centimeter (cm)
- 1 pound (lb) = 454 grams
- 1 quart (qt) = 0.946 Liters (L)
- 1 miles (mi) = 5280 feet (ft)
- 1 quart (qt) = 2 pints (pt)
- 1. What is 0.0032 gallons in L? **0.0032 gal** $\left(\frac{4 \text{ qt}}{1 \text{ gal}}\right) \left(\frac{0.946 \text{ L}}{1 \text{ qt}}\right) = 0.012 \text{ L}$

2. Convert 4.1 kJ to cal.
4.1 kJ
$$\left(\frac{1,000 \text{ J}}{1 \text{ kJ}}\right) \left(\frac{1 \text{ cal}}{4.184 \text{ J}}\right) = 979.92 \dots \Rightarrow 980 \text{ cal } (2 \text{ s.f.})$$

3. Convert 55 oz to lb.

$$\frac{55 \text{ oz}}{1 \text{ oz}} \left(\frac{28.35 \text{ g}}{1 \text{ oz}}\right) \left(\frac{1 \text{ lb}}{454 \text{ g}}\right) = 3.434 \dots \Rightarrow 3.4 \text{ lb} (2 \text{ s.f.})$$

- 4. A person is 168 centimeters tall. How many feet is this? $\frac{168 \text{ cm}}{2.54 \text{ cm}} \left(\frac{1 \text{ ft}}{12 \text{ in}}\right) = 5.5118 \dots 5.51 \text{ ft} (3 \text{ s.f.})$
- 5. A bottle of water has a volume of 500 mL. How many quarts is this?

$$\frac{500 \text{ mL}}{\left(\frac{10^{-3} \text{ L}}{1 \text{ mL}}\right)} \left(\frac{1 \text{ qt}}{0.946 \text{ L}}\right) = 0.5 \text{ qt (1 s.f.)}$$

6. A precast concrete block has a mass of 0.250 tons. What is the mass of this concrete block in grams?

0.250 tons $\left(\frac{2,000 \text{ lb}}{1 \text{ ton}}\right) \left(\frac{454 \text{ g}}{1 \text{ lb}}\right) = 227,000 \text{ g} (3 \text{ s.f.})$

7. The pressure reading from a barometer is 742 mm Hg. Express this reading in kilopascals, kPa.

40 Unit 2 742 mm Hg $\left(\frac{1.013 \times 10^5 \text{ Pa}}{760 \text{ mm Hg}}\right) \left(\frac{1 \text{ kPa}}{1,000 \text{ Pa}}\right) = 98.9 \text{ kPa}$

- 4 quarts = 1 gallon (gal)
- 1 ounce (oz) = 28.35 grams(g)
- 1 cal = 4.184 J
- 1 ton = 2000 lb
- 760 mm Hg = 1.013×10^5 Pa



LIQUID-BASED VOLUME UNITS RELATED TO CUBED LENGTH VOLUME UNITS

How many cubic millimeters (cm³) are in a 2-L bottle of soda? At first glance, it may seem like the units are not related because the unit "cm³" does not contain "L" and instead contains "cm" which is a length, not a volume. But, the unit of cm³ 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.

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

The most commonly memorized of these is

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

PRACTICE

1. How many milliliters are in 853 cm³?

 $\frac{853 \text{ cm}^3}{1 \text{ cm}^3} = 853 \text{ mL}$

2. How many cubic centimeters are in 3597 mL?

 $3,597 \text{ mL}\left(\frac{1 \text{ cm}^3}{1 \text{ mL}}\right) = 3,597 \text{ cm}^3$

- 3. Convert 0.020 L to cm³. 0.020 L $\left(\frac{1,000 \text{ mL}}{1 \text{ L}}\right) \left(\frac{1 \text{ cm}^3}{1 \text{ mL}}\right) = 20 \text{ cm}^3$
- 4. How many cubic centimeters are in 8500 gallons?

8,500 gal $\left(\frac{4 \text{ quarts}}{1 \text{ gal}}\right) \left(\frac{0.946 \text{ L}}{1 \text{ quart}}\right) \left(\frac{1,000 \text{ mL}}{1 \text{ L}}\right) \left(\frac{1 \text{ cm}^3}{1 \text{ mL}}\right) = 3.2 \times 10^7 \text{ cm}^3$

5. How many cubic meters are in 8500 gallons?

$$3.2 \times 10^7 \text{ mL} \left(\frac{1 \text{ m}^3}{1 \times 10^6 \text{ mL}} \right) = 3.2 \times 10^1 \text{ or } 32 \text{ m}^3$$

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:



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?



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

$$\frac{15.6 \text{ L}}{\left(\frac{0.70 \text{ kg}}{1 \text{ L}}\right)} = 10.9 \text{ kg}$$

- 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. <u>1 gram</u> or <u>29.7 kJ</u> 1 gram
 - b. How much heat in kilojoules is released when 4.274 grams of liquid ethanol are burned?

 $4.274 \text{ g}\left(\frac{29.7 \text{ kJ}}{1 \text{ g}}\right) = 126.9 \text{ kJ}$

- 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/cm³.
 - a. Write out a unit map to solve this problem. grams \rightarrow cm³
 - b. Solve the problem.

$$\frac{242.6 \text{ g}}{4 \text{ s.f.}} \left(\frac{\text{cm}^3}{7.86 \text{ g} (3 \text{ s.f.})} \right) = 30.87 \text{ cm}^3$$

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 × 10³ kJ of energy per day to survive, how many days could he survive on his fat alone?

 $\frac{11 \text{ kg}}{1 \text{ kg}} \left(\frac{1000 \text{ g}}{1 \text{ kg}} \right) \left(\frac{38 \text{ kJ}}{1 \text{ g}} \right) \left(\frac{1 \text{ day}}{8.0 \times 10^3 \text{ kJ}} \right) = 52.25 \rightarrow 53 \text{ days} (2 \text{ s.f.})$

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

 $\frac{1.5 \text{ mg}}{\text{kg body weight}} = \frac{20. \text{ mg}}{\text{mL}}$ b. Solve the problem. $72.6 \text{ lb} \left(\frac{1 \text{ kg}}{2.205 \text{ lb}}\right) \left(\frac{1.5 \text{ mg}}{1 \text{ kg}}\right) \left(\frac{\text{mL}}{20. \text{ mg}}\right) = 12 \text{ mL}$ body weight dose in mg dose in mL

CONVERTING BOTH NUMERATOR & DENOMINATOR

EXAMPLE

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

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

Potentially helpful information:

1. Average gasoline mileage for a particular sports utility vehicle is 13 miles per gallon of gasoline. Express this miles/gal value in km/L.

$$\frac{13 \text{ mi}}{\text{gal}} \times \left(\frac{1.609 \text{ km}}{1 \text{ mi}}\right) \times \left(\frac{1 \text{ gal}}{4 \text{ qt}}\right) \times \left(\frac{1 \text{ qt}}{0.946 \text{ L}}\right) = 5.5 \text{ km/L}$$

2. The kidneys of a normal adult male filter 125 milliliters of blood per minute. How many gallons of blood are filtered per day?

$$\frac{125 \text{ mL}}{1 \text{ min}} \times \left(\frac{60 \text{ min}}{1 \text{ hr}}\right) \times \left(\frac{24 \text{ hr}}{1 \text{ day}}\right) \times \left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) \times \left(\frac{1 \text{ qt}}{0.946 \text{ L}}\right) \times \left(\frac{1 \text{ gal}}{4 \text{ qt}}\right) = 47.6 \text{ gal/day} (3 \text{ s.f.})$$

Dosage Problems



In nursing school, you will often be asked to solve dosage problems. Dosage is the amount of medication prescribed to a patient, and is often dependent on the patient's body mass/ weight. These problems can be solved using unit analysis (dimensional analysis).



1. The daily dose of ampicillin (an antibiotic) for treatment of an ear infection is 115 mg/kg of body weight. What is the daily dose in mg for a 55-kg person?

55 kg body weight $\left(\frac{115 \text{ mg antibiotic}}{1 \text{ kg body weight}}\right) = 6325 \rightarrow 6300 \text{ mg antibiotic}$

2. A physician ordered 0.50 mg of atrophine intramuscularly (injected directly into muscle) for a patient. If atropine is available as 0.10 mg/mL of solution, how many milliliters is needed to be administered?

0.50 mg atrophine $\left(\frac{1 \text{ mL solution}}{0.10 \text{ mg atrophine}}\right) = 5.0 \text{ mL solution (2 s.f.)}$

3. A 28-kg child is to receive 25 mg ampicillin/kg body weight. If stock on hand is 250 mg/capsule, how many capsules should be given?

28 kg body weight $\left(\frac{25 \text{ mg drug}}{1 \text{ kg body weight}}\right) \times \left(\frac{1 \text{ capsule}}{250 \text{ mg drug}}\right) = 2.8 \text{ capsules}$

4. A patient in the hospital is given an intravenous fluid that must deliver 1000 cc of a dextrose (sugar) solution over eight 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? Round the answer to two significant figures.

$$\frac{1000 \text{ cc}}{8 \text{ hr}} \times \left(\frac{1 \text{ hr}}{60 \text{ min}}\right) \times \left(\frac{15 \text{ drops}}{1 \text{ cc}}\right) = 31.25 \rightarrow 31 \text{ drops/min}$$

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/cm³** 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.)

 $d = \frac{m}{v} = \frac{18.9 \text{ g}}{1.12 \text{ mL}} = 16.9 \text{ g/m}$ No, it's not pure gold since its density is different.

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

 $\frac{0.871 \text{ g}}{\left(\frac{L}{1.29 \text{ g}}\right)} = 0.675 \quad \text{or} \quad d = \frac{m}{v} \rightarrow v = \frac{m}{d} = \frac{0.871 \text{ g}}{1.29 \text{ g/L}} = 0.675 \text{ L}$

- 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.)

$$d = \frac{225 \text{ g}}{236.6 \text{ mL}} = 0.951 \text{ g/mL}$$

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

Yes, since its density is less than that of water.

4. A cup of sugar has a volume of 237 mL. What is the mass of the cup of sugar if the density is 1.59 g/mL? (Answer is 377 grams.)

 $\frac{237 \text{ mL}}{\text{mL}} \left(\frac{1.59 \text{ g}}{\text{mL}}\right) = 377 \text{ g} \text{ or } d = \frac{\text{m}}{\text{v}} \rightarrow \text{m} = \text{v} \cdot d = (237 \text{ mL})(1.59 \text{ g/mL}) = 377 \text{ g}$

5. 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.

1 mL of $H_2O = 1.00$ g so 1,000 mL of $H_2O = 1,000$ g H_2O

1 mL of gasoline = 0.68 g so 1,000 mL of gasoline = 680 g gasoline

 A crumpet (English muffin) recipe calls for 175 grams of flour. According to Julia Child's data, the density of flour is 0.620 g/mL. How many mL of flour are needed for this recipe? (Answer is 282 mL.)

175 g $\left(\frac{1 \text{ mL}}{0.620 \text{ g}}\right)$ = 282 mL or d = $\frac{\text{m}}{\text{v}} \rightarrow \text{v} = \frac{\text{m}}{\text{d}} = \frac{175 \text{ g}}{0.620 \text{ g/mL}}$ = 282 mL

- 7. 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 c. Asphalt 1.2 g/mL Sink
 - b. Mercury 10.6 g/mL Sink d. Cork 0.26 g/mL Float
- 8. 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)

1.00 lb Hg
$$\left(\frac{454 \text{ g}}{1 \text{ lb}}\right)\left(\frac{1 \text{ mL}}{13.6 \text{ g}}\right) = 33.4 \text{ mL}$$
 or $d = \frac{m}{v} \rightarrow v = \frac{m}{d} = \frac{454 \text{ g}}{13.6 \text{ g/mL}} = 33.4 \text{ mL}$

 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)

By displacement, volume is 5.7 mL – 2.8 mL = 2.9 mL

$$d = \frac{m}{v} = \frac{32.6 \text{ g}}{2.9 \text{ mL}} = 11.2 \rightarrow 11 \text{ g/mL}$$

Temperature



$$^{\circ}C = \frac{^{\circ}F - 32}{1.8}$$
 $K = ^{\circ}C + 273$

- What is the difference between kinetic and potential energy? Kinetic energy is energy of motion. Potential energy is stored energy or energy of position.
- What are heat and temperature? How do they differ?
 Heat is the transfer of energy. Temperature is the measure of the average kinetic energy.
 With temperature, the molecules at 100 °C move faster on average molecules at 0 °C.
- 3. Temperature is a measure of _____.
 - a. average speed of the molecules
 - b. average mass of the molecules
- c. average kinetic energy of the molecules
- d. amount of heat of the molecules
- 4. Define the Fahrenheit, Celsius, and Kelvin temperature scales in terms of freezing and boiling points of water.

Water freezes at 32 °F, 0 °C, 273 K. Water boils at 212 °F, 100 °C, 373 K. °C = ((°F – 32)/(1 · 8)) K = °C + 273

- What is the mathematical relationship between Celsius and Kelvin and between Celsius and Fahrenheit scales?
 See above.
- 6. Convert –30 °C to Fahrenheit and Kelvin temperature. (–22 °F, 243 K)
 °F = 1 · 8(–30 °C) + 32 = –22 °F
 K = –30 °C + 273 = 243 K
- 7. Convert 65 °F to Celsius and Kelvin temperature. (18 °C, 291 K)
 °C = ((65 °F 32)/(1 · 8)) = 18 °C
 K = 18 + 273 = 291 K

Energy Units



CONVERSIONS

Common units of energy are joules, kilojoules, calories, Calories = kilocalories.

1. A defibrillator gives a high energy shock output of 360 J. How many calories (cal) of energy is this?

 $360 J \left(\frac{1 cal}{4.184 J} \right) = 8.6 cal$

- 2. A person uses 650 kilocalories (kcal) on a long walk.
 - a. How many joules of energy is this? **650 kcal** $\left(\frac{1,000 \text{ cal}}{1 \text{ kcal}}\right) \left(\frac{4.184 \text{ J}}{1 \text{ cal}}\right) = 2.7 \times 10^6 \text{ J}$
 - b. How many Calories is this?
 650 kcal = 650 Cal *Note: 1 Cal = 1,000 cal = 1 kcal
 - c. If the person ate a candy bar that was 230 Calories, did he/she "burn it off"? Yes!
- 3. A person burns 2100 kJ/hour swimming, and 750 kcal/hour running. Which activity burns more energy per hour?

2,100 kJ hr (1 kcal 4.184 kJ) = 502 kcal/hr

4. Sleeping burns 60 kcal/hour. In the short story "Rip Van Winkle" by Washington Irving (1819), Rip Van Winkle sleeps for 20 years. How many kcal did he burn?

20 yr
$$\left(\frac{365 \text{ d}}{1 \text{ yr}}\right)\left(\frac{24 \text{ hr}}{1 \text{ d}}\right)\left(\frac{60 \text{ kcal}}{1 \text{ hr}}\right) = 10,512,000$$
 $(1.1 \times 10^7 \text{ kcal})$

NUTRITION

Typical energy values for various food types are carbohydrates (4 kcal/g), fats (9 kcal/g), and protein (4 kcal/g).

 A nutrition label for Toll House chocolate chip cookies states for an 80-g sample (1/16 of a recipe, or about 3 cookies), there are 2.7 grams of protein, 46.6 grams of total carbohydrates, and 18.1 grams of fat. How many total Calories (kilocalories) are in this serving of cookies?

2.7 g protein
$$\left(\frac{4 \text{ kcal}}{g}\right) = 10.8 \text{ kcal}$$

46.6 g carbs $\left(\frac{4 \text{ kcal}}{g}\right) = 186 \text{ kcal}$
10.8 + 186 + 163 = 360 kcal) = 360 calories
48 Unit 2 18.1 g fat $\left(\frac{9 \text{ kcal}}{g}\right) = 163 \text{ kcal}$

NUTRITION LABEL WORKSHEET

All packaged foods are required to display a standardized nutrition label. This nutrition label contains information about the caloric content, amount of fat, protein, carbohydrates, and other required nutrients.

Examine the following nutrition labels and answer the questions.

- 1. a. What is the serving size for the food label? _____
 - b. What does this mean?_____

- 2. a. How many calories are contained within one serving of this food?
 - b. What does this mean?_____

Serving Size 1 Servings Per C	Bar (85g) ontainer 4	Fac	cts					
Amount Per Serv	ing							
Calories 170	Ca	lories fron	n Fat 50					
		% Dail	y Value *					
Total Fat 6g Saturated F Trans Fat 0	at 4g g		9% 19%					
Polyunsaturated Fat 0.5g								
Monounsaturated Fat 1g								
Cholesterol 13mg 4%								
Sodium 83mg			3%					
Total Carboh	ydrate 33g	1	11%					
Dietary Fibe	er 4q		16%					
Sugar 25g								
Protein 3g								
Vitamin A 110°	%•	Vitamin	C 2%					
Calcium 10%	•	Iron 3%	, o					
*Percent Daily Valu Your daily values n your calorie needs.	es are based o nay be higher o	on a 2,000 ca or lower depe	alorie diet. ending on					
Total Fat	Less than	<u>2,000</u> 65a	2,500 80a					
Sat Fat	Less than	20g	25g					
Cholesterol	Less than	300mg	300mg					
Sodium	Less than	2,400mg	2,400mg					
Total Carbohydrate		300g	375g					
Dietary Fiber		25g	30g					
Fat 9 •	Carbohydrate	•4 •	Protein 4					

3. How many calories would you take in if you ate the whole box of bars in one sitting?

CHEM& 121 WORKBOOK

UNIT 3 ATOMS AND ISOTOPES



Periodic Table



- Which element in row 1A has the fewest number of protons?
 Hydrogen (H)
- The element with 15 protons is in what row?
 Third
- 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 six common elements with symbols derived from their Latin name. Write the symbol of each element next to the names below
 - a. Iron Fec. Tin Sne. Mercury Hgb. Silver Agd. Copper Cuf. Lead Pb
- 4. Where are the noble gases on the periodic table? Why are they called *noble* gases? Group 8A (or 18, in modern numbering system) (last column) They are called noble gases because they are the least reactive, most inert/stable.
- Where are the alkali metals on the periodic table? Why are they called alkali metals? Group 1A They react in H₂O to form alkaline substances (high pH).
- 6. Where on the periodic table are the elements that exist as gases at standard temperature and pressure?

Noble gases (Group 8A) H_2 , O_2 , N_2 , F_2 , Cl_2

 7. How are metals different from metalloids? How do we exploit the properties of metalloids for our computing needs?
 Metalloids only have some of the properties of metals. They are semiconductors which are needed in electronics/computers.

Periodic Table Search



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



- 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

3. Identify the location of all the members

7. Identify the transition elements (B group elements).

b. alkaline earth metals

c. noble gases

d. halogens

of Group 3A.

- 8. Locate the inner transition elements.
- 9. Identify the members of the

 - b. Actinide series
- 10. Find the representative elements (main group elements, A group elements).



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.



¹H and ²H are *isotopes* of hydrogen.

¹²C and ¹³C 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. a. In terms of subatomic particles, explain how ${}^{13}C^-$ got a negative charge compared to ${}^{13}C$ (neutral, no charge).

 $^{13}C^{-}$ gained an electron, 6 protons (+) + 7 electrons (-) = overall, -1

b. In terms of subatomic particles, explain why ²³Na⁺ has a positive charge.

```
<sup>23</sup>Na<sup>+</sup> has 11 protons (+) but only 10 electrons (-).
Overall, it is +1 charge.
```

c. 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)?

```
Electrons (see Question 3)
^{13}C = 6 electrons ^{13}C^- = 7 electrons \leftarrow ion
```

- 5. Based on the model,
 - a. What do all carbon atoms and ions have in common?
 All have 6 protons.
 - b. What do all hydrogen atoms and ions have in common?
 All have 1 proton.
 - c. How many protons, neutrons, and electrons would be in ¹H⁺? (**NOTE:** ¹H⁺ is not in the model! See if you can predict what it would be ...)

```
<sup>1</sup>H<sup>+</sup> will have 1 proton, 0 neutrons, and 0 electrons.
```

d. In what way is ¹⁶O^{2–} and ²³Na⁺ similar in atomic structure? Same number of electrons. 6. Look at the periodic table. What does the *atomic number* (above each atomic symbol) represent in terms of an atom's subatomic particle?
Atomic Number --> 1

There are two possible answers; choose the BEST one! Number of protons (always ture!) Number of electrons (not always)



- 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 (atomic number).
- ¹²C and ¹³C are isotopes of carbon. ¹H and ²H are isotopes of hydrogen. What structural feature is different in *isotopes* of a particular element?
 Number of neutrons differ for various isotopes of a given element.
- 9. 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 ²H isotope based on its subatomic particles. Does that work for all the other isotopes as well? Mass number = # of neutrons + # of protons

Mass Number $\rightarrow {}^{2}$ H

10. 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.

a. Use the ¹H and ²H and ¹H⁻ (first row) in the model to determine the approximate mass of each individual sub-

1 H (average) <u>→1.0079</u> Atomic Mass

atomic particle. Note: These calculations are only to estimate—these are not accurate. Neutrons are actually heavier than protons.

A neutron weighs about 1 amu ${}^{2}H = 2.0140 \text{ amu} - {}^{1}H 1.0078 \text{ amu} = 1.0062 \text{ amu}$ An electron weighs about 0.0005 amu ${}^{1}H^{-} = 1.0083 \text{ amu} - {}^{1}H 1.0078 \text{ amu} = 0.0005 \text{ amu} (e^{-})$ A proton weighs about 1 amu

¹H = 1.0078 amu – 0.0005 amu (e⁻) = 1.0073 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 each, but e⁻ are almost 0 amu.

11. Bonus question: There are 3 isotopes of hydrogen: ¹H, ²H (deuterium), and ³H (tritium). If the average mass of all hydrogen isotopes is 1.0079 amu, which isotope is in greatest abundance (amounts) in nature?

Hydrogen (1.0079) is closest to 1, so ¹H (1 amu) is most abundant.

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 ¹⁸O atom contains the same number of protons, neutrons, and electrons.

False. ¹⁸O has 8 protons, 8 electrons, and 10 neutrons.

- **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. $Mg \rightarrow Mg^{2+} + 2e^{-}$ (HINT: Mg^{2+} or Mg^{2-} ?) b. $F + e^{-} \rightarrow F^{1-}$ c. $Al \rightarrow Al^{3+} + 3e^{-}$ (HINT: e^{-} or $2e^{-}$ or $3e^{-}$?) d. $Ca^{2+} + 2e^{-} \rightarrow Ca$
- 4. How many protons, neutrons, and electrons are found in each of the following?

b.	²³ Na ⁺ 11, 12, 10	d.	³⁵ Cl ⁻ 17, 18, 18	f.	¹⁵ N 7, 8, 7	h.	²⁷ Al ³⁺ 13, 14, 10
a.	²⁴ Mg 12, 12, 12	C.	³⁵ Cl 17, 18, 17	e.	⁵⁶ Fe ³⁺ 26, 30, 23	g.	¹⁶ 0 ^{2–} 8, 8, 10

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 but different number of neutrons.

 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

Notes: Isotopes and Atomic Mass

In 1996 a skeleton was found on the bank of the Columbia River near Kennewick, WA. Originally the bones were estimated to be from the 1800s, but carbon-14 dating indicated that the bones were 9300 years old!

What's carbon-14 dating? Let's first start with what carbon-14 is. All atoms of an element are not exactly identical.

Recall that there are three major subatomic particles present in every atom. Protons and neutrons are found in the nucleus of an atom; electrons are found outside of the nucleus, but they're still part of the atom.

All atoms of an element have the same number of **protons** in the nucleus.

But, the **isotopes** of an element are atoms that differ in the number of **neutons** in the nucleus.

Some naturally occurring isotopes of the element carbon (not drawn to scale!):



The number of protons in an atom is called the **atomic number** (denoted by the symbol **Z**). It defines the identity of the atom (which element it is).

Which of the three types of subatomic particles are considered to have significant mass? Not the electrons! The sum of the number of **protons** and **neutrons** in an atom is called the **mass number** (denoted by the symbol **A**).

How do we name an isotope of carbon?

From below: **A** = mass number (p + n)

Z = atomic number (p)

X = element symbol

Isotope Symbol Notation: AZ = 6 protons + 8 neutrons 6 = 6 protons This is also called carbon -14

An isotope of nitrogen has 8 neutrons. Determine the values for Z and A, and then write an isotope symbol.

 $_5$ N $\rightarrow \frac{15}{7}$ N z = 7, A = 15

How many protons, neutrons, and electrons are present in an atom of ²²⁶₈₈Ra ? 88 protons, 226 – 88 = 138 neutrons, 88 electrons

If an atom has a mass number of 41 and 22 neutrons, it is an isotope of which element? K 41 – 22 = 19 protons ⁴¹/₁₉? 19 = atomic number for potassium ⁴¹/₁₉K

ATOMIC MASS

How much does a carbon atom weigh? It depends, right? Are we talking about carbon-12, carbon-13, or carbon-14? Since neutrons have mass and the numbers of neutrons in isotopes are different, then the masses of the isotopes of an element will be different, too.

lsotope	Isotope Mass	% Abundance	proton 1.0073 amu neutron 1.0087 amu electron 0.00055 amu	
carbon-12	exactly 12 amu	98.938%		
carbon-13	13.0034 amu	1.078%	Combined Law of	
carbon-14	14.0032 amu	trace	Mass & Energy!	

So, what's the mass recorded on the periodic table for each element? This is a **weighted** average mass and it's called the **atomic mass**. It's calculated just like you might calculate your grade in a class.

EXAMPLE

If you earned an average of 85% on quizzes, 50% on homework, and 87% on exams, while quizzes, homework, and exams were worth 25%, 10%, and 65%, respectively, of your overall grade, then you would calculate your grade as $(.25 \times 85) + (.10 \times 50) + (.65 \times 87) = 83\%$. (Much higher than 74% from a regular average calculation!)

Let's verify the atomic mass of carbon, 12.01 amu, using the isotope data listed above. **Atomic mass = [(12 amu)0.98938] + [(13.0034 amu)0.01078] + [(14.0032 amu)0]**

"trace" (very little)

= 11.873 amu + 0.14018 amu = 12.013 amu Yes!

Weighted Averages and Atomic Mass



Use the concept and calculation of weighted average to answer the following.

1. Frank Einstein received the following report card:

Chem 100 (5 credits) 2.0 grade points (letter grade = C) Math 105 (5 credits) 3.0 grade points (letter grade = B)

Tennis (2 credits) 4.0 grade points (letter grade = A)

Explain why his GPA (grade point average) is not 3.0. Then calculate what his GPA should be.

Different number of credits is not the same for all classes.

(5/12 · 2)	0.833
+ (5/12 · 3)	+ 1.25
+ (2/12 · 4)	+ 0.667
	2.75 GPA

2. Copper has two isotopes, copper-63 and copper-65. Based on the atomic mass of copper, which isotope is most abundant?



- 3. An element exists as two different isotopes. Refer to the data table for the masses and abundances. Predict an approximate atomic mass for this element.
 - a. 85 amu c. 86 amu
- e. greater than 86 but

b. 87 amu

- d. greater than 85 but
- less than 87 amu
- less than 86 amu

Isotope Mass (amu)	Percent Abundance (%)
84.91	72.2
86.91	27.8

Review: Isotope Notation



Isotope Symbol Notation: ${}^{A}_{7}X$



Isotope Name: element-A

X = chemical symbol

Z = atomic number

A = mass number

Fill in the blanks. Assume neutral atoms.

lsotope Name	lsotope Symbol	Atomic Number	Mass Number	Number of Protons	Number of Electrons	Number of Neutrons
calcium-40	⁴⁰ 20Ca	20	40	20	20	20
magnesium-24	²⁴ 12Mg	12	24	12	12	12
mercury-201	²⁰¹ 80 <mark>Hg</mark>	80	201	80	80	121

Fill in the blanks. This table contains ions.

lon Symbol	Z	А	Number of Protons	Number of Electrons	Number of Neutrons	Charge
<mark>64</mark> 30 ²⁺	30	64	30	29	34	+
⁵⁵ ₂₅ Mn ³⁺	25	55	25	22	30	3+
³¹ 15	15	31	15	15	16	0

Balanced—overall neutral!

CHEM& 121 WORKBOOK

UNIT 4 ELECTROMAGNETIC RADIATION AND BOHR MODEL



Activity: Energy, Wavelength, and Frequency



ELECTROMAGNETIC RADIATION (LIGHT WAVES)



- Frequency (ν) 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
- Wavelength (λ) 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**

shorter

4. True or false? Higher energy light has longer wavelengths than lower energy light. False





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	
5. 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?

Red Urange Yellow Green Blue Indigo	Violet
-------------------------------------	--------

6. 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 (shorter wavelength)**

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.

	Symbol	Conversion
Coloriaa	Cal	1 Cal = 4.184 J
Calones	Udi	1 Cal = 0.001 kcal
laulaa	I	1 J = 0.0002390061 Cal
Joules	J	1 J = 0.000239006 kcal
Kilooolorio	kaal	1 kcal = 4184 J
Kilocalorie	KCal	1 kcal = 1000 Cal

EXERCISES

- Sketch a diagram of a wave and label a crest, a trough, a wavelength, and amplitude.
 Crest (peak) →
 Amplitude (height/intensity)
- 2. Define frequency of a wave. Number of cycles per unit time (usually Hertz, cycles/sec).
- 3. What are the units and symbol/abbreviations for each of these wave properties?

a.	energy <mark>E</mark>	b.	wavelength λ	C.	frequency v or f
	unit: J		unit: meter		unit: Hertz (Hz)

4. a. If the frequency (v) of a wave is increased, what happens to the energy (E) of the wave?

b. Based on your answer, which relationship is correct? (< means "proportional to")

Directionally proportional $\rightarrow (E \propto v)$ or $E \propto (1/v)$

5. a. If the frequency (v) of a wave is increased, what happens to the wavelength (λ) of the wave?



b. Based on your answer, which relationship is correct?

$$v \propto \lambda$$
 or $v \propto (1/\lambda)$



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:



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



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. (v.e.)
 - a. How many valence electrons are in Li, N, and Ne based on your drawings above (Question 9)? **1**, **5**, **8**
 - b. What group (column) are Li, N, and Ne in? 1, 5, 8
 - c. Is there a connection between the number of valence electrons and the group number? **Group number = number of v.e.**
 - d. How many valence electrons are in P, drawn in Question 3? 5
 - e. Why do you think it's similar to N? (**HINT:** Location in periodic table.) Same group/ same number of y.e.
 - f. Without drawing them out, how many valence electrons do you predict are in a sodium (Na) and an argon (Ar) atom? Na = 1 v.e. (Group 1A) Ar = 8 v.e. (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 els 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.



This is called an **electron transition**. If given enough energy, the electron can undergo an $n=1 \rightarrow n=4$ transition, as shown above.

5. What are other possible transitions if the electron absorbs less energy than shown above?

$$n = 1 \rightarrow n = 2$$
 $n = 1 \rightarrow n = 3$

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 conserved (not created/destroyed). e^- absorb energy in one process—in the reverse process, energy is reeased.

7. Only considering the four energy levels shown, what are the possible electron transitions as the atom releases energy?

$$n = 4 \rightarrow 3$$
 $n = 4 \rightarrow 2$ $n = 4 \rightarrow 1$

8. Of the transitions you listed above, which one releases the MOST energy? Explain.

4 \rightarrow 1 (greatest change/distance)

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 \rightarrow 3 (smallest energy change, red) n = 4 \rightarrow 2 (green)
```

$n = 4 \rightarrow 1$ (largest energy change, violet)

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.



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.

Radio waves, microwaves, infrared radiation, visible light, ultraviolet light, x-rays, gamma rays

- 2. Which color of light comes from a greater energy transition, red or blue? **Blue, higher E photon**
- 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 = number of v.e.
- 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 v.e.
- 5. An electron drops from $n = 4 \rightarrow 3$, and then $n = 3 \rightarrow 1$. 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 $n = 4 \rightarrow 1$? Same; conservation of E
- What is the meaning of QUANTIZED energy?
 There are only certain "chunks" of energy (quanta) absorbed or released by atoms.
- 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 all energy values would be emitted → resulting in a continuous spectrum (rainbow) rather than a line spectrum.



Calculate Your Annual Radiation Dose



Source	
1. Cosmic radiation at sea level (from outer space)	30 mrem
What is the elevation (in feet) of your town? (sea level = 30 mrem; 1650 ft = 35 mrem; 3300 ft = 40 mrem)	mrem
Airline travel: 1000 miles traveled = 1 mrem	mrem
2. Ground	
What region of the US do you live in?	
(coastal state = 23 mrem; Rocky Mountain plateau = 90 mrem; all other US regions = 46 mrem)	mrem
3. Air (radon-222)	200 mrem
4. Food and water, (e.g., potassium)	40 mrem
5. Building materials (brick = 7, wood = 4, concrete = 8)	
Your house	mrem
Place of work/school	mrem
6. Medical and dental diagnostics and other	
Dental x-rays (10 mrem per visit)	mrem
Do you wear a plutonium powered pacemaker? (100)	mrem
How many medical x-rays do you receive per year? (20 mrem each)	mrem
Do you have porcelain crowns or false teeth? (0.7)	mrem
Nuclear medical procedures (14 mrem each)	mrem
7. Nuclear Weapons test fallout	1 mrem
8. Other/Misc	
Are x-ray luggage inspection machines used at your airport? (0.002)	mrem
Do you use gas lantern mantles when camping? (0.003)	mrem
Do you wear a luminous wristwatch (LCD)? (0.06)	mrem
Do you watch TV? (1)	mrem
Do you use a computer terminal? (0.1)	mrem
Do you have a smoke detector in your home? (0.008)	mrem
9. Do you live within 50 miles of a nuclear power plant? (0.009)	0 mrem
10. Do you live within 50 miles of a coal fired power plant? (0.03)	0 mrem
GRAND TOTAL =	mrem

From https://www.epa.gov/radiation/calculate-your-radiation-dose

NOTES: The amount of radiation exposure is usually expressed in a unit called millirem (mrem). In the United States, the average person is exposed to an effective dose equivalent of approximately 360 mrem (whole-body exposure) per year from all sources (NCRP Report No. 93).

The dose calculator is based on the American Nuclear Society's brochure, "Personal Radiation Dose Chart." The primary sources of information we relied on are the National Council on Radiation Protection and Measurements Reports #92–95 and #100. Please remember that the values used in the calculator are general averages and do not provide precise individual dose calculations. CHEM& 121 WORKBOOK

UNIT 5 NUCLEAR CHEMISTRY





PART 1: TYPES OF RADIATION

Substances that are radioactive spontaneously release high energy particles or rays (radiation). After a substance releases radiation, it becomes more stable (or less radioactive). There are three types of radiation that may be released—alpha, beta, and gamma. Their properties are shown below.

Type of Radiation	Symbol	Mass (amu)	Charge	Penetrating Power	Can Be Stopped By
Alpha	⁴ ₂ He or ⁴ ₂ α	4 amu	+2	low	Paper, clothing, air
Beta	_1 ⁰ e or_1 ⁰ β	0 amu	-1	intermediate	Sheet of metal or wood
Gamma	0 0Y	0 amu	0	high	Several inches of lead or concrete

Electron capture: A type of radioactive decay where the nucleus of an atom absorbs a K or L shell electron and converts a proton into a neutron.

The decay scheme for electron capture is:

$$^{Z}X_{A}+e^{-}\rightarrow ^{Z}Y_{A-1}+\nu +\gamma$$

Positron: Positron emission or beta plus decay (β^+ **decay**) is a subtype of radioactive decay called beta decay, in which a proton inside a radionuclide nucleus is converted into a neutron while releasing a positron.

Example of positron decay:

$$^{23}_{12}Mg \rightarrow {}^{23}_{11}Na + {}^{0}_{1}e + {}^{0}_{0}v$$

- Discuss what the symbols mean in the table. What does the top number represent? The bottom? How do these symbols differ from the isotope labels you used before?
 The top number represents the mass and the bottom represents the number of protons.
- Which type of radiation is the heaviest in terms of mass? Which is the lightest? Heaviest: alpha Lightest: beta and gamma
- 3. If you were given three radioactive rocks (one that emitted each of the three kinds of radiation) and you were forced to keep two of them, which one would you give away, and why? How would you minimize your exposure to the rocks you have to keep?

Keep alpha and beta because they have lower penetrating power than gamma. Increase its distance from you and/or sheild yourself with lead.

PART 2: NUCLEAR EQUATIONS

What causes some substances to be radioactive while others are not? An atom is radioactive when its nucleus is too large or has an imbalance of neutrons and protons. When the nucleus is unstable, it releases radioactive particles (α , β , and/or γ) and becomes a new nucleus.

1. When an atom's nucleus changes, does the atom become a different element? Explain. Yes—the number of protons changes.

The following is an example of what happens when uranium-238 undergoes alpha decay:

Equation 1
$${}^{238}_{92}U \rightarrow {}^{234}_{90}Th + {}^{4}_{2}\alpha$$

Parent daughter alpha particle (decay)

The arrow divides what we start with before decay (on the left) and what we end up with after decay (on the right).

2. Verify whether your answer to Question 4 above is correct, based on Equation 1 given above.

```
92 \rightarrow 90 protons
```

- $\mathbf{U} \rightarrow \mathbf{Th}$
- 3. Explain why the alpha particle is shown on the right side of Equation 1 instead of the left side.

The alpha particle is produced and ejected from nucleus.

- 4. The "parent nuclide" is the original isotope that produces "daughter nuclides." Label the parent and daughter nuclides in Equation 1.
- Compare the total mass of everything on the left side of Equation 1 with the right side of Equation 1. What do you notice?
 Conservation of mass.
- 6. Balance the following nuclear equations by providing the missing daughter nuclide.

a.
$$^{224}_{88}$$
Ra $\rightarrow ^{4}_{2}\alpha + ^{220}_{86}$ Rn

b.
$$^{249}_{98}Bk \rightarrow ^{0}_{1}B + ^{249}_{98}Cf$$

c. $^{145}_{60}$ Nd + $^{0}_{-1}e \rightarrow$

NOTE: Part B involves beta decay. This is possible because a neutron can release a beta particle and a proton.

Fill in the blanks with "increases" or "decreases": During alpha decay, the mass number decreases by four and the atomic number decreases by two. During beta decay, the mass number doesn't change, but the atomic number increases by one.
 A neutron is decaying to a proton and a beta particle.

- 8. Write a balanced nuclear equation for the following processes.
 - a. Americium-241 (²⁴¹Am) is in smoke detectors and undergoes alpha decay. $^{241}_{95}Am \rightarrow ^{237}_{93}Np + ^{4}_{2} \propto$ (alpha)
 - b. Positron emission tomography (PET) often uses the positron emission of fluorine-18. ${}^{18}_{9}F \rightarrow {}^{0}_{+1}\beta + {}^{18}_{8}O$ Positron ${}^{0}_{+1}\beta$

PART 3: HALF LIFE

Just as different radioactive samples emit different types of radiation, they also emit radiation at different rates. When a radioactive sample emits radiation, it becomes more stable over time. This process is called radioactive decay. Some isotopes decay slowly and others quickly. The rate at which radiation is emitted (also called the rate of decay) is called the half-life.

DEFINITION

half-life = the time required for half the mass of a radioactive isotope to decay

- 9. Suppose you start with 1 gram of a radioactive sample. How many grams do you have after
 - a. one half-life? 0.5 g c. three half-lives? 0.125 g e. five half-lives? 0.03125 g
 - b. two half-lives? 0.25 g d. four half-lives? 0.0625 g f. ten half-lives? 0.000977 g (~0 g)
- 10. A radioactive sample is composed of 60 milligrams of iodine-131, which has a half-life of 8.0 days. How many milligrams of iodine-131 are left after 24 days?

```
60 mg \rightarrow 30 mg \rightarrow 15 mg \rightarrow 7.5 mg after 3 half lives
```

24 days $\left(\frac{\text{half life}}{8.0 \text{ days}}\right) = 3.0 \text{ half lives}$

11. Radioactive carbon-dating uses the principles of nuclear decay to find out how old an object is (assuming the object was or was not part of a living organism, and is not much older than 50,000 years). The half-life of carbon-14 is 5730 years. A mammoth skeleton has a carbon-14 content 12.50% of that found in living organisms. How old is the skeleton? (HINT: How many half-lives have passed to have 12.5% of a sample remain? See Question 12.)

100 \rightarrow 50 \rightarrow 25 \rightarrow 12.5 (3 half lives have passed)

76 Unit 5 $3 \text{ half lives} \left(\frac{5,730 \text{ yrs}}{1 \text{ half life}}\right) = 17,190 \text{ yrs}$

Balancing Nuclear Equations



- 1. Draw the symbol of each type of radiation and include its mass and charge.
 - a. Alpha $\frac{4}{2} \propto$ c. Gamma $\frac{0}{0} \gamma$
 - b. Beta $_{1}^{0}\beta$ d. Positron $_{+1}^{0}\beta$
- 2. Which isotope is produced when krypton-81 undergoes beta decay?

Write the balanced nuclear equation. ${}^{81}_{36}$ Kr $\rightarrow {}^{0}_{-1}\beta + {}^{81}_{37}$ Rb

3. By what decay process does thorium-230 decay to radium-226?

Write the balanced nuclear equation. $^{230}_{90}$ Th $\rightarrow ^{226}_{88}$ Ra + $^{4}_{2} \propto$ (alpha decay)

- 4. Positron emission tomography (PET) is a medical imaging technique that uses positron emitters, such as oxygen-15.
 - a. What type of radiation is *detected* with PET?
 Positron _1⁰β
 - b. Write the equation for the positron decay of oxygen-15. ${}^{15}_{8}0 \to {}^{0}_{-1}\beta + {}^{15}_7N$

Half-Life

- Hyperthyroidism (an overactive thyroid) is sometimes treated by using radioactive iodine, ¹³¹₅₃I, which shrinks the thyroid gland and lessens symptoms. It has a half-life of 8.02 days.
 - a. A patient is instructed to take 100 μ g of ¹³¹I. If the patient only takes one dose, approximately how much iodine-131 remains in his/her body after 48 days?

48 days $\left(\frac{\text{half life}}{8 \text{ days}}\right) \sim 6$ half lives; 100. $\mu g \times \left(\frac{1}{2}\right)^6 - \text{ or divide by 2 six times} = 1.56 \ \mu g$

b. About how many days will it take for the 100- μg sample to decay to an activity below 2.0 $\mu g?$

100 $\xrightarrow{1}$ 50 $\xrightarrow{2}$ 25 $\xrightarrow{3}$ 12.5 $\xrightarrow{4}$ 6.25 $\xrightarrow{5}$ 3.125 $\xrightarrow{6}$ 1.563

```
About 6 half lives ... 8 days × 6 half lives/day = 48 days
```

2. The decay rate of the carbon in an old bone is one-eighth that of the carbon in a living organism. The half-life of carbon-14 is 5730 years. How old is the bone?

1/8 is how many half lives?

 $1 \rightarrow \frac{1}{2} \rightarrow \frac{1}{4} \rightarrow \frac{1}{8}$ (3 half lives × ((5,730 yr)/(half life)) = 17,190 yrs

- 3. Carbon-14 dating is most reliable for objects between 1000 and 50,000 years. For objects older than 50,000 years, the decay rate of uranium-238 can be used. Uranium-238 decays to lead-206 and has a half-life of 4.46×10^9 years. The age of a rock is the time elapsed since the rock solidified. Suppose a rock from a planet in another solar system was obtained, and it was determined that 75% of the uranium-238 originally present had decayed to lead-206.
 - a. What is the approximate age of the rock?

2 half lives × $\left(\frac{4.46 \times 10^9 \text{ yrs}}{\text{half life}}\right)$ = 8.92 × 10⁹ yrs

b. Our solar system formed around 4.5 billion years ago. Based on the result in part a (above), is the other solar system younger or older than ours?
 Older, by about 4.5 billions years! (4.5 × 10⁹)

CHEM& 121 WORKBOOK

UNIT 6 COVALENT BONDS AND STRUCTURE OF MOLECULES



Covalent Nomenclature 1



To determine if the compound is *covalent (molecular)*, look for either

- a nonmetal bonded to a nonmetal or a metalloid bonded to a nonmetal.
- To name, use Greek prefixes to indicate the number of atoms present in the compound. Prefixes are listed in your text book. Metalloids are also named in the same format as molecular compounds.

Number	Prefix	Number	Prefix
1	mono-	6	hexa-
2	di-	7	hepta-
3	tri-	8	octa-
4	tetra-	9	nona-
5	penta-	10	deca-

NOTE: Mono is usually omitted for the first element.

EXAMPLES

• CO_2 carbon dioxide • N_2O_3 dinitrogentrioxide • CO carbon monoxide

To write the formula, use subscripts to identify the number of atoms present.

EXAMPLES

• dinitrogen tetraoxide N_2O_4 carbon tetrachloride CCl₄

Some other traditional names you need to memorize.

• water H_2O ammonia NH₃ ■ methane CH₄

PRACTICE PROBLEMS FOR COVALENT COMPOUNDS

- 1. Write the names of the following binary covalent compounds: (Remember to use prefixes here.)
 - a. SO₂ Sulfur dioxide
 - b. PCl₃ Phosphorus trichloride
 - c. NO₂ Nitrogen dioxide
- 2. Write the formulas for the following:
 - a. Phorphorus pentachloride PCI₅
- b. lodine trichloride ICI₃

- d. H_2S Hydrogen sulfide NOTE: For H in front we do not need to use prefixes (exception).
- e. SF₆ Sulfur hexafluoride
- f. CBr₄ Carbon tetrabromide
- c. Dinitrogen tetraoxide N₂O₄
- d. Diarsenic pentasulfide As₂S₅

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Covalent Nomenclature 2



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. Therefore, you simply have to state the chemical formula using words and prefixes, and continue using binary endings.

EXAMPLE

SiCl₄ is silicon tetrachloride. (**NOTE:** Mono is usually omitted from the first word.)

1.	XeO ₃ Xenon trioxide	8. dihydrogen monoxide <mark>H₂O</mark>	
2.	OF ₂ Oxygen difluoride	9. xenon tetrafluoride XeF₄	
3.	SO ₃ Sulfur trioxide	10. diiodine pentoxide <mark>I₂0</mark> 5	
4.	P ₂ O ₅ Diphosphorus pentoxide	11. chlorine monoxide ClO	
5.	Cl ₂ 0 Dichlorine monoxide	12. carbon tetrafluoride CF₄	
6.	NI ₃ Nitrogen triiodide	13. tetraphosphorus triselenid	e <mark>P₄Se</mark> 3
7.	SCI ₂ Silicon dichloride	14. phosphorus pentabromide	PBr ₅



The valence electrons are the outermost electrons in an atom. These are the electrons stored in the highest (*n*) occupied energy level (shell) that is furthest from the nucleus. For main group elements, the group number (1A to 8A) indicates the number of valence electrons for an atom in that group. Electrons that are not valence electrons are core electrons.

EXAMPLES

- Carbon is in Group 4A, so it has 4 valence electrons. $\dot{\mathbf{C}}$ Lewis structure
- Carbon has 2 electrons in n = 1, 4 electrons in n = 2; n = 2 is highest energy level, so C has 4 valence electrons.
- Bromine has 35 total electrons. It's in Group 7A so it has 7 valence electrons.

PRACTICE

How many valence electrons does an atom of sulfur have? 6 v.e. (group 6A) ·Š·

How many total electrons does an atom of calcium have? How many core electrons? How many valence electrons? **Ca has 20 e⁻ total, 2 v.e. ·Ca**·

ATOMS BOND TO ACHIEVE FULL VALENCE SHELLS—"THE OCTET RULE"

Atoms of main group elements will undergo chemical reactions via transfer (gain or loss) of electrons or sharing of electrons in order to achieve full valence shells that result in energetic stability. A filled valence shell for a main group element has 8 valence electrons, hence "The Octet Rule." The exception is for hydrogen or helium that only need a maximum of 2 valence electrons to fill n = 1 energy level.

Transfer of electrons involves ionic bonding and results in ionic compounds; sharing of electrons involves covalent bonding and results in covalent molecules.

For the transfer of electrons to form ionic compounds, metal atoms tend to lose electrons, while nonmetal atoms gain electrons.

Elements in the same group tend to have similar chemical properties because atoms in the same group will react to bond so that a filled valence shell is achieved (Octet Rule followed).

PRACTICE

How many electrons would an atom of sulfur gain to achieve an octet? 2 e⁻

How many electrons would an atom of calcium lose to achieve an octet?

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, CI, 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

$$NO_2^-$$
 [: $\ddot{\mathbf{0}} - \ddot{\mathbf{N}} = \ddot{\mathbf{0}}$:]²⁻
 $N = 5e^-$
 $O = 6e^- - 2 = 12e^-$
 $17e^- + 1e^-(-) = 18e^-$

Lewis Dot Structures



Draw the Lewis dot structure for each of these substances.

NF₃
$$SO_4^{2-}$$

5 + 3(7) = 26 v.e. 6 + 4(6) + 2 = 32 v.e.
 $\stackrel{!}{:}\overset{'}$



$$NH_4^+$$

5 + 4(1) - 1 = 8 v.e.
 $\begin{bmatrix} H \\ H - N - H \\ H \end{bmatrix}^+$

SiCl₄
4 + 4(7) = 32 v.e.
$$\ddot{C}l^{:}$$

 $\ddot{C}l - Si - \ddot{C}l^{:}$
 $\ddot{C}l^{:}$

SCl₂ 6 + 2(7) = 20 v.e. :<u>ĊI - Ŝ - Ċ</u>I: C_2H_4 4(2) + 4(1) = 12 v.e. $H_H > C = C < H_H$

$$PH_3$$
 $\underline{C}H_2O$
 $5 + 3(1) = 8$ v.e.
 $4 + 2(1) + 6 = 12$ v.e.

 $H - \ddot{P} - H$
 0 :

 $H - \ddot{P} - H$
 $H - C - H$

$$NO_3^-$$

5 + 3(6) + 1 = 24 v.e.
 $\begin{bmatrix} \ddot{O} = N \\ \ddot{O} \vdots \end{bmatrix}^-$

SO₃ 6 + 6(3) = 24 v.e. ;Ö: , S ©: ;Ö[~]S ö:

0 ₃	CO
6 × 3 = 18 v.e.	4 + 6 = 10 v.e.
:Ö – Ö = Ö	:C ≡ 0:

$$CH_3OH$$
 $CHCl_3$
 $4 + 3(1) + 6 + 1 = 14$ v.e.
 $4 + 1 + 3(7) = 26$ v.e.

 H
 $\stackrel{!}{Cl}$
 $H - \stackrel{!}{C} - \stackrel{!}{O} - H$
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$
 H
 $\stackrel{!}{Cl}$

$$CO_2$$
 N_2H_2
 $4 + 6(2) = 16$ v.e.
 $5(2) + 2(1) = 12$ v.e.

 $\ddot{O} = C = \ddot{O}$
 $H - \ddot{N} = \ddot{N} - H$

$$H_{3}O^{+}$$

 $3(1) + 6 - 1 = 8 \text{ v.e.}$
 $\begin{bmatrix} H \\ 0 \\ H \end{bmatrix}^{+}$

Notes: Bonding in Molecules



HOW TO DRAW LEWIS DOT STRUCTURES (2-D DRAWINGS—SHOW ALL PAIRS OF ELECTRONS)

TYPICAL BONDING PATTERNS FOR ATOMS IN ORGANIC COMPOUNDS

(single/double/triple bonds and lone pairs)



Summarize the pattern for each element:

Element	# of Bonds	# of Ion e ⁻ Pair
С	4	0
Ν	3	1
0	2	2
H and Halogens	1	0

These patterns are similar for elements in the same group of the periodic table.

- What is the typical bonding pattern for P? (How many bonds and how many lone pairs?)
 P is like N, so 3 bonds + 1 lone pair
- 2. Fill in the missing electrons (lone pairs or bonds) in this molecule.



CENTRAL VS. SURROUNDING (OUTER, TERMINAL) ATOMS

The molecule shown above is a larger molecule, with many central atoms and some surrounding atoms. A central atom is an atom that is bonded to two or more atoms. A terminal atom is only bonded to one other atom.

3. Circle all of the central atoms in this molecule.



ELECTRON GROUPS

Count the # of regions of electrons around a central atom.

An electron group is a region of electrons, either pair(s) of bonding electrons or lone pairs.

- 1 lone pair = 1 region of electrons
- 1 single bond = 1 region of electrons
- 1 double bond = 1 region of electrons
- 1 triple bond = 1 region of electrons

Question: How many electron groups are around each central atom in this molecule? = 2e = 3e = 4e Electron groups repel each other and spread out in 3-D space (VSEPR model)



ELECTRON GEOMETRY (OR ELECTRON ARRANGEMENT)

The electron geometry is the name that corresponds with the number of electron groups. It's a name that describes how many regions of electrons there are around a central atom and how they are arranged in 3-D space. This includes both lone pairs and bonds to atoms (doesn't matter if single, double, or triple—each counts as one electron region). Ask yourself: how would the atoms and lone pairs on a central atom be spread out in 3-D space?

- 2 electron groups = linear electron arrangement (180°)
- 3 electron groups = trigonal planar arrangement (120°)
- 4 electron groups = tetrahedral arrangement (109.5°)

BOND ANGLES

What are the angles between these groups?

DRAW A 3-D SKETCH FOR A MOLECULE

Once you've figured out the # of electron groups and electron geometry around each central atom in a molecule, draw the molecule in 3-D, using these generic templates. Replace the letters with the actual element symbols in the molecule.



Electron Geometry and Molecular Shape

The molecular shape is the name that describes how only the atoms are arranged in 3-D space. Look at the atoms (not the lone pairs) and choose the name that matches their arrangement.

When a central atom has no lone pairs on it, the names for the electron geometry and molecular shape are the same. When there are lone pairs, the molecular shape name will be different.

# Atoms Bonded to Central Atom	# Lone Pairs on Central Atom	# Electron Groups around Central Atom	Electron Geometry	Bond Angles	Molecular Shape	Linear 3- DiDeav ing
4	0	4	tetrahedral	109.5°	tetrahedral	
3	1	4	tetrahedral	109.5°	trigonal pyramidal	
2	2	4	tetrahedral	109.5°	bent	x Ä:
3	0	3	trigonal planar	120°	trigonal planar	X X X X
2	1	3	trigonal planar	120°	bent	x A x
2	0	2	linear	180°	linear	X X

1. Fill in the missing blanks.

2. Label the molecular shape at each central atom in this molecule.

What are the bond angles for the carbons in the hexagonal ring?
 Trigonal planar + 120°

Trigonal planar \rightarrow 120°

4. What are the bond angles for the oxygens with the two lone pairs?





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				
нйн	c=:0;	[:0N=0:]		
Geometry: Tetrahedral	Geometry: Linear	Geometry: Trigonal planar		
Shape: <mark>Bent</mark>	Shape: Linear	Shape: Trigonal planar		
C≡ H	:сі: н—с—сі: н	: <u>ö</u> — <u>s</u> = <u>ö</u> :		
Geometry: Linear	Geometry: Tetrahedral	Geometry: Trigonal planar		
Shape: Linear	Shape: Tetrahedral	Shape: <mark>Bent</mark>		
F F	:;; c=;;:	:CI: :CI-P-CI:		
Geometry: Tetrahedral	Geometry: Linear	Geometry: Tetrahedral		
Shape: Bent	Shape: Linear	Shape: Trigonal pyramidal		
H H H H	H H H H H	о: ∥ н—с—н		
Geometry: Tetrahedral	Geometry: Tetrahedral	Geometry: Trigonal planar		
Shape: Trigonal pyramidal	Shape: Tetrahedral	Shape: Trigonal planar		
	:ö—-ö==0:	H—Äs—H H		
Geometry: Trigonal planar	Geometry: Trigonal planar	Geometry: Tetrahedral		
Shape: Trigonal planar	Shape: Bent	Shape: Trigonal pyramidal		

Review: VSEPR



1. True or False? A molecule with three atoms bonded to the center can have a tetrahedral shape. **Explain.**

It can have $4 e^-$ groups (tetrahedral geometry) but it will have trigonal pyramidal shape if bonded to 3 atoms (will have a lone pair as well). $\cdots \leftarrow$ One lone pair

Ar ← Three atoms
 Provide the possible electron geometries and molecular shapes for the following bond angles:

		-			
			109.5°	120°	180°
	e^-	geometry	Tetrahedral	Trigonal planar	
	Mo	olecular shape	Tetrahedral, trigonal	Trigonal planar, ben	t
3.	Dra	aw the Lewis str	ucture for formaldehyde	, <u>С</u> Н ₂ О.	:0:
	a.	There is (are) <mark>2</mark>	single bond(s) and <mark>1</mark> dou	ible bond(s) in <u>C</u> H ₂ O.	Ĩ
	b.	There is (are) Ø	lone pair(s) on the cent	ral atom for <u>C</u> H ₂ O.	H´Ŭ`H
	C.	What is the bon	nd angle involved in this	molecule? Circle one:	109.5° 120° 180°
	d.	<u>C</u> H ₂ O has a tr i lar shape.	i <mark>gonal planar</mark> electron	geometry and a tri	<mark>gonal planar</mark> molecu-

4. a. Can a molecule with a tetrahedral geometry have a bent shape? **Explain.**

Yes, if it has two atoms and two lone pairs on the central atom.

b. Can a molecule with a trigonal planar geometry have a trigonal pyramidal shape? **Explain.**

No, this is a tetrahedral geometry (not trigonal planar).

c. Can a molecule with a linear geometry ever have a bent shape? Explain.
 No! Linear geometry only leads to linear shape.

Electronegativity

Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons. Use the Electronegativity Scale to answer the questions below.

- What happens to the electronegativity values across each period from left to right?
 Generally increases
- 2. What happens to the electronegativity values of each group from bottom to top?

Generally increases

 If you have a bond between a metal atom and a nonmetal atom, which of the two is more electronegative? Explain your thinking.



Nonmetals appear to be much more electronegative than metals.

- Where are the atoms with the greatest electronegativity values located? Are they metals, nonmetals, or metalloids?
 Nonmetals. Upper right corner of periodic table.
- Where are the atoms with the lowest electronegativity values located? Are they metals, nonmetals, or metalloids?
 Metals. Lower left corner of the periodic table.
- 6. Metals often are referred to as <u>electropositive</u>. Explain why. **They are opposite of electronegative**.
- 7. Why do you think the noble gases do not have electronegativity values? They do not tend to share e⁻ (bond) with other elements (valence shells are complete).
- 8. Circle the atom in each pair below that will attract shared electrons more strongly.

a.	С	or Cl	c. I or In	e.	As	or Na
b.	Rb	or Br	d. Ag or S	f.	Н	or Se

9. Which two atoms in the periodic table form the most polar bond?

Lower left (lowest EN) \rightarrow Fr—F \leftarrow Upper right (highest EN)

- 10. List at least three examples of pairs of atoms with nonpolar covalent bonds.
 H—H F—F 0—0
- 11. If the difference in electronegativity between two bonded atoms is greater than 2.1, then the bond is considered ionic. List three examples of pairs of atoms with ionic bonds. Difference in EN = Δ EN LiF NaCl K—0 Δ EN = |1-4| = 3 Δ EN = |0.9-3| = 2.1 Δ EN = |0.8-3.5| = 2.7
- 12. If the difference in electronegativity between two bonded atoms is less than 2.1, then the bond is considered polar covalent. List three examples of pairs of atoms with polar covalent bonds.

C--ON--OS--F $\Delta EN = |2.5 - 3.5| = 1$ $\Delta EN = |3 - 3.5| = 0.5$ $\Delta EN = |2.5 - 4| = 1.5$

- 13. Metal atoms tend to form cations with positive charges. Is this consistent with the electronegativity of metal atoms? Why or why not?
 Yes, they are not as attracted to e⁻ (electropositive) K⁺ → K⁺
- 14. Provide two explanations for why nonmetals tend to form anions.
 Anions are negatively charged ions.
 1. Gain e⁻ to complete valence shell.
 2. Are more EN so are attracted to e⁻.
- Sulfur forms both ZnS and SF₂. Is sulfur the most electronegative element in both compounds? Why or why not?
 ZnS has more EN. SF₂ has less EN. (Based on values of EN in comparison to other element.)
- 16. Arrange these bonded pairs in order of increasing polarity from the least polar to the most polar: C—H, H—O, N—H, and H—F.

 $\begin{array}{ccccc} \textbf{C--H} & \textbf{N--H} & \textbf{O}--\textbf{H} & \textbf{H}--\textbf{F} \leftarrow \text{ most polar} \\ \Delta \textbf{EN} & 0.4 & 0.9 & 1.4 & 1.9 \end{array}$

17. Explain how you would use the electronegativity scale to determine both the direction and the degree of polarity of a bond between two different atoms.
 Ex. C-0 or C-0

For direction: the dipole points in the direction of the atom with a higher EN value.

For degree of polarity: The greater the difference in EN values, the more polar the bond.

18. Which of these pairs of atoms would result in the most polar bond? _____

The least polar bond?		Arr	ange the	ese bond	led pair:	s in ord	ler of
increasing polarity from the lea	ast polar	r to the	most po	lar: C—H	I, C—S,	H—F, C	C—N,
C—O, and H—Br. Least polar	C—S	С—Н	C—N	H—Br	C—0	H—F	Most polar
	2.5	2.5	3.0	2.8	3.5	4.0	
	- 2.5	- 2.1	- 2.5	- 2.1	- 2.5	<u> </u>	
	Ø	0.4	0.5	0.7	1.0	1.9	

Activity: Bond Polarity and Molecular Polarity

PART 1: BOND POLARITY

Each kind of atom has a certain attraction (or greediness) for the electrons in the bond they share.

If two atoms in a covalent bond are the same (0-0), the attraction for the electrons in the bond is the same and the electrons are evenly shared between the two atoms. We call this a **NON-POLAR bond** (0.0–0.4 \rightarrow pure covalent (non-polar)).

If the atoms are different (one X and one Y), then the attraction is different, and the electrons are not evenly shared (such that one end is more positively charged than the other, and one end is more negatively charged. We call this a **POLAR bond** (0.5–1.9 \rightarrow polar covalent). (Like two poles of a magnet—opposites!)

H – C ≡ N:

- Draw the Lewis Dot Structure and/or build a model of HCN. Note that it has a linear electron geometry.
 C—H |2.55 2.20| = 0.35
 - a. Is the H—C bond in the molecule POLAR or NON-POLAR? Less than 0.5 so "non-polar"
 - b. Is the C—N bond in the molecule POLAR or NON-POLAR? [2.55 3.04] = 0.49 ~ 0.5

In a polar bond, which atom is more greedy? It's the one that is more **ELECTRONEGATIVE** (has a higher electronegativity value)(>2 \rightarrow ionic).



Figure 2. Pauling electronegativity values for selected elements.

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- 2. Does the electronegativity value tend to increase or decrease from left to right?
- 3. Does the electronegativity value tend to increase or decrease from bottom to top?
- Which element in the periodic table has the highest electronegativity value? F The lowest electronegativity value? Cs (in the whole table)

K (in the partial table shown above)

The atom with the higher electronegativity value (see table above) is more greedy for the electrons in the bond. This greediness creates an unevenly shared polar bond. The greedy atom becomes slightly negatively charged, δ -. (The δ sign means "partial.") The less greedy atom becomes δ +.

- 5. In HCN, which atom is more electronegative? C or N
- 6. Which best describes the distribution of electrons in the CN bond? Explain briefly.



The more EN atom pulls the e⁻ toward itself, resulting in more negative charge.

7. The second one is correct. This distribution of electrons is called a dipole. Another way to draw it is like this: $\overrightarrow{C-N}$

PART 2: MOLECULAR POLARITY

- 8. Draw the Lewis Dot Structure and/or build a model of CO₂. Does this molecule contain polar bonds? Yes ⁱ = C = Ö
- 9. Even though this molecule contains polar bonds, the molecule is symmetrical. The polarity cancels out and the molecular polarity is zero (nonpolar). Draw a picture of CO₂ and the two dipoles. See how the dipoles both point in opposite directions? (NOTE: Dipoles both pointing inward would result in the same thing—zero polarity.)
- 10. A good rule of thumb: If a molecule doesn't contain lone pairs on the central atom (non-bonding electrons), the dipoles cancel and it is nonpolar overall. If a molecule contains lone pairs on the central atom, the dipoles do not cancel and the molecule is polar.

Complete the table below for each of these shapes with their (1) electron geometry, (2) molecular shape, and (3) molecular polarity (polar or nonpolar).



11. True or False? A molecule can be non-polar even though it contains polar bonds. Explain. The dipoles could cancel each other out, which occurs if the molecular shape is highly symmetrical (no lone pairs on the central atom).



ELECTRONEGATIVITY (EN)

Electronegativity (EN) is a measure of an atom's ability to attract valence electrons when bonding to another atom. The more electronegative an atom is the greater the attraction for the valence electrons and the higher the electronegativity value. Electronegativity is based on a 4.0 scale. Fluorine has the highest electronegativity of the elements at 4.0.

Sketch: Draw a rough sketch of the outline of the Periodic Table here and label the trends in EN across periods and down groups. Use your textbook or the internet to look up EN values.

TYPE OF BONDING BASED ON DIFFERENCE IN ELECTRONEGATIVITY BETWEEN TWO ATOMS

Calculate the difference in electronegativity (ΔEN) of the two atoms involved in a bond. If there is no difference or a slight difference, the bond is a nonpolar covalent bond; if there is moderate difference, the bond is polar covalent bond; and if there is a large difference, the bond is an ionic bond, indicating transfer (gain or loss) of electrons instead of sharing of electrons.

Here are the ranges listed in your textbook:

• $\Delta EN = 0$ to 0.4 is nonpolar covalent • $\Delta EN = 0.5$ to 1.9 is polar covalent • $\Delta EN = 2.0 +$ is ionic

In reality, chemical bonding is a continuum of polarity, and these ranges are only one perspective. In practice, many chemists and scientists don't calculate ΔEN values to determine bond polarity, but instead use rough guidelines. The types of elements (metal, semimetal, nonmetal) in a compound can be used for a quick method of distinguishing between covalent and ionic bonding (see more in another section below), but doesn't help distinguish between polar covalent and nonpolar covalent bonds.

BOND POLARITY: NONPOLAR COVALENT BONDS VS. POLAR COVALENT BONDS

Sort these bonds into three major groups based on their ΔEN . Bonds to sort: C—O , C—H, C—N, N—H, C—C, O—H.

nonpolar covalent C—C, C—H polar covalent C—O, C—N very polar covalent 0—H, N—H

Notes: Molecular Polarity



A polar molecule is a molecule that has areas of slightly negative charge and areas of slightly positive charge as the result of unequal distribution of valence electrons across the whole molecule. The molecule can be thought of as having a + "pole" and a – "pole," like the Earth has a N pole and a S pole.

A nonpolar molecule has non imbalance in the sharing of valence electrons (no "poles").

The polarity of molecules is important because it determines how molecules interact ("stick") with other molecules. Same premise again here, opposite charges attract. These interactions between molecules are called intermolecular forces.

HOW TO DETERMINE MOLECULAR POLARITY

A molecule must have polar bonds and an asymmetric shape to lead to an imbalance in electron distribution and therefore be POLAR.

- 1. Draw the Lewis Dot Structure for the molecule.
- 2. Count # of electron groups around central atom and determine electron geometry.
- 3. Draw molecule in 3-D to show bond angles and lone pairs on central atom.
- 4. Draw bond dipole arrows next to each polar bond.
- 5. If the bond dipoles do not cancel out in 3-D space and there is a net "pull" of the electrons to one side of the molecule, then the molecule is POLAR.

Is CH₂F₂ polar or nonpolar?



Lewis Dot Structures, Geometry, and Polarity



Formula of Compound	Lewis Dot Structure	Electron Geometry	Molecular Shape	Bond Angle	Polar?
Cl ₂ chlorine 7 + 7 = 14 v.e.	:Čl — Čl:	Linear	Linear	180°	Non-polar
H ₂ O water 2(1) + 6 = 8 v.e .	H – Ö – H or H – Ö – H	Tetrahedral	Bent	109.5°	Polar
CH ₄ methane 4 + 4(1) = 8 v.e .	$H = \begin{bmatrix} H & \text{or} & H \\ I & -H & H \end{bmatrix} = \begin{bmatrix} H & H \\ I & H \end{bmatrix}$	Tetrahedral	Tetrahedral	109.5°	Non-polar
NH ₃ ammonia 5 + 3(1) = 8 v.e .	H [∕] Ń∖ H H	Tetrahedral	Trigonal pyramidal	109.5°	Polar
HF hydrogen fluoride 1 + 7 = 8 v.e .	+→ H−Ë:	Linear	Linear	180°	Polar
N ₂ nitrogen 5 + 5 = 10 v.e .	:N ≡ N:	Linear	Linear	180°	Non-polar

Formula of Compound	Lewis Dot Structure	Electron Geometry	Molecular Shape	Bond Angle	Polar?
CO ₂ carbon dioxide 4+6(2) = 16 v.e.	⇔ c=c=ö	Linear	Linear	180°	Non-polar (symmetric)
<u>C</u> ₂ H ₂ acetylene 4(2) + 2(1) = 10 v.e.	Two central atoms H – C ≡ C – H	Give answer for each central atom Linear Linear	Give answer for each central atom Linear Linear	Give answer for each central atom 180° 180°	Overall non-polar (symmetric)
SO ₂ sulfur dioxide 6 + 2(6) = 18 v.e.	0 = \$ ↔ ★ 0 :	Trigonal planar	Bent	120°	Polar (S—O) and bent shape



NOTE: A "C" labeled at an end of a bond means that a C atom is required at that position for that functional group (e.g., ether, ketone, ester). Bonds without atom labels at an end mean that either C or H atoms may be present at those positions (e.g., aldehyde, carboxylic acid, amine, amide).

Skeletal Structures



Organic Compounds are prevalent in nature and comprise most of the compounds found in biological systems (for example, proteins, lipids, carbohydrates, and vitamins). Organic compounds contain carbon and other elements such as hydrogen, oxygen and nitrogen (and many more). Due to carbon's ability to form up to four bonds, these molecules can get very large, from molecules having less than 10 carbon atoms, to many thousands of carbon atoms! Due to the many numbers of carbons and hydrogens in organic molecules, there is a shorthand notation for molecules which is called a **skeletal structure**.

It's much easier to draw and read organic molecular structures by using this simplified system of notation. At the end of every bond is a carbon unless noted otherwise (O, N, etc.). Every neutral carbon has four bonds, so often the hydrogens that are attached are implied (not drawn).

EXAMPLE

Let's look at the molecule methamphetamine, which is a highly addictive drug.



- 1. Use the expanded structure to count how many carbons **10** and hydrogens **15** are in a methamphetamine molecule.
- 2. Practice drawing the carbons into the skeletal structure. Note that carbons are assumed where two lines meet and at the end of a line where there is no label.

3. Practice drawing the hydrogens into the skeletal structure. Note that hydrogens are drawn in to give carbon four bonds.
4. Atoms other than carbon and hydrogen need to be shown explicitly, as well as their hydrogens. Confirm this is the case with the nitrogen in the molecule.

 Lone pairs are often implied in organic structures for elements other than C and H. The molecule methamphetamine has 1 implied lone pairs on the N atom in order to satisfy the octet rule.

PRACTICE

ŇH₂ Serotonin is a neurotransmitter, often called the "happiness molecule." CH2 H₂C Н 1. How many carbons? 10 С ΗÖ C 2. How many hydrogens? 12 CH 3. How many nitrogens? 2 HC 4. How many total lone pairs? 4 H C н

Drawing Organic Structures



Chemical Formula	Expanded Structure	Condensed Structure	Skeletal Structure (Shorthand Notation)
C ₄ H ₁₀		H ₃ C H ₂ C CH ₃	
C ₅ H ₁₂ 0	$\begin{array}{cccccccccccccccccccccccccccccccccccc$	H ₃ C CH ₂ CH ₂ CH ₃ CH ₂ CH	ОН
C2H402	$ \begin{array}{c} 0 \\ H \\ H$:0: H ₃ C— С — ёН	°O∶ ↓↓ ÖH
C7H9N	H H C C H H C C H		NH ₂
C7H₀O	$ \begin{array}{c} H & 0 \\ H & H \\ H & C \\ H & C \\ H \\ H$	HC HC HC HC HC HC HC HC HC HC HC HC HC H	O H Or O
C ₅ H ₁₀		$H_{2}C$	

CHEM& 121 WORKBOOK

UNIT 7 SOLIDS, LIQUIDS, GASES, AND INTERMOLECULAR FORCES



Notes: Intermolecular Forces



Questions to get you thinking: Why does ice float? How do DNA strands stick together? Methane and hexane are both hydrocarbons, but why is hexane a liquid and methane is a gas?

Questions to get us started:

- What are some words that begin with "inter"? internet, interstate (highway)
- What are some words that begin with "intra"? intranet, intravenous
- What do the word prefixes "intra" and "inter" mean? within between

The covalent bonds that hold atoms together in molecules are intramolecular forces, which greatly influence the **chemical** properties of molecules. **Covalent bonds**

On the other hand, **intermolecular forces "IMFs"** are attractive forces that occur **between** molecules and influence the **physical** properties of molecules.

REVIEW OF POLARITY OF BONDS

Using the periodic table only, from least polar to most polar: C—H, N—H, O—H, H—H, and F—H.

least polar H—H C—H N—H O—H F—H most polar

Smallest ∆EN

Greatest ΔEN (change in electronegativity)

THREE MAJOR TYPES OF INTERMOLECULAR FORCES

Let's focus on the intermolecular forces that exist between neutral molecules (sometimes also referred to as van der Waals forces). There are three major types of IMF:

dispersion forcesdipole-dipole forceshydrogen bonds

We'll be using the following relationship while we learn about the three types of IMFs:

As the strength of intermolecular forces increases, the boiling point (bp) also increases because more energy is needed to disrupt the forces of attraction between molecules.



DISPERSION FORCES



The larger the atom or molecule the more electrons there are, and the electron cloud can be distorted more easily. This is called polarizability and it results in stronger dispersion forces.

F—F vs. Br—Br Larger and more massive, so more dispersion forces. \uparrow IMFs

EFFECT OF MOLECULAR SIZE ON STRENGTH OF DISPERSION FORCES

CH ₃ CH ₃	CH ₃ CH ₂ CH ₂ CH ₃
Propane	Butane ← More mass
44 amu bp = – 42°C	58 amu bp = 0°C

On your own: Rank CCl₄, CBr₄, and CH₄ in order of increasing polarizability.

CH_4 , CCI_4 , CBr_4 Highest mass, \uparrow IMFs. Polarizability is the degree to which there can be induced dipoles. EFFECT OF MOLECULAR SHAPE ON STRENGTH OF DISPERSION FORCES

The shape of molecules can also impact how easily the electron cloud can be distorted (polarizability), and therefore also the strength of dispersion forces.









H₃C

CH₃

Neopentane

 C_5H_{12}

bp = 9.6°C

Fewer IMFs than long chain molecules

DIPOLE-DIPOLE FORCES

Dipole-dipole forces exist between **polar** molecules.

Both attractive and repulsion forces occur, but there is a net attraction overall.

 $\begin{array}{c} \begin{array}{c} & & & & & & \\ & & & \\ & & & \\$

Which one of these two molecules is nonpolar? Propane

Which one is polar? Label the partial charges on the polar molecule using δ + or δ - or dipole arrows. Dimethyl ether

How would polar molecules align when they are near each other? **Such that opposite charges are together.**

On your own: Would CS_2 or CH_3F have dipole-di- $\ddot{S} = C = \ddot{S}$ pole forces? Non-polar



HYDROGEN BONDS

Hydrogen bonds are very strong dipole-dipole forces. "SUPER" dipole-dipole forces!

The term "hydrogen bonds" can be a bit confusing as they are not covalent bonds, but are instead attractive forces between molecules.

Special requirements for hydrogen bonding to occur: **Must be between molecules with** *very* **polar bonds**.

Definition of a hydrogen bond:

A special type of dipole–dipole interaction between a very electropositive H (covalently bonded to N, O, or F) and an N, O, or F on another molecule.

A dashed line (-----) is typically used to draw hydrogen bonds.



Draw the possible H-bonds between H₂O and NH₃ molecules.

Let's look at how hydrogen bonding influences boiling points. Determine whether each of these molecules is capable of hydrogen bonding.



On your own: Would hydrogen bonding be present in samples of CH_3F , HF, or CH_3OCH_3 ? Draw hydrogen bond interactions for the substances capable of hydrogen bonding (you'll need to draw more molecules).



SUMMARY OF RELATIVE STRENGTHS OF THE THREE TYPES OF IMFS

weakest	disnersion	dipole-dipole	H bonding	strongest
WEakest	uispersion	uipoie-uipoie	n bonung	Sciuliyesi

And, a review question for you: What are hydrogen bonds? Are they covalent bonds to H atoms? **No** Or, forces between molecules with very polar bonds? **Yes**

RELATIVE STRENGTHS OF IMFS COMPARED TO CHEMICAL BONDS

Here is some data to look at:

- Energy required to disrupt intermolecular forces in 18 grams (1 mole) of water to cause physical state changes:
 - solid to liquid: 6.02 kJ
 liquid to gas: 40.7 kJ
- Energy required to break both 0—H covalent bonds in the same amount of H₂O: 928 kJ

So, which are stronger, IMFs or covalent bonds?



Review: Intermolecular Forces



1. Determine the Lewis structure, electron geometry, shape, and polarity of each of the following. Then identify the strongest type of interactive force (dispersion, dipole-dipole, hydrogen bonding) that occurs between molecules in each of the following substances.



2. Predict (circle) which of each of the following in each pair has the higher boiling point. Explain your choice briefly.

a. NaCl or HCl
 b. Br₂ or HBr
 Br₂ - dispersion
 HBr - H bonding
 C. H₂O or H₂S
 H₂O - H bonding
 H₂S - dipole-dipole
 d. C₂H₆ or C₈H₁₈
 Both have dispersion forces;
 Both have dispersion forces;
 C₈H₁₈ - greater mass
 3. Identify whether each of the following statements about water is true or false.

a.	Water can dissolve many polar and ionic compounds.	True	or False
b.	Polar substances dissolve with nonpolar substances.	True	or False
C.	Hydrophobic substances are likely non-polar.	True	or False
d.	Intermolecular forces are stronger than covalent bonds.	True	or False
e.	When water boils, its covalent bonds break.	True	or False
_			

4. Draw a hydrogen bond (as a dotted line) that can occur between two water molecules.



5. Describe at least two properties of water which make it such a unique compound. Explain these properties using the concept of intermolecular forces.

High boiling point (100°C) – requires a lot of energy for liquid to gas phase change due to strong IMFs (H bonds).

High surface tension – strong tendency to attract itself/minimize surface area due to strong IMFs (H bonds).

Identifying Organic Functional Groups in Molecules

Functional Groups: Groups of atoms within molecules that have particular chemical/physical properties—chemical reactivity, boiling point, etc.

Identify the type of compounds that contain the following functional groups:



The following are interesting molecules you may have heard of before. Answer the questions about each one and circle/identify the functional groups present.



Activity: Properties of Organic Molecules



(Adapted from *General, Organic, and Biological Chemistry: A Guided Inquiry* by Michael Garoutte.)

 Two atoms share a covalent bond. If the electronegativity difference is zero, the bond is non-polar. As the electronegativity difference increases, the bond becomes increasingly polar. Using a periodic table instead of electronegativity values, rearrange the following bonds from least polar (left) to most polar (right): C—H, N—H, O—H, H—H, F—H

Least polar H—H C—H N—H O—H F—H Most polar

2. For our purposes, the two least polar bonds above will be called "non-polar." The rest will be "polar." Make a note of this above.

MODEL 1: SOME PHYSICAL DATA FOR SELECTED ORGANIC MOLECULES

	Skeletal Structure	Molecular Weight (amu)	Dipole (Debyes)	Boiling Point (°C)	Solubility (g/100 mL H ₂ O)			
Alkanes								
Propane	CH ₂ H ₃ C CH ₃	44	0	-42	0.007			
Butane	CH ₂ CH ₃	58	0	0	0.006			
Alcohols	Alcohols							
Ethanol	ОН	46	1.7	78	High (∞)			
Propanol	ОН	60	1.7	82	High (∞)			
Butanol	ОН	74	1.7	118	6.3			
Ethers	Ethers							
Dimethyl ether	0	46	1.3	-23	High (∞)			
Diethyl ether		74	1.2	35	6.9			

- 3. According to the dipole moments in the data table (Model 1), which functional group is most POLAR? Circle one: alkane alcohol ether
- 4. According to the data table, the general trend shows that the more polar a compound is, the **higher** its boiling point.
- 5. For each of the class of compounds below, indicate how the boiling point changes as the molecular weight (MW) increases.
 - a. alkanes b. alcohols c. ethers As MW increases, boiling point increases for each of these.
- 6. Intermolecular forces are the attractions molecules have for one another. The stronger the intermolecular forces, the higher / lower its boiling point.
- 7. Find one molecule from each functional group (alkane, alcohol, ether) with roughly the same molecular weight (within 5 amu). Write their names and structures below, and rank them in order of low to high boiling point.

Example: propane dimethyl ether ethanol $bp = -42 \ ^{\circ}C$ $0 \ bp = -23 \ ^{\circ}C$ $OH \ bp = 78 \ ^{\circ}C$

8. Based on your answer above, which functional group appears to have ...

the strongest intermolecular forces?			Circle one:	alkanes	alcohols	ethers
the weakest?	Circle one:	alkanes	alcohols	ethers		

Circle the type of bonds found in each type of functional group below. What type of bond seems to account for the most polar functional group? 0—H

Functional Group	Type of Bond					
Alkane	С—Н	C—C	C—0	0—H		
Alcohol	С—Н	C—C	C—0	0—Н		
Ether	С—Н	C—C	C—0	0—H		

- 10. Using the data in the table (Model 1), it appears that polar functional groups are **more** water soluble?
- 11. How does water solubility change as the number of carbons in the molecules increases? **It decreases.**
- 12. Draw the dipoles for a water molecule using δ + and δ symbols:

 δ_{+}^{0}

Raise your hand and ask your instructor for the "magnetic water molecules." Use them to answer the rest of the questions:

- 13. To what regions of one molecule is another molecule attracted?
 - Sketch a drawing and use words and dotted lines to describe the attraction between the two water molecules.
 - Label "covalent bonds" and label "attractive forces" separately in your diagram.



- 14. When you break the two water magnets apart, are any covalent bonds breaking? During evaporation or vaporization (boiling), are covalent bonds broken or intermolecular forces? No—you are only breaking the attractive forces.
- 15. The attraction between two water molecules is called "hydrogen bonding." Hydrogen bonds are not covalent bonds—they are much weaker compared to covalent bonds and more easily broken.

Hydrogen bonds are usually shown by dotted lines. Water is not the only compound to hydrogen bond—others may also do so.

Only one of these correctly represents a hydrogen bonding interaction. Circle it. (If you don't know, ask your instructor for a methanol CH₃OH magnet.)



HINT: Is a C—H bond polar? No!

- 16. Describe what is wrong with each of the other three pictures above. See above.
- 17. Which of these samples would exhibit hydrogen bonding? Circle the ones that can hydrogen bond to other molecules like itself.



Intermolecular Forces and Physical States



c. HCN

- 1. Which of the three types of intermolecular forces is the weakest? **Dispersion**
- 2. What type(s) of intermolecular forces are present in the following?
 - a. N_2 b. CH_3NH_2 H bonding $:N \equiv N:$ (non-polar) (verv polar Dispersion

 $H - \overset{\overset{\leftarrow}{}}{C} \equiv \overset{\leftarrow}{N}$: Dipole-dipole N—H bonds) For the substance(s) above that are capable of hydrogen bonding, draw molecular level drawings that show the hydrogen bonding (used dashed lines to represent hydrogen bonding). On your drawing point out the difference between an intramolecular force

and an intermolecular force. CH3NH2



- 3. Acetone (a ketone with a molecular formula of C_3H_6O) evaporates more quickly than water at room temperature.
 - a. Which compound is more volatile? **Acetone. Volatile is the tendency to vaporize.**
 - b. Explain this phenomenon in terms of the intermolecular forces present in a sample of each compound. (What IMFs are present in each? Which IMFs are stronger? How does this info relate to volatilities?)

 $H_3C \sim C \sim CH_3$ Dipole-dipole Weaker so greater H bonding Stronger tendency to vaporize (less volative) H₂C ~ C ~ CH₂ (liquid \rightarrow gas)

4. Which of the following is a liquid (rather than a gas) at room temperature? Explain. CH₃SH or CH₃OH

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CH<sub>3</sub>OH – H bonding
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Stronger IMF = higher boiling point.

- 5. In which physical state do the particles have a definite volume but an indefinite shape? Liquids – takes the shape but not always the volume of its container.
- 6. In which physical state does the volume of the sample greatly exceed the actual volume of the particles?

Gases – they fill the volume of the container. 0 0 0 0 0 0 0 7. How does the pressure of a gas relate to its volume? (Assume constant amount of gas and constant T.) Draw a series of pictures to illustrate your answer (e.g., gas filled syringes, balloons, etc.).



- 8. Identify the changes of state occurring in each of the following. Also specify the process as endothermic or exothermic.
 - a. Ice cubes get smaller in a freezer.



b. Gaseous tin atoms form as a solid on a glass surface.



Is heat energy added or removed from a system undergoing condensation? Explain using the terms "kinetic energy" and "intermolecular forces."
 Gas particles move faster (higher kinetic energy) and have weaker IMFs than liquids.

Energy must be removed to decrease kinetic energy and allow for stronger IMFs to form for the phase change.

 $\begin{array}{rcl} \textbf{Gas} & \rightarrow & \textbf{liquid} \\ \textbf{(more energy) (less energy)} \end{array}$

E is released/removed.

Exothermic.

CHEM& 121 WORKBOOK

UNIT 8 IONIC COMPOUNDS



Notes: Ionic Compounds



Type I: "A" metals

Type II: "B" transition metals

There are two major ways that atoms achieve electron configurations similar to the noble gases: either *share* electrons or *transfer* electrons. Covalent molecules and compounds are formed from the sharing of electrons. Let's focus now on the *transfer* of electrons in the formation of **ionic compounds**.

The main group elements chemically react in ways that allow atoms to attain the same number of valence electrons as the noble gases. Atoms react to achieve an **octet** of electrons.

Get to the				C	N	0	F	Ne
Gas Fast!	Na	Mg	AI					
# Valence e ⁻ ?	1	2	3	4	5	6	7	8
Gain or Lose e ⁻ ?	Lose	Lose	Lose	N/A	Gain	Gain	Gain	N/A
How Many?	1 e ⁻	2 e ⁻	3 e ⁻		3 e ⁻	2 e ⁻	1 e ⁻	
Net Charge = (# p ⁺) – (# e ⁻)	+	2+	3+		3–	2–	_	
Element Symbol with Net Charge	Na ⁺	Mg ²⁺	Al ³⁺		N ^{3–}	0 ^{2–}	F	

TRANSFERRING ELECTRONS

When electrons are transferred *from one* atom *to another* atom, the atoms can no longer be neutral atoms; they now have charges. The charged atoms are called **ions**.

Anions

• The attractive force between positively and negatively charged particles is called an **ionic bond**, and the resulting substance is called an **ionic compound**.

_

- One atom's loss is another's gain: the electrons lost by a metal atom are simultaneously gained by a nonmetal atom, but it's not always a 1:1 ratio ... it depends on the atoms involved and how many e⁻ they need to gain/lose.
- There is always a balance of + and charges because ionic compounds have net zero charge.
- The resulting 3-D structure of stacked ions is called a **crystal lattice**.
- The smallest repeating unit of a crystal lattice is called a formula unit instead of a molecule.

⁺ Cations

Na & Cl

Simple drawing:

[Na]⁺ [:Ċ̈́l:][−]

Compound formula: **NaCl** Compound name: **Sodium chloride**

Equations using Lewis structures:

 $\begin{array}{l} Na\cdot \rightarrow Na^{+}+1 \; e^{-} \\ :\ddot{C}l:+1 \; e^{-} \rightarrow :\ddot{C}l: \end{array}$

Mg & F

Simple drawing: [**Mg]²⁺ [:Ĕ:]⁻** [:Ĕ:]⁻ Compound formula: MgF₂ Compound name: Magnesium fluoride

Equations using Lewis structures:

 $\begin{array}{l} {}^{\cdot}Mg {}^{\cdot} \longrightarrow \left[Mg\right]^{2+} + 2 \; e^{-} \\ {}^{\cdot}\ddot{F} {}^{\cdot} + e^{-} \longrightarrow \left[{}^{\cdot}\ddot{F} \; {}^{\cdot}\right]^{-} \\ {}^{\cdot}\ddot{F} {}^{\cdot} + e^{-} \longrightarrow \left[{}^{\cdot}\ddot{F} \; {}^{\cdot}\right]^{-} \end{array}$



Charges of lons



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): N/A +1 +2 N/A -2 -3 _1 Assume all others are variable charged. Fixed +3 \downarrow +3 +2 \downarrow +1 +2 \downarrow
- 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	0 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; C⁴⁻ is carbide and Si⁴⁻ is silicide.

NAMES OF METAL CATIONS (POSITIVELY-CHARGED IONS)

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

EXAMPLES

Atom/Element	lon	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 CI chlorine CI_2 chloride ion CI^-

lonic 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	lon Formula (Show Charge)	Cation?	Anion?	Name of Ion
Na	Na ⁺	X		Sodium ion
0	0 ^{2–}		X	Oxide
Zn	Zn ²⁺	X		Zinc ion
F	F		X	Fluoride
К	К+	X		Potassium ion
N	N ^{3–}		X	Nitride
AI	Al ³⁺	X		Aluminum ion
Р	P ³⁻		X	Phosphide
Sr	Sr ²⁺	X		Strontium ion
S	S ^{2–}		X	Sulfide

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.

а.	Barium and Oxygen Ba ²⁺ [:Ö:] ²⁻ → BaO	Element	Lewis Dot Structure	Lewis Dot Structure after Reaction (HINT: Cation, Anion, or Neither?)
		0	. <u>ö</u> .	[:Ö:] ^{2–}
b. Al [A	Aluminum and Flourine Al] ³⁺ 3[:بَّ::] ⁻	Ва	·Ba·	[Ba] ²⁺
		CI	÷Ċŀ	[:Ċ]:] [−]
	Determine and to the	AI	٠À	[AI] ³⁺
с. Р [I	Potassium and lodine [K] ⁺ [:į:] ⁻	Ar	:Är:	:Är:
		Са	·Ca·	[Ca] ²⁺
d.	Calcium and Nitrogen 3[Ca]²⁺ 2[:៉៉:]³⁻	N	٠Ņ٠	[:៉៉:] ^{3–}
		К	K٠	[K] ⁺
		I	÷Ï÷	(jji)⊤

Notes: Naming Binary Ionic Compounds



Binary compound means there are two different elements present in the compound.

First, naming ions on their own ...

Constant charge cations:

Variable charge cations:

Anions:

metal i	ons: element name and "ion"	Na ⁺ Mg ²⁺		
nonme	t al ions: (stem of element nam	e + "—ide") a	nd "ion"	
F	fluoride ion	0	Oxide	
CI	Chloride ion	S	Sulfide	
Br	Bromide ion	Ν	Nitride	
I	lodide ion	Р	Phosphide	
compound name = metal and nonmetal stem + "–ide"		What are these co	<i>What are the ions that make up these compounds?</i>	
NaCl	sodium chloride		$Na^+ \& CI^-$	
KCI	Potassium chloride		\mathbf{K}^+ and \mathbf{CI}^-	
MgF_2	Magnesium fluoride		${ m Mg}^{2+}$ and ${ m F}^-$	
Cs ₂ 0	Caesium oxide		Cs^+ and O^{2-}	

Some transition metals form more than one type of ion; for example, Fe²⁺ and Fe³⁺. We differentiate these two ionic forms of iron by using Roman numerals to specify the charge:

 Fe^{2+} is iron (II) and Fe^{3+} is iron (III)

in compounds	compound formula	compound name
$Pb^{2+} \& Cl^{-}$	PbCl ₂	Lead (II) chloride
Pb^{4+} & Cl^{-}	PbCl ₄	Lead (IV) chloride

Find v	where	on	Periodic	Table?
i inu v	1010	011	i chiouio	Tuble:

+1+2 +2 +1 Use Roman Numerals?

No (constant charge)

Yes (variable charge)

MONATOMIC VS. POLYATOMIC IONS

The ions described so far have been individual atoms that attain a charge; these ions are monatomic ions. **Polyatomic ions** are formed when *a group of atoms collectively gain or lose electrons.*

lons	Compound Formula
ammonium	NH4 ⁺
hydronium	H ₃ 0 ⁺
hydroxide	OH⁻
carbonate	C03 ²⁻
hydrogen carbonate (bicarbonate)	HCO ₃ -
nitrate	N03 ⁻
nitrite	N02 ⁻
phosphate	P04 ³⁻
hydrogen phosphate	HP04 ²⁻
dihydrogen phosphate	H ₂ PO ₄ ⁻
chromate	Cr04 ²⁻
dichromate	Cr ₂ 0 ₇ ²⁻
acetate	$C_2H_3O_2^-$ or CH_3COO^-
sulfate	\$0 ₄ ²⁻
sulfite	\$0 ₃ ²⁻
hydrogen sulfate (bisulfate)	HSO ₄ -
hydrogen sulfite (bisulfite)	HSO ₃ ⁻
perchlorate	CI04 ⁻
chlorate	CI03_
chlorite	CI02 ⁻
hypochlorite	CIO ⁻
cyanide	CN ⁻
permanganate	Mn0 ₄ ⁻
thiosulfate	S ₂ O ₃ ²⁻
oxalate	C ₂ O ₄ ²⁻
borate	B0 ₃ ³⁻
citrate	$C_{6}H_{5}O_{7}^{3-}$ or $C_{3}H_{5}O(COO^{-})_{3}$

Let's draw the Lewis dot structure of the carbonate ion, CO_3^{2-} . The formula indicates that there is 1 carbon atom, 3 oxygen atoms, and 2 electrons in addition to the valence electrons contributed by the C & O atoms.

Can more than one structure be drawn for $CO_3^{2-?}$ And, draw the Lewis structure for NH_4^+ .



IONIC COMPOUNDS INVOLVING POLYATOMIC IONS

If one carbonate ion has a 2– charge, what would be the formula for the compound that forms between

calcium ion and carbonate ion? **CaCO**₃

sodium ion and carbonate ion? Na₂CO₃

ammonium ion and carbonate ion? (NH₄)₂CO₃

COMMON POLYATOMIC IONS

The list of polyatomic ions will be provided for you to use on tests. You do not need to memorize them, but you do need to be familiar with the names and formulas so that you can efficiently find them in the list.

Notice that most of the names end in the suffixes -ate or -ite.

The exceptions to this are hydroxide and cyanide. All other ions with the –ide ending are monatomic ions, and their charge is determined by their position on the periodic table.

lons	Compound Formula	Compound Name
Pb ²⁺ & SO ₄ ²⁻	PbSO ₄	lead (II) sulfate
NH ₄ ⁺ & NO ₃ ⁻	NH ₄ NO ₃	ammonium nitrate
Ca ²⁺ & PO ₄ ³⁻	Ca ₃ (PO ₄) ₂	calcium phosphate
Fe ³⁺ & OH ⁻	Fe(OH) ₃	iron (III) hydroxide

Bonding Recap: Attention all carbon-based lifeforms! Carbon doesn't tend to form cations or anions ... how does it make all kinds of organic compounds? And, what's up with diatomic molecules, like O_2 and N_2 ? Transferring e⁻ won't work. When bonding together, nonmetal atoms (upper right-hand corner of the periodic table) like to pool their resources. They share! Aww, isn't that cute? When metal atoms and nonmetal atoms react (from opposite sides of the periodic table), they don't share with each other—there's a winner and a loser (transfer e⁻)—resulting in ionic compounds.

Tutorial: Ionic Nomenclature



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

TYPES OF CATIONS (POSITIVE IONS)

- 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	$AI^{3+} O^{2-}$	AI_2O_3
---	----------------	------------------	-----------

- Iron (II) phosphide $Fe^{2+}P^{3-}$ Fe_3P_2
- Calcium nitrate
 Ca²⁺ NO₃¹⁻ Ca(NO₃)₂
- Copper (I) carbonate $Cu^{1+} CO_3^{2-} Cu_2CO_3$
- Barium oxide Ba²⁺ O²⁻ BaO
 - **NOTE:** The charges are already balanced so the ratio is 1:1.
- Calcium hydroxide
 Ca²⁺ OH¹⁻
 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

- Na₂O Sodium oxide (Group IA—no Roman numeral needed)
- Cr₂S₃ Chromium (III) sulfide (Most transition metals need Roman numerals.)
- BaSO₄ Barium sulfate (Group IIA—no Roman numeral needed)
- Pb(OH)₂ Lead (II) hydroxide (Pb is a main group metal with variable charge.)
- Cu₂CO₃ Copper (I) carbonate (Most transition metals need Roman numerals.)
- FeSO₄ Iron (II) sulfate (Most transition metals need Roman numerals.)

Ionic Nomenclature 1

- 1. These are binary ionic compounds (two ions involved). The second ion carries a binary ending of –ide.
 - a. BeS Beryllium sulfide
 - b. CaH₂ Calcium hydride
 - c. Nal Sodium iodide
 - d. Mg0 Magnesium oxide
 - e. AlBr₃ Aluminum bromide
 - f. Li₂S Lithium sulfide
 - g. ZnF₂ Zinc fluoride
 - h. K₃N Potassium nitride

- i. AgCl Silver chloride
- j. AIN Aluminum nitride
- k. BeCl₂ Beryllium chloride
- I. Cal₂ Calcium iodide
- m. KF Potassium fluoride
- n. MgBr₂ Magnesium bromide
- o. Ag₂S Silver sulfide
- p. ZnO Zinc oxide
- 2. These are ternary ionic compounds (they involve a polyatomic ion). First, recognize which ion is used. Then use your ion chart to name the compound.
 - a. NH₄Br **Ammonium bromide**
 - b. AI(OH)₃ Aluminum hydroxide
 - c. Li₃PO₄ Lithium phosphate
 - d. Mg(C₂H₃O₂)₂ Magnesium acetate
 - e. Zn(HSO₄)₂ Zinc hydrogen sulfate or zinc bisulfate
 - f. (NH₄)₂CO₃ Ammonium carbonate

- g. NaClO₃ Sodium chlorate
- h. Ca(CN)₂ Calcium cyanide
- i. KHCO₃ Potassium hydrogen carbonate or potassium bicarbonate
- j. Mg₃(PO₄)₂ Magnesium phosphate
- k. Al(NO₂)₃ Aluminum nitrite
- I. Ag₂SO₄ Silver sulfate
- 3. These are binary or ternary and involve a variable charged metal (you must indicate the charge as a Roman numeral in the name).
 - a. Cr₂O₃ Chromium (III) oxide
 - b. CoN Cobalt (III) nitride
 - c. Fel₃ Iron (III) iodide
 - d. Co(HSO₄)₂ Cobalt (II) bisulfate
 - e. Cu₂O Copper (I) oxide

- f. Hg(ClO₃)₂ Mercury (II) chlorate
- g. NiBr₂ Nickel (II) bromide
- h. CuSO₃ Copper (II) sulfite
- i. Ni(NO₃)₃ Nickel (III) nitrate
- j. Fe(CH₃COO)₂ Iron (II) acetate



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 $ ightarrow$ cation \downarrow	chloride Cl ⁻	oxide S ²⁻	nitrate S04	sulfite S0 ^{2–}	phosphate CO ₃ ^{2–}
sodium Na⁺	NaCl	Na ₂ S	Na ₂ SO ₄	Na ₂ SO ₃	Na ₂ CO ₃
calcium Ca ²⁺					
aluminum Al ³⁺					
iron (II) Fe ²⁺					
iron (III) Fe ³⁺					
copper (I) Cu ⁺	CuCl	Cu ₂ S	Cu ₂ SO ₄	Cu ₂ SO ₃	Cu ₂ CO ₃
copper (II) Cu ²⁺	CuCl ₂	CuS	CuSO4	CuSO ₃	CuCO ₃
chromium (II) Cr ²⁺					
silver Ag⁺					
zinc Zn ²⁺					
ammonium NH ⁺ 4	NH₄CI	(NH ₄) ₂ S	(NH ₄) ₂ SO ₄	(NH ₄) ₂ SO ₃	(NH ₄) ₂ CO ₃

Ionic Nomenclature 2



1. Write the formula for each of the following compounds:

a.	iron (II) chloride	g.	sodium iodide
b.	barium nitride	h.	magnesium oxalate
C.	zinc chromate	i.	calcium chlorite
d.	chromium (III) oxide	j.	aluminum fluoride
e.	tin (IV) oxide	k.	copper (II) nitrate
f.	lithium chloride	I.	aluminum sulfide

2. Write the names for each of the following compounds:

a.	K ₂ SO ₃	g.	AlBr ₃
b.	Mg ₃ N ₂	h.	Al ₂ (SO ₄) ₃
C.	HgCl ₂	i.	AgN0 ₃
d.	Fe(NO ₂) ₂	j.	Ca ₃ (PO ₄) ₂
e.	PbBr ₂	k.	Na ₂ CO ₃
f.	Ca(C ₂ H ₃ O ₂) ₂	I.	Mn(0H)2

ANSWERS TO IONIC NOMENCLATURE 2

You can use these answers as another worksheet.

1. Write the formula for each of the following compounds:

a.	FeCl ₂	g.	Nal
b.	Ba ₃ N ₂	h.	MgC_2O_4
C.	ZnCrO ₄	i.	$Ca(CIO_2)_2$
d.	Cr ₂ O ₃	j.	AIF_3
e.	SnO ₂	k.	$Cu(NO_3)_2$
f.	LiCl	I.	Al_2S_3

2. Write the names for each of the following compounds:

a.	potassium sulfite	g.	aluminum bromide
b.	magnesium nitride	h.	aluminum sulfate
C.	mercury (II) chloride	i.	silver nitrate
d.	iron (II) nitrite	j.	calcium phosphate
e.	lead (II) bromide	k.	sodium carbonate

f. calcium acetate I. manganese (II) hydroxide

Ionic and Covalent Nomenclature



Identify the following compounds as ionic or covalent (molecular).

1. Write the names of the following:

a. Na₂S _____ j. Fe₂O₃_____ k. CuNO₃_____ b. P₂O₅ _____ Sn(SO₃)₂ c. FeO _____ Ι. d. NH₄CI m. AIBr₃ n. Al₂(SO₄)₃ _____ e. Ba(NO₃)₂ _____ o. Cu₂S_____ f. PbCrO₄_____ g. AgCl _____ p. CCl₄ _____ q. $Zn(C_2H_3O_2)_2$ _____ h. K₃N _____ i. N₂O₃ _____ r. $Fe_2(Cr_2O_7)_3$ _____ 2. Write the formula for the following compounds: a. barium bromide _____ cobalt (II) sulfide _____ j. b. nickel (II) chlorate _____ k. chromium (III) sulfate _____ aluminum sulfite _____ c. copper (II) chloride _____ Ι. d. sodium iodide _____ m. sodium phosphate e. carbon monoxide_____ n. calcium sulfide _____ f. magnesium nitride cobalt (II) nitrate _____ 0. g. iron (II) oxide calcium acetate p. phosphorous trichloride _____ h. dinitrogen tetrachloride_____ q. i. barium sulfate phosphorous pentachloride r.

(Answers are given on the next page.)

ANSWER SHEET

You can use the answer sheet as a worksheet. That is, work backwards and see if you can write the formulas and the names.

1.	Write the names of the following:				
	a.	sodium sulfide ionic	j.	iron (III) oxide ionic	
	b.	diphosphorous pentoxide covalent	k.	copper (I) nitrate ionic	
	C.	iron (II) oxide ionic	I.	tin (IV) sulfite ionic	
	d.	ammonium chloride ionic	m.	aluminum bromide ionic	
	e.	barium nitrate ionic	n.	aluminum sulfate ionic	
	f.	lead (II) chromate ionic	0.	copper (I) sulfide ionic	
	g.	silver chloride ionic	p.	carbon tetrachloride covalent	
	h.	potassium nitride ionic	q.	zinc acetate ionic	
	i.	dinitrogen trioxide covalent	r.	iron (III) dichromate ionic	
n	۱۸/				

2. Write the formula for the following compounds:

a.	BaBr ₂ ionic	j.	CoS ionic
b.	Ni(ClO ₃) ₂ ionic	k.	Cr ₂ (SO ₄) ₃ ionic
C.	CuCl ₂ ionic	I.	Al ₂ (SO ₃) ₃ ionic
d.	Nal ionic	m.	Na ₃ PO ₄ ionic
e.	CO covalent	n.	CaS ionic
f.	Mg ₃ N ₂ ionic	0.	Co(NO ₃) ₂ ionic
g.	FeO ionic	p.	$Ca(C_2H_3O_2)_2$ ionic
h.	N ₂ Cl ₄ covalent	q.	PCl ₃ covalent
i.	BaSO ₄ ionic	r.	PCI ₅ covalent

Review: Ionic and Covalent Compounds



1. Write the Lewis structure of Mg.

Mg∙

2. Write the Lewis structure of the compound NaBr and name the compound.

 $Na^{+}Br^{+} \rightarrow Na^{+}[Br^{+}Br^{+}]^{-}$

3. Use Lewis theory to predict the formula of the compound that forms between magnesium and nitrogen and name the compound.

Mg₃N₂

4. Write the Lewis structure for CO and name this compound.

 $4 + 6 = 10 \text{ ve}^-$: \overrightarrow{C} : \overrightarrow{O} : \rightarrow : $\overrightarrow{C} \equiv 0$:

5. Write the Lewis structure for H_2CO . + 4 + 6 = 12 ve⁻ ; ; ; ; ; ; ; ;

6. Write the Lewis structure for the $C - H \rightarrow H - C - H$

 $7 + 6 + 1 = 14 \text{ ve}^-$

- 7. Write the Lewis structure for the NO_2^- ion. Include resonance structures. $5 + 2(6) + 1 = 18 \text{ ve}^- : \ddot{O} - \ddot{N} - \ddot{O} : \Rightarrow : \ddot{O} - \ddot{N} = \ddot{O} : \iff \ddot{O} = \ddot{N} - \ddot{O} :$
- 8. Predict the molecular geometry of CINO (N is the central atom).

7 + 5 + 6 = 18 ve⁻ :
$$\ddot{C}I - \ddot{N} - \ddot{O}: \rightarrow :\ddot{C}I - \ddot{N} = O: \rightarrow CI - \tilde{N} \ge O$$

9. Predict the molecular geometry of the SO₃²⁻ ion.
6 + 18 + 2 = 26 ve⁻ : $\ddot{O}:$ Trigonal pyramid
: $\ddot{O} < \overset{\circ}{S} > \ddot{O}:$

- 10. Polar, covalent, or ionic? Using the Pauling electronegativity values, predict whether the bond formed between each pair of atoms is pure covalent, polar covalent, or ionic.
 - a. I and I Pure Covalent b. Cs and Br Ionic c. P and O Polar Covalent
- 11. Determine whether the following molecules are polar or nonpolar covalent.
 - a. CH₄ Nonpolar c. H₂S Polar
 - b. CO₂ Nonpolar d. NH₃ Polar

CHEM& 121 WORKBOOK

UNIT 9 THE MOLE AND MOLAR MASS



Notes: The Mole



We have to use mass and the chemist's counting unit—the mole—as a way to count chemical particles because atoms, molecules, and ions are too small to see to count by eye.

1 mole = 6.022×10^{23} particles

MAIN IDEAS OF THE MOLE

• The mole can be used for any object—atoms, molecules, formula units, ...

"Formula units" used instead of "molecules" for ionic compounds.

1 mol Al = 6.022×10^{23} Al atoms1 mol1 mol H20 = 6.022×10^{23} H20 molecules1 mol1 mol NaCl = 6.022×10^{23} NaCl formula units

1 mol donuts = 6.022×10^{23} donuts 1 mol \$ bills = 6.022×10^{23} \$ bills

• The mole is related to mass.

EXAMPLES

- 1 mole = # of atoms in exactly 12 g of carbon-12 atoms = 6.022 × 10²³ particles
- All atomic masses on the Periodic Table are relative to carbon-12 (1 amu = 1/12th of a C-12 atom), so the mass of 1 mole of any element is equal to the atomic mass in units of grams. This mass is referred to as the molar mass (MM) or molar weight (MW).

EXAMPLES OF MOLAR MASS

Element	Compound
6	(Molecule or Formula Unit)
С	H ₂ 0: Calculate Molar Mass
12.01	2(1.008) + 15.99 =

1 Mole of	# of Particles	Mass
Ar (M)		
Ca (I)		
NaCl (I.C.)		
SO ₂ (M.C.)		

CALCULATIONS INVOLVING THE MOLE & MASS

Use molar mass Use Avogadro's # Mass \leftarrow Mole \leftarrow # Particles # g/mol from periodic table One mole = 6.022 × 10²³ EXAMPLES

1. How much does a 13.6 mol sample of copper weigh in grams?

13.6 mol Cu × $\frac{63.55 \text{ g}}{1 \text{ mol Cu}}$ = 864.28 g = 864 g Cu

2. How many moles of CO_2 molecules are in 0.25 grams of CO_2 ?

$$0.25 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol } \text{CO}_2}{44.01 \text{ g } \text{CO}_2} = 0.00568 \text{ mol} = 0.0057 \text{ mol } \text{CO}_2$$

3. How many formula units of NaCl are in 0.7 moles of NaCl? How much would this sample of NaCl weigh?

0.7 mol NaCl ×
$$\frac{6.022 \times 10^{23} \text{ formula units}}{1 \text{ mol NaCl}} = 4.22 \times 10^{23} = 4 \times 10^{23} \text{ NaCl formula units}$$

0.7 mol NaCl ×
$$\frac{58.44 \text{ g NaCl}}{1 \text{ mol NaCl}} = 40.91 \text{ g} = 40 \text{ g NaCl}$$

4. 63.9 g of sucrose ($C_{12}H_{22}O_{11}$) contains how many moles of $C_{12}H_{22}O_{11}$ molecules? And, how many $C_{12}H_{22}O_{11}$ molecules is this?

$$63.9 \text{ g } C_{12}H_{22}O_{11} \times \frac{1 \text{ mol } C_{12}H_{22}O_{11}}{342.34 \text{ g } C_{12}H_{22}O_{11}} = 0.187 \text{ mol } C_{12}H_{22}O_{11}$$

$$\times \frac{6.022 \times 10^{23} \text{ C}_{12}H_{22}O_{11} \text{ molecules}}{1 \text{ mol } C_{12}H_{22}O_{11}} = 1.12 \times 10^{23} \text{ C}_{12}H_{22}O_{11} \text{ molecules}$$

5. What is the mass of 7.53×10^{25} molecules of water?

$$7.53 \times 10^{25} \text{ H}_2 \text{ 0} \text{ molecules} \times \frac{1 \text{ mol H}_2 \text{ 0}}{6.022 \times 10^{23} \text{ H}_2 \text{ 0} \text{ molecules}} \times \frac{18.02 \text{ g}}{1 \text{ mol H}_2 \text{ 0}} = 2,253 \text{ g} = 2,253 \text{ g}$$

CONNECTING CHEMICAL FORMULAS TO THE MOLE CONCEPT

H ₂ 0 Molecules	H Atoms	0 Atoms
٨	2 H atoms	1 O atoms
۸ م	4 H atoms	2 0 atoms
	24 H atoms	12 0 atoms
1 dozen	2 dozen H atoms	1 dozen 0 atoms
6.022×10^{23}	1.204 × 10 ²⁴ H atoms	6.022 × 10 ²³ 0 atoms
1 mole	2 mole H atoms	1 mole 0 atoms

Microscale: $1 H_2 0$ molecule contains **2** H atoms and **1** 0 atoms.

Macroscale: $1 \mod H_2 0$ molecules contains 2 moles H and 1 moles 0.

EXAMPLES

1. How many moles of H atoms are in 2.9 mol of H_2O ?

2.9 mol H₂0 × $\frac{2 \text{ mol H}}{1 \text{ mol H}_20}$ = 5.8 mol H

2. How many H atoms are in 5.3 moles of CH_4 ?

5.3 mol CH₄ ×
$$\frac{4 \text{ mol H}}{1 \text{ mol CH}_4}$$
 = 21.2 mol H
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 eggs3. 750 sheets = 1.5 reams of paper
- 2. 4 score of years = **80** years
- 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 " \times 10 ^" (times 10 to the power of). *Use dimensional analysis for each one.*

- 5. 2.40×10^{20} atoms of sodium = 4.0×10^{-23} moles of sodium atoms 1 mol Na atoms 24 Na atoms \times – 6.022×10^{23} Na atoms 6. 15.0 moles of Pb²⁺ ions = 9.03×10^{24} Pb²⁺ ions 15.0 mol Pb²⁺ ions × $\frac{6.022 \times 10^{23}$ Pb²⁺ ions}{21} = 1 mol Ph²⁺ 7. 9.0×10^{23} atoms of silver = 1.5×10^{-21} moles of silver atoms 1 mol Ag atoms 900 Ag atoms \times — 6.022×10^{23} Ag atoms 8. 5.36 moles of $NH_3 = 3.23 \times 10^{24}$ molecules of NH_3 molecules 6.022 × 10²³ NH₃ molecules 5.36 mol NH₃ × – 1 mol NH₃ 9. 5.00×10^{22} molecules of chlorine (Cl₂) gas = 8.30 × 10⁻² moles of chlorine (Cl₂) molecules 1 mol Cl₂ molecules 5 molecules $Cl_2 \times 6.022 \times 10^{23}$ Cl₂ molecules 10. 4.00 moles of barium atoms = 2.41×10^{24} atoms of barium 4.00 mol Ba atoms × $\frac{6.022 \times 10^{23} \text{ Ba}}{-10^{23} \text{ Ba}}$ atoms 1 mol Ba atoms
- 11. Why is a mole such a big counting unit compared to ones we use every day (such as the dozen)? Atoms are extremely small.

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 H₂O? 18.016 g H₂O/mol 2 H = 2 × 1.008 = 2.016 10 \rightarrow = 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 ? ? mol $H_2O = 22.5 \text{ g } H_2O | 1 \text{ mol } H_2O = 1.25 \text{ mol } H_2O | 18.016 \text{ g } H_2O | 125 \text{ mol } H_2O |$
- 6. How many grams are in 0.250 moles of NaCl? ? g NaCl = 0.250 mol NaCl 58.44 g NaCl = 14.6 g NaCl
- 7. What is the molar mass of C₂H₅OH (ethanol)? 2 C $2 \times 12.01 = 24.02$ 6 H $6 \times 1.008 = 6.048$ 0 $= \frac{16.00}{46.068 \approx 46.07 \text{ g/mol}}$
- 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) ? mol $C_2H_5OH = 25.0 \text{ mol } C_2H_5OH = 0.785 \text{ g} = 19.625 \text{ g} + 1 \text{ mol } C_2H_5OH = 0.426 \text{ mol$

Activity: Moles within Moles





An elephant has one trunk and four legs.

A molecule of methane, CH₄ contains one carbon and four hydrogens.

- 1. An elephant has 4 legs and 1 trunk.
- 2. One **dozen** elephants have **4 dozen** legs and **1 dozen** trunks.
- 3. One dozen elephants is 12 elephants. (How many?)

See a relationship between the answers in #1 and #2? How they are similar and how they are different? Provide a sentence or two below. Use complete sentences.

Same numbers, different units.

Since each elephant has four legs and one trunk, then one dozen elephants will have four dozen legs and one dozen trunks.

- 4. A molecule of methane has 1 carbon atom(s) and 4 hydrogen atom(s).
- 5. A **MOLE** of methane has **1 mole(s)** of carbon atoms and **4 mole(s)** of hydrogen atoms.
- 6. A MOLE of methane is 6.022×10^{23} methane molecules. (How many?)
- 7. 0.25 moles of methane has 0.25 × 1 = 0.25 moles of carbon atoms and 0.25 × 4 = 1 moles of hydrogen atoms.
 0.25 mol CH₄ × 6.022 × 10²³ CH₄ molecules = 1.5 × 10²³
- 8. 0.25 moles of methane is 1.5×10^{23} methane molecules. (How many?)¹ mol CH₄ molecules
- 88 grams of carbon dioxide is 2 mole(s) of CO₂, which is 2 mole(s) of carbon atoms, and 4 mole(s) of oxygen atoms. Molar mass of CO₂ is ~44 g/mol.

NOTE: Do the moles of carbon atoms and moles of oxygen atoms add up to the moles of carbon dioxide? (They don't necessarily have to.) **No. 2 mol CO₂ \neq 2 mol C + 4 mol O**. The answer to the last problem is no. There is no Law of Conservation of Moles. Moles do not need to be conserved. Perform this calculation, using some information you already provided in the previous question (#9): There is the Law of Conservation of Mass though!

88 grams of carbon dioxide is also 24 grams of carbon atoms and 64 grams of oxygen atoms. Show your work here:

$$88 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol } \text{CO}_2}{44 \text{ g } \text{CO}_2} \times \frac{1 \text{ mol } \text{C}}{1 \text{ mol } \text{CO}_2} \times \frac{12. \text{ g } \text{C}}{1 \text{ mol } \text{C}} = 24 \text{ g } \text{C}$$

$$88 \text{ g } \text{CO}_2 \times \frac{1 \text{ mol } \text{CO}_2}{44 \text{ g } \text{CO}_2} \times \frac{2 \text{ mol } \text{O}}{1 \text{ mol } \text{CO}_2} \times \frac{16 \text{ g } \text{O}}{1 \text{ mol } \text{O}} = 64 \text{ g } \text{O}$$

NOTE: Do the grams of carbon atoms and grams of oxygen atoms add up to the grams of carbon dioxide? Should they? **88 g = 24 g + 64 g. Yes, because of Law of Conservation of Matter.**

- 11. 254 grams of carbon dioxide is 3.48×10^{24} molecules of carbon dioxide. In these molecules, 3.48×10^{24} are carbon atoms and 6.95×10^{24} are oxygen atoms.
- 12. 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 1 mol of NH_3 1 mol H_2O

```
√7.04 g < 18.02 g →
```

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

$$48 \text{ g } \text{CO}_2 \times \frac{1 \text{ mole } \text{CO}_2}{44 \text{ g } \text{CO}_2} \times \frac{1 \text{ mol } \text{C}}{1 \text{ mol } \text{CO}_2} = 1.09 \text{ mol } \text{C} \qquad 12 \text{ g } \text{C} \times \frac{1 \text{ mole } \text{C}}{12 \text{ g } \text{C}} = 1 \text{ mol } \text{C}$$

c. There are equal numbers of nitrogen atoms in one mole of NH₃ and one mole of N₂. False

$$1 \mod \mathsf{NH}_3 \times \frac{1 \mod \mathsf{N}}{1 \mod \mathsf{NH}_3} = 1 \mod \mathsf{N} \qquad 1 \mod \mathsf{N}_2 \times \frac{2 \mod \mathsf{N}}{1 \mod \mathsf{N}_2} = 2 \mod \mathsf{N}$$

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
 1 mol Cu

$$100 \text{ g Cu} \times \frac{1 \text{ mor Cu}}{63.55 \text{ g Cu}} = 1.57 \text{ mor Cu} \qquad 100 \text{ g CuO} \times \frac{1 \text{ mor CuO}}{79.55 \text{ g CuO}} \times \frac{1 \text{ mor Cu}}{1 \text{ mor CuO}} = 1.26 \text{ mor Cu}$$

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, NiCl₂. False

$$100 \text{ mol NiCl}_2 \times \frac{2 \text{ mol Cl}}{1 \text{ mol NiCl}_2} = 200 \text{ mol Cl atoms} \qquad 200 \text{ mol atoms} > 100 \text{ mol atoms}$$

f. There are more hydrogen atoms in 2 moles of NH₃ than in 2 moles of CH₄. False

$$2 \mod \text{NH}_3 \times \frac{3 \mod \text{H}}{1 \mod \text{NH}_3} = 6 \mod \text{H}$$

$$2 \mod \text{CH}_4 \times \frac{4 \mod \text{H}}{1 \mod \text{CH}_4} = 8 \mod \text{H} \qquad 6 \mod < 8 \mod \text{H}$$

Moles within Moles



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

EXAMPLES

- 1 mole of (NH₄)₂SO₄ = 2 moles of nitrogen
- 1 mole of (NH₄)₂SO₄ = 8 moles of hydrogen
- 1 mole of (NH₄)₂SO₄ = 1 mole of sulfur
- 1 mole of (NH₄)₂SO₄ = 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)

 $2 \text{ mol } \text{KNO}_3 \times \frac{3 \text{ mol } 0 \text{ atoms}}{1 \text{ mol } \text{KNO}_3} = 6 \text{ mol } 0 \text{ atoms}$

2. How many oxygen atoms are there in 2.50 moles of O_2 molecules? (3.01 × 10²⁴ atoms)

2.50 mol O₂ molecules × $\frac{2 \text{ mol O atoms}}{1 \text{ mol O}_2 \text{ molecules}}$ × $\frac{6.022 \times 10^{23} \text{ O atoms}}{1 \text{ mol O atoms}}$ = 3.01 × 10²⁴ O atoms

- 3. A mole of H_2O and a mole of O_2
 - a. have the same mass. 18.02 vs. 32 g
 - b. contain one molecule each. 1 mole = 6.022×10^{23} molecules \neq 1 molecule
 - c. have a mass of 1 g each.

(d.) contain the same number of molecules.

- 4. One molecule of sulfur contains 8 sulfur atoms. Then one mole of sulfur molecules will contain
 - a. 8 g of sulfur.
 - (b.) 8 moles of sulfur atoms.
 - c. 6.02×10^{23} sulfur atoms.
 - d. 8 sulfur atoms.
- 5. How many oxygen atoms are there in 48.0 g of 0 and 48.0 g of O_2 ?

48.0 g O atoms ×	1 m	ol O atoms	6.022 × 1	0 ²³ 0 atoms	_ 1 81 v 1	n ²⁴	Ω atoms		
	16.0	0 g O atoms î	1 mol	0 atoms	- 1.01 × 10 (o atoms		
/2 0 a 0 molecules ×		1 mol O ₂ mole	cules	2 mol 0	atoms		6.022 × 10 ²³	³ O atoms	-1.81×10^{24} O atoms
	69 × 1	32.00 g O ₂ mol	ecules	1 mol 0 ₂ m	olecules	Ŷ	1 mol 0	atoms	

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,	$\longleftrightarrow \longrightarrow$	malaa		mass
molecules, formula units)		mores	$\leftarrow \rightarrow$	(grams)

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

Molar mass Mg = 24.31 g/mol

10.00 g Mg × $\frac{1 \text{ mol Mg}}{24.31 \text{ g Mg}}$ = 0.41135 mol Mg = 0.4114 mol Mg

How many grams of iron are there in 34.77 moles of iron?
 Molar mass Fe = 58.55 g/mol

34.77 mol Fe × $\frac{55.85 \text{ g Fe}}{1 \text{ mol Fe}}$ = 1941.9 g Fe = 1942 g Fe

How many moles are there in 3.493 grams of lye, NaOH?
 Molar mass NaOH = 22.99 + 16.00 + 1.01 = 40.00 g/mol

3.493 g NaOH × $\frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}}$ = 0.087325 mol NaOH = (0.08733 mol NaOH)

How many grams are there in 0.275 moles of Na₃(PO₄)₂?
 Molar mass Na₃(PO₄)₂ = (3 × 22.99) + (2 × 30.97) + (8 × 16.00) = 258.91 g/mol

$$0.275 \text{ mol } \text{Na}_3(\text{PO}_4)_2 \times \frac{258.91 \text{ g } \text{Na}_3(\text{PO}_4)_2}{1 \text{ mol } \text{Na}_3(\text{PO}_4)_2} = 71.200 \text{ g} = (71.2 \text{ g } \text{Na}_3(\text{PO}_4)_2)$$

 How many molecules are in 7.47 moles of NH₃? How many total atoms are there? Molar mass NH₃ = 14.01 + 3(1.01) = 17.04 g/mol

7.47 mol NH₃ ×
$$\frac{6.022 \times 10^{23} \text{ NH}_3 \text{ molecules}}{1 \text{ mol NH}_3} = 4.498 \times 10^{24} = 4.50 \times 10^{24} \text{ NH}_3 \text{ molecules}$$

7.47 mol NH₃ ×
$$\frac{6.022 \times 10^{23} \text{ NH}_3 \text{ molecules}}{1 \text{ mol NH}_3} \times \frac{4 \text{ atoms}}{1 \text{ NH}_3 \text{ molecule}} = 1.799 \times 10^{25} = 1.80 \times 10^{25} \text{ atoms}$$

6. What is the mass in grams of 1.000 × 10¹² (1.000 trillion) atoms of gold?
 Molar mass Au = 196.97 g/mol

 $1.000 \times 10^{12} \text{ Au atoms} \times \frac{1 \text{ mol Au}}{6.022 \times 10^{23} \text{ Au atoms}} \times \frac{196.97 \text{ g Au}}{1 \text{ mol Au}} = 3.2708 \times 10^{-10} = (3.271 \times 10^{-10} \text{ g Au})$

7. How many grams of ammonium nitrate, NH_4NO_3 , contain 8.3×10^{25} formula units? Molar mass $NH_4NO_3 = (2 \times 14.01) + (4 \times 1.01) + (3 \times 16.00) = 80.06$ g/mol

 $8.3 \times 10^{25} \text{ NH}_{4} \text{NO}_{3} \text{ formula units} \times \frac{1 \text{ mol NH}_{4} \text{NO}_{3}}{6.022 \times 10^{23} \text{ NH}_{4} \text{NO}_{3} \text{ formula units}} \times \frac{80.06 \text{ g NH}_{4} \text{NO}_{3}}{1 \text{ mol NH}_{4} \text{NO}_{3}}$ $= 11,034.5 \text{ g NH}_{4} \text{NO}_{3} = (1.1 \times 10^{4} \text{ g NH}_{4} \text{NO}_{3})$

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

 $0.034 \text{ g H}_{2}0 \times \frac{1 \text{ mol H}_{2}0}{18.02 \text{ g H}_{2}0} \times \frac{6.022 \times 10^{23} \text{ H}_{2}0 \text{ molecules}}{1 \text{ mol H}_{2}0} = 1.136 \times 10^{21} = (1.1 \times 10^{21} \text{ H}_{2}0 \text{ molecules})$

9. How many atoms of oxygen are there in 0.0327 moles of Na_2SO_4 ? (**HINT**: There are 4 moles of 0 in each mole of Na_2SO_4 .)

 $0.0327 \text{ mol Na}_2 \text{SO}_4 \times \frac{4 \text{ mol } 0}{1 \text{ mol Na}_2 \text{SO}_4} \times \frac{6.022 \times 10^{23} \text{ 0 atoms}}{1 \text{ mol } 0} = 7.877 \times 10^{22} = 7.88 \times 10^{22} \text{ 0 atoms}$

10. 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?)

 $35.0 \text{ mol } (\text{NH}_4)_2 \text{SO}_4 \times \frac{2 \text{ mol N}}{1 \text{ mol } (\text{NH}_4)_2 \text{SO}_4} \times \frac{6.022 \times 10^{23} \text{ N atoms}}{1 \text{ mol N}} = 4.2154 \times 10^{25} = 4.22 \times 10^{25} \text{ N atoms}$

CHEM& 121 WORKBOOK

UNIT 10 CHEMICAL REACTIONS AND STOICHIOMETRY



Notes: Reaction Rates and Energies



COLLISION THEORY

Chemical equations show us what reacts and what is produced in a chemical reaction, but how exactly do reactions take place? In order for a reaction to occur:

1. Collisions must occur. Molecules must collide.

Concentration of reactants The amount of a substance per volume.



(more collisions)

Physical state of reactants **Gases (more collisions) > liquids > solids**.

2. Collisions must be of sufficient energy.

Activation energy Minimum amount of E required for reaction to occur.

Temperature Measure of the average kinetic energy of particles.

3. Sometimes the colliding particles need to be oriented in a specific way.

EXAMPLE Proper orientation leads to reaction.

(Interesting Question: What major classification of reaction can occur without satisfying rule #1? Synthesis, **Decomposition**, Single-replacement, Double-replacement? See pg. 154.)

No collisions. (Only one reactant!)

FACTORS AFFECTING REACTION RATES

The rate of a reaction is the speed at which a reaction takes place. It is determined by measuring the change in concentration of a reactant or product in a specific amount of time.

Concentration ↑ number of collisions, ↑ reaction time
 Temperature ↑ temperature, ↑ number of collisions, ↑ reaction time
 Use of catalyst lowers the activation energy of a reaction which ↑ rate
 Surface area ↑ surface area, ↑ number of collisions, ↑ rate

ENERGY CHANGES IN REACTIONS

Chemical reactions (and physical changes like melting and boiling) are accompanied by energy transfer between the system and the surroundings.

Exothermic	Endothermic		
Energy is released.	Energy is absorbed.		
Exo = "exit"	Endo = "into"		
Ex. Gas $ ightarrow$ liquid $ ightarrow$ solid (physical change)	Ex. Solid \rightarrow liquid \rightarrow gas (physical change)		
Combustion (making fire!) (chemical change)	Cold pack (chemical change)		

Chemical reactions (and physical processes) can occur both in a forward and reverse direction. The reverse reaction/process has an equal, but opposite energy change.

REACTION ENERGY DIAGRAMS

Energy diagrams visually display the energy relationships in a given reaction. Typical components of energy diagrams include:

- Axes: Energy on y-axis, Reaction pathway (progress) on x-axis
- Energy of the reactants; Energy of the products
- Overall change in energy for the reaction (ΔE)
- Activation energy (AE)
- Transition state (TS) The structure of the reactants as it changes into products.



Reaction Rates and Energies



- What are three main things that can increase the rate of a chemical reaction?
 ↑ concentration, use a catalyst, ↑ temperature
- Give a definition for chemical equilibrium in your own words.
 Chemical equilibrium refers to a state of a reversible reaction in which the rate of the forward reaction equals the rate of the reverse reaction.
- Consider the following reaction: Ca + 2 HBr → CaBr₂ + H₂ + Heat. Is the reaction endothermic or exothermic? How can you tell?
 Exothermic, heat is a product (released).
- 4. Drawing Energy Diagrams
 - a. Draw an energy diagram for the generic chemical reaction $A + B \rightarrow C$ that has the following characteristics: exothermic and fast reaction. Label the axes, reactants, transition state, products, activation energy, and heat of reaction on your diagram.
 - b. How would the energy diagram be different if the reaction was a slow reaction?
 - c. How would the energy diagram be different if a catalyst was used?



5. How would each of the following change the rate of the following reaction?

 $2 \text{ Mg} + 0_2 \rightarrow 2 \text{ MgO} + \text{Heat}$

- a. The reaction is cooled to 0 °C. Faster Slower No change
- b. The reaction is done in an atmosphere of pure O_2 instead of in air (air is a mixture of gases that include O_2). Faster Slower No change

Higher concentration of O₂.

Notes: Balancing Chemical Equations



The phrase "two plus two equals four" can be expressed using a mathematical equation as 2 + 2 = 4. Similarly, in chemistry, substances that react together to make new substances as part of a chemical change can be expressed using symbols in a chemical equation. Here are the main components:

- Chemical formulas are used to represent the substances involved in the reaction.
 - Substances present at the start: \rightarrow chemical arrow
 - Substances present at the end:
- Instead of an equal sign, use ...
 - General form: $\mathbf{A} \rightarrow \mathbf{B}$
- When more than 1 substance is present on either "side" of the reaction equation, use +. $A + B \rightarrow C + D$
- The order in which the substances are written, on either side of the equation, does not matter.
- Include the physical state (in parentheses) of each substance.
 - Solid (s)
 Liquid (l)
 Gas (g)
 Aqueous solution (aq)
- Equations can be written for chemical changes (chemical reactions) or physical changes.
 A (s) + B (I) → AB (s) A (s) → A (I)

EXAMPLE

Expressing a "word equation" as a chemical equation:

Nitrogen dioxide, a common air pollutant, is produced in combustion engines when nitrogen gas reacts with oxygen gas at very high temperatures.

```
First attempt at the chemical equation:

N_2(g) + O_2(g) \rightarrow NO_2

have 2 N atoms only 1 N atom
```

Is our first attempt at the equation reasonable in terms of the types and numbers of atoms present? Does our first attempt follow the Law of Conservation of Matter?

LAW OF CONSERVATION OF MATTER

An atom in a chemical reaction doesn't change identity. Atoms are neither created nor destroyed; they're just rearranged!

Instead of changing chemical formulas, we indicate how many "multiples" of a chemical formula are needed to balance/conserve atoms. The "multiples" are whole numbers written before the chemical formula, and we call these **coefficients**.

MICROSCOPIC VS. MACROSCOPIC VIEWPOINTS

On the microscopic scale, the coefficients refer to molecules, but on the macroscopic scale the coefficients refer to moles.

MAIN IDEAS OF BALANCING CHEMICAL EQUATIONS

- 6. Do not add new chemicals to the equation. Why?
- 7. Do *not* change chemical formulas (don't change the subscripts). Why?
- 8. Balance the equation by using coefficients placed before the different chemical formulas. Reduce the coefficients to the lowest whole number ratio (2, 4, 6 reduce to 1, 2, 3).

STEP-BY-STEP APPROACH FOR BALANCING CHEMICAL EQUATIONS

- 1. Place a "1" in front of the formula with the largest number of atoms. (Tie breaker: choose formula with greater number of elements.) Let's call this the "starting formula."
- 2. Insert coefficients that balance the elements in the starting formula that appear in other *compounds* (have more than one element present); for now, skip single-element only formulas.

Suggested approach:

- a. Start with elements in the starting formula that are only in one other compound.
- b. Then, deal with all other elements from the starting formula.
- c. Lastly, deal with all other elements in compounds.
- 3. Place coefficients in front of formulas of uncombined elements that balance those elements.

- 4. Two options when dealing with coefficients:
 - a. If you get into an "odd/even" situation where there are an odd number of one type of atom on one side and an even number of that type of atom on the other side, go back to the starting formula and double the coefficient, then double all other coefficients established thus far.
 - b. Use fractions if necessary. Clear any fractions at the end of the process by multiplying all coefficients by the lowest common denominator.
- 5. What to do in the case of fractions:

 $1.5 \ 0_2 \rightarrow 2(1.5 \ 0_2) \rightarrow 3 \ 0_2$

- 6. Reduce coefficients to simplest set of whole numbers. Remove any "1" coefficients.
- 7. Double check that the equation is now balanced.

Example:

LET'S TRY A FEW

2 NO (g) + 1 O ₂ (g)	$\gamma) \rightarrow 2$	NO ₂ (g)
	21	Ν	1 2
	4 3	0	24
$2 \operatorname{KClO}_3(s) \to 2$	KCI (<i>g</i>)	+ 3 0 ₂	₂ (g)
	2 1	K	1 2
	2 1	CI	1 2
	6 3	0	2 6

Aqueous lead (II) acetate reacts with aqueous potassium iodide to form solid lead (II) iodide and aqueous potassium acetate.

 $Pb(CH_3COO^-)_{2(aq)} + 2 KI(aq) \rightarrow PbI_{2(s)} + 2 K(CH_3COO^-)_{aq}$

Balancing Chemical Equations



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 H₂, N₂, O₂, F₂, Cl₂, Br₂, I₂. 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.
- 1. **2** Cu + **1** $0_2 \rightarrow$ **2** CuO
- 2. $1 S_8 + 24 F_2 \rightarrow 8 SF_6$
- 3. **1** P_4O_{10} + **6** $H_2O \rightarrow$ **4** H_3PO_4
- 4. $1 \text{ZnS} + 2 \text{HCI} \rightarrow 1 \text{ZnCI}_2 + 1 \text{H}_2\text{S}$
- 5. **1** C_5H_{12} + **8** $O_2 \rightarrow$ **5** CO_2 + **6** H_2O
- 6. $1 C_2 H_5 OH + 3 O_2 \rightarrow 2 CO_2 + 3 H_2 O$
- 7. $2 \text{ NH}_3 + 1 \text{ O}_2 \rightarrow 2 \text{ NO} + 3 \text{ H}_2$
- 8. **2** $\operatorname{Fe}_2 S_3 + 9 O_2 \rightarrow 2 \operatorname{Fe}_2 O_3 + 6 \operatorname{SO}_2$
- 9. **2** C_6H_6 + **15** $O_2 \rightarrow$ **12** CO_2 + **6** H_2O
- 10. $2 \text{ NaN}_3 \rightarrow 2 \text{ Na} + 3 \text{ N}_2$
- 11. $_$ Na₂SO₄ + $_$ KCl \rightarrow $_$ K₂SO₄ + $_$ NaCl

Balancing Word Equations



1. Solid lithium carbonate reacts with aqueous hydrochloric acid (HCI) to produce liquid water, carbon dioxide gas, and aqueous lithium chloride.

```
Li_2CO_3 + 2 HCI \rightarrow 2 LiCI + 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 (CH₃CH₂OH or C₂H₅OH) will burn in oxygen gas to produce liquid water and carbon dioxide gas.

 $\mathrm{C_2H_5OH} + 3~\mathrm{O_2} \rightarrow 3~\mathrm{H_2O} + 2~\mathrm{CO_2}$

4. Hydrogen gas and oxygen gas can be burned, under certain conditions, to form liquid hydrogen peroxide (officially called dihydrogen dioxide).

 $1 \text{ H}_2 + 1 \text{ O}_2 \rightarrow 1 \text{ H}_2 \text{ O}_2$

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.

 $AI(NO_3)_3 + 3 \text{ KF} \rightarrow 3 \text{ KNO}_3 + \text{ AIF}_3$

7. Solid silicon and oxygen gas can react to form a solid disiliconhexaoxide.

 $2 \text{ Si} + 3 \text{ } \text{ } 0_2 \rightarrow \text{ } \text{ } \text{Si}_2 \text{ } \text{ } 0_6$

8. Aqueous ammonium chloride will react with aqueous calcium hydroxide to form aqueous calcium chloride and aqueous ammonium hydroxide.

```
2 \text{ NH}_4\text{Cl} + \text{Ca}(\text{OH})_2 \rightarrow \text{CaCl}_2 + 2 \text{ NH}_4\text{OH}
```

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 (HCI) to form a precipitate of silver chloride and hydrogen gas.

 $2 \text{ Ag} + 2\text{HCI} \rightarrow 2 \text{ AgCI} + \text{H}_2$

Tutorial: Classifying Types of Reactions



- 2. Which of the equations above describe each type of reaction? List all that apply.
 - a. Synthesis **a**, **d**, **h** d. Double-replacement **c**, **e**, **g**
 - b. Single-replacement **j** e. Combustion **f**
 - c. Decomposition **b**, **i**

Activity: Redox Reactions



WHAT IS A REDOX REACTION?

Redox is an abbreviation for "oxidation-reduction" reactions. These reactions involve the transfer of electrons. Oxidation and reduction are opposite processes and are always paired (when one occurs, the other must also occur). OIL RIG is a memory device to remember the definitions of oxidation and reduction.

Memory Device: OIL RIG

Oxidation = loss of electrons (e⁻)

Oxidation **I**s **L**oss of e⁻

Reduction = gain of electrons (e⁻)

<u>**R</u>eduction <u>I</u>s <u>G**ain of e⁻</u></u>

A. Na \rightarrow Na⁺ + e⁻

B. $Mg^{2+} + 2e^- \rightarrow Mg$

- 1. Are the electrons that appear on the products side of Equation A gained or lost by sodium?
- 2. Are the electrons that appear on the reactants side of Equation B gained or lost by magnesium ion?
- 3. Which equation represents an oxidation reaction? (A) or B **HINT**: Use OIL RIG
- 4. Which equation represents a reduction reaction? A or **B HINT**: Use OIL RIG

IDENTIFY WHETHER A REACTION IS A REDOX REACTION OR NOT

Redox reactions involve transfers of electrons, so neutral atoms and molecules which become ions or vice versa indicate redox has occurred. This is common in synthesis, decomposition, single-replacement, and combustion reactions.

For example, $2 \operatorname{Na}(s) + \operatorname{MgCl}_2(aq) \rightarrow 2 \operatorname{NaCl}(aq) + \operatorname{Mg}(s)$

- 5. What type of reaction is this? Single replacement
- 6. Write the charge of Na (*s*) and Mg (*s*) as zero above their symbol in the equation above.
- 7. Write the charge of Mg (in MgCl₂) above its symbol. Do the same with Na (in NaCl).
- 8. When Na (s) turns into NaCl, (circle the correct choice)
 - a. Na gains / loses electrons. b. Na is oxidized / reduced.

- 9. When Mg (in MgCl₂) turns into Mg (s),
 - a. Mg gains / loses electrons. b. Mg is oxidized / reduced.
- 10. The species that is oxidized is called the "reducing agent." Which reactant is the reducing agent in this reaction? Na (s) / MgCl₂ Na is oxidized and called the reducing agent.
- The species that is reduced is called the "oxidizing agent." Which reactant is the oxidizing agent in this reaction? Na (s) / MgCl₂ Mg²⁺ (in MgCl₂) is reduced and called the oxidizing agent.

Answer the following questions about this reaction: $BaS + PtO \rightarrow BaO + PtS$ +2-2 + 2-2 +

- 12. Write the charge of each element above its symbol.
- 13. Do the charges of any of the elements change from reactants side to products side? **No**—charges
- 14. Has there been a transfer of electrons? **No** Is this a redox reaction? **No stay the same.**
- 15. This is a double replacement reaction. Are double replacement reactions also redox reactions? No

Redox occurs in many processes in biology, such as the electron transport chain and photosynthesis. There is another way to think of redox for biological/organic molecules:

Oxidation is a gain of oxygen or loss of hydrogen. **Reduction** is a gain of hydrogen or loss of oxygen.

16. Based on these definitions, determine if the following molecules are oxidized [0] or reduced [H]:



17. In biology, the cofactor NAD+ turns into NADH. Is NAD+ oxidized / reduced?

NAD+ gained H In a combustion reaction, the fuel (usually a hydrocarbon) is oxidized by the oxygen present. In this reaction, $C_5H_{12}(I) + 8 O_2(g) \rightarrow 5 CO_2(g) + 6 H_2O(g)$, which reactant is **Oxidized Reduced** 18. being oxidized? C_5H_{12}/O_2 Turns into C_0_2/H_2O . 19. being reduced? C_5H_{12}/O_2 20. the oxidizing agent? C_5H_{12}/O_2O_2 is oxidizing C_5H_{12} . 21. is the reducing agent? C_5H_{12}/O_2O_2 C_5H_{12}/O_2O_2 .

Notes: Stoichiometry



1 sandwich includes the following:







2 Slices of Bread

3 Slices of Tomato

1 Piece of Lettuce

Basic Problem: If you have an unlimited supply of tomato and lettuce, but only 12 slices of



Theoretical Yield: If you have 12 slices of bread, 4 pieces of lettuce, and 6 slices of tomato, how many sandwiches can you make? **Let's see! (below)**

Limiting Reactant: What ingredient limits the number of sandwiches that you can make?



Theoretical Yield: The *calculated* amount of product that can theoretically be made based on the amount of the limiting reactant available. **Tomato is the limiting reactant since** only 2 sandwiches can be made!

% **Yield:** So far we've been focusing on how many sandwiches we could make ... but what if our 20-lb kitty eats one of our sandwiches? Using the previous example (12 bread, 4 lettuce, 6 tomato), let's calculate the % yield ...

% Yield
$$\frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100$$

% Yield = $\left(\frac{1 \text{ sandwich}}{2 \text{ sandwiches}}\right) \times 100\% = 50\% \text{ yield}$

IN THE LANGUAGE OF CHEMISTRY ...

A generic chemical equation: $2A + B + 3C \rightarrow D$ Mole ratios! (Like a recipe!) Some possible conversion factors: $\frac{2 \mod A}{1 \mod B}$ $\frac{2 \mod A}{3 \mod C}$ $\frac{1 \mod B}{1 \mod D}$ 1 mol B/3 mol C 2 mol A/1 mol D, etc.!

If you want to make 5.7 moles of D, how many moles of C do you need?

5.7 mol D $\left(\frac{3 \text{ mol C}}{1 \text{ mol D}}\right)$ = 17.1 mol C is needed (round to 17 mol C)

If you have 2 mol of A and 4 mol of C, how many moles of D can you make? What is the limiting reactant?



In the lab we don't measure moles, but we can weigh items in grams. We need to be able to transition between grams and moles, as well as from moles to grams and use the molar coefficients in a balanced equation to correlate moles of one substance with moles of another substance.

EXAMPLES

How many grams of CO_2 will be produced by burning 66.0 g C_7H_{16} (heptane)? (**HINT:** Start by writing a balanced equation for the combustion of heptane.)

$$\begin{array}{c} C_{7}H_{16} + \underline{11} \ 0_{2} \rightarrow \underline{7} \ CO_{2} + \underline{8} \ H_{2}O \\ \\ \hline 66.0 \ g \ C_{7}H_{16} \ \left(\begin{array}{c} 1 \ \text{mole} \ C_{7}H_{16} \\ \hline 100.23 \ g \ C_{7}H_{16} \end{array} \right) \left(\begin{array}{c} 7 \ \text{mole} \ CO_{2} \\ \hline 1 \ \text{mol} \ C_{7}H_{16} \end{array} \right) \left(\begin{array}{c} 44.01 \ g \ CO_{2} \\ \hline 1 \ \text{mole} \ CO_{2} \end{array} \right) \end{array} \right) \\ \hline \end{array} \\ \hline \end{array} \\ \begin{array}{c} \hline 203 \ g \ CO_{2} \end{array} \\ \hline \end{array}$$

How many grams of heptane can be burned if 300. g O₂ is available?

$$300 \text{ g O}_2 \left(\frac{1 \text{ mole O}_2}{32.00 \text{ O}_2} \right) \left(\frac{1 \text{ mole C}_7 \text{H}_{16}}{11 \text{ moles O}_2} \right) \left(\frac{100.23 \text{ g C}_7 \text{H}_{16}}{1 \text{ mole C}_7 \text{H}_{16}} \right) \in 85.4 \text{ g C}_7 \text{H}_{16}$$

How many moles of CO_2 can be made from 300. g O_2 ? How many grams of CO_2 is this? $300 \text{ g } O_2 \left(\frac{1 \text{ mole } O_2}{32.00 \text{ g } O_2} \right) \left(\frac{7 \text{ moles } CO_2}{11 \text{ moles } O_2} \right) = 6 \text{ moles } CO_2$ $6 \text{ moles } CO_2 \left(\frac{44.01 \text{ g } CO_2}{1 \text{ mole } CO_2} \right) = 263 \text{ g } CO_2 \text{ (or 300 g if one s.f.)}$

What if only 176 grams of CO₂ was actually produced from 300. g O₂? What is the % yield? % Yield = $\left(\frac{176 \text{ g CO}_2 \text{ (actual yield)}}{263 \text{ g CO}_2 \text{ (theoretical yield)}}\right) \times 100\% = 66.9\% \text{ yield}$

The Mole House



A TOOL FOR STOICHIOMETRY



Stoichiometry 1



1. Ammonia gas reacts with oxygen gas according to the following equation:

```
4 \text{ NH}_3 + 5 \text{ O}_2 \rightarrow 4 \text{ NO} + 6 \text{ H}_2\text{O}
```

a. How many moles of oxygen gas are needed to react with 23 moles of ammonia? (29 mole)

 $\frac{23 \text{ mol NH}_3}{4 \text{ mol NH}_3} = 29 \text{ mol O}_2$

b. How many grams of NO are produced when 25 moles of oxygen gas react with an excess of ammonia? $(6\overline{0}0 \text{ g})$

```
\frac{25 \text{ mol } \text{O}_2}{5 \text{ mol } \text{O}_2} \left(\frac{4 \text{ mol } \text{NO}}{5 \text{ mol } \text{O}_2}\right) \left(\frac{30.01 \text{ g } \text{NO}}{1 \text{ mol } \text{NO}}\right) = 600.2 \text{ g } \text{NO} \rightarrow 600 \text{ g } \text{NO} (1 \text{ s.f.})
```

 c. If 24 grams of water are produced, how many moles of nitrogen monoxide are formed? (0.89 mole)

 $\frac{24 \text{ g H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \left(\frac{4 \text{ mol NO}}{6 \text{ mol H}_2\text{O}}\right) = 0.89 \text{ mol NO}$

2. The compound calcium carbide, CaC₂, is made by reacting calcium carbonate with carbon at high temperatures. The UNBALANCED EQUATION for the reaction is

$$\mathbf{2} \operatorname{CaCO}_3 + \mathbf{5} \operatorname{C} \rightarrow \mathbf{2} \operatorname{CaC}_2 + \mathbf{3} \operatorname{CO}_2$$

- a. Balance the equation.
- b. How many moles of carbon are required to produce 5.0 moles CO₂? (8.3 mole)

$$5.0 \text{ mol } \text{CO}_2 \left(\frac{5 \text{ mol } \text{C}}{3 \text{ mol } \text{CO}_2} \right) = 8.3 \text{ mol } \text{C}$$

c. How many grams of calcium carbide are produced when 4.00 moles of carbon react with an excess of calcium carbonate? (102 g)

4.0 mol C
$$\left(\frac{2 \text{ mol CaC}_2}{5 \text{ mol C}}\right) \left(\frac{64.10 \text{ g CaC}_2}{1 \text{ mol CaC}_2}\right) = 102.56 \rightarrow 100 \text{ or } 1.0 \times 10^2 \text{ g CaC}_2$$

d. How many moles of carbon dioxide are produced when 55 grams of calcium carbonate react with an excess of carbon? (0.83 mole)

 $\frac{55 \text{ g CaCo}_3}{100.09 \text{ g CaCO}_3} \left(\frac{1 \text{ mole CaCO}_3}{2 \text{ mol CaCO}_3}\right) \left(\frac{3 \text{ mol CO}_2}{2 \text{ mol CaCO}_3}\right) = 0.82 \text{ mole CO}_2$

e. How many grams of carbon are needed to react with 453 grams of calcium carbonate? (136 g)

 $453 \text{ g CaCO}_3 \left(\frac{1 \text{ mole CaCO}_3}{100.09 \text{ g CaCO}_3}\right) \left(\frac{5 \text{ mol C}}{2 \text{ moles CaCO}_3}\right) \left(\frac{12.01 \text{ g C}}{1 \text{ mole C}}\right) = 136 \text{ g C}$

f. How many grams of calcium carbonate are needed to form 598 grams of calcium carbide? (934 g)

$$\frac{598 \text{ g CaC}_2}{64.10 \text{ g CaC}_2} \left(\frac{1 \text{ mole CaC}_2}{64.10 \text{ g CaC}_2} \right) \left(\frac{2 \text{ mol CaCO}_3}{2 \text{ mol CaC}_2} \right) \left(\frac{100.09 \text{ g CaCO}_3}{1 \text{ mol CaCO}_3} \right) = 934 \text{ g CaCO}_3$$

Stoichiometry 2



1. For the given combustion reaction of **octane**, **C**₈**H**₁₈, answer the following questions:

(Answers to the questions are given in parenthesis.)

 $\mathbf{2}~\mathbf{C_8H_{18}+25}~\mathbf{0_2} \rightarrow \mathbf{16}~\mathbf{CO_2+18}~\mathbf{H_2O}$

a. Write 5 possible molar ratios from this equation.

2 mol C ₈ H ₁₈	2 mol C ₈ H ₁₈	25 mol O ₂	16 mol CO ₂	2 mol C ₈ H ₁₈
$25 \text{ mol } 0_2$	16 mol CO_2	16 mol CO_2	18 mol H ₂ O	18 mole H_2O

b. How many moles of CO_2 would be produced by reacting 0.67 moles of octane with excess of oxygen? (Amount of oxygen is not involved in the calculation.) (5.4 mol CO_2)

 $0.67 \text{ mol } C_8H_{18} \left(\frac{16 \text{ mol } CO_2}{2 \text{ mol } C_8H_{18}}\right) = 5.4 \text{ mol } CO_2$

c. How many moles of H_2O would be produced by reacting 0.67 moles of octane with excess of oxygen? (6.0 mol H_2O)

 $0.67 \text{ mol } C_8 H_{18} \left(\frac{18 \text{ mol } H_2 0}{2 \text{ mol } C_8 H_{18}} \right) = 6.0 \text{ mol } H_2 0$

d. If we react 225 g of octane C_8H_{18} with oxygen, how many moles of O_2 are required? (24.7 mol O_2)

 $225 \text{ g } C_8 \text{H}_{18} \left(\frac{1 \text{ mole } C_8 \text{H}_{18}}{114.26 \text{ g } C_8 \text{H}_{18}} \right) \left(\frac{25 \text{ mol } O_2}{2 \text{ mol } C_8 \text{H}_{18}} \right) = 24.6 \text{ mol } O_2$

e. If we react 225 g of octane C_8H_{18} with excess oxygen, how many moles of CO_2 are produced? (15.8 mol CO_2)

 $\frac{225 \text{ g } \text{C}_8 \text{H}_{18}}{114.26 \text{ g } \text{C}_8 \text{H}_{18}} \left(\frac{16 \text{ mol } \text{CO}_2}{2 \text{ mol } \text{C}_8 \text{H}_{18}} \right) = 15.8 \text{ mol } \text{CO}_2$

$2\ C_8 H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2 O$

f. If we react 225 g of octane C_8H_{18} with excess oxygen, how many moles of H_2O are produced? (17.8 mol H_2O)

 $\frac{225 \text{ g } \text{C}_8 \text{H}_{18}}{114.26 \text{ g } \text{C}_8 \text{H}_{18}} \left(\frac{18 \text{ mol } \text{H}_2 \text{O}}{2 \text{ mol } \text{C}_8 \text{H}_{18}} \right) = 17.7 \text{ mol } \text{H}_2 \text{O}$

g. If we wish to make 7.5 mol CO_{2,} how many grams of C₈H₁₈ will be used? (2 significant figures rounds to 110 g C₈H₁₈)

7.5 mol CO₂ $\left(\frac{2 \text{ mol } C_8 H_{18}}{16 \text{ mol } CO_2}\right) \left(\frac{114.26 \text{ g}}{1 \text{ mol } C_8 H_{18}}\right) = 107 \rightarrow 110 \text{ g} (2 \text{ s.f.})$

h. If we wish to make 7.5 mol $CO_{2,}$ how many grams of O_2 do we need? (380 g O_2)

7.5 mole $CO_2\left(\frac{25 \text{ mol } O_2}{16 \text{ mol } CO_2}\right)\left(\frac{32.00 \text{ g } O_2}{1 \text{ mol } O_2}\right) = 375 \rightarrow 380 \text{ g } O_2 (2 \text{ s.f.})$

i. If we wish to make 7.5 mol CO₂, how many grams of H₂O will be produced? (150 g H₂O) 7.5 mol CO₂ $\left(\frac{18 \text{ mol H}_2O}{16 \text{ mol O}_2}\right) \left(\frac{18.02 \text{ g H}_2O}{1 \text{ mol H}_2O}\right) = 152 \rightarrow 150 \text{ g} (2 \text{ s.f.})$

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)?



b. What is the limiting reactant? How can you tell?

Eggs \rightarrow limit us to only 36 cookies

2. Ammonia gas reacts with oxygen gas according to the following equation:

```
4 \text{ NH}_3 + 5 \text{ O}_2 \rightarrow 4 \text{ NO} + 6 \text{ H}_2 \text{O}
```

 Calculate the number of moles of oxygen gas needed to react exactly with 7 moles of NH₃.

 $\frac{7 \text{ mol NH}_3}{4 \text{ mol NH}_3} \left(\frac{5 \text{ mol } \mathbf{O}_2}{4 \text{ mol NH}_3} \right) = 8.75 \rightarrow 9 \text{ mol } \mathbf{O}_2$

b. If you had 7 moles of $\rm NH_3$ and 6 moles of oxygen gas, what would be the limiting reactant?

```
0<sub>2</sub>
```

7 moles of NH_3 – need 9 mol O_2 to react with 7 mol NH_3 .

```
6 moles of oxygen gas – only have 6 moles 0_2 \Rightarrow this will get used up first!
```

c. If you had 7 moles of $\rm NH_3$ and 10 moles of oxygen gas, what would be the limiting reactant?

NH₃

7 moles of NH_3 – limiting reactant.

10 moles of oxygen gas – plenty of $\rm O_2$ (only need 9 mol) will have 1 mol leftover (excess).

3. Aluminum reacts with aqueous sulfuric acid to produce an aluminum sulfate aqueous solution and hydrogen gas.

2 AI
$$(s)$$
 + 3 H₂SO₄ $(aq) \rightarrow$ AI₂(SO₄)₃ (aq) + 3 H₂ (g)

a. What would be the limiting reactant if you had 2 moles of Al and 2 moles of sulfuric acid? H₂SO₄

$$2 \mod AI \left(\frac{3 \mod H_2}{2 \mod AI}\right) = 3 \mod H_2$$

$$2 \mod H_2 SO_4 \left(\frac{3 \mod H_2}{3 \mod H_2 SO_4}\right) = 2 \mod H_2$$
(less, limited to 2 mol)

```
Pick one product to solve for ... e.g., H<sub>2</sub>.
```

b. What would be the limiting reactant if you had 3 moles of Al and 3 moles of sulfuric acid? H₂SO₄

 $3 \text{ mol Al} \left(\begin{array}{c} 3 \text{ mol H}_2 \\ \hline 2 \text{ mol Al} \end{array} \right) = 4.5 \text{ mol H}_2 \quad \begin{array}{c} 3 \text{ mol H}_2 \text{SO}_4 \\ \hline 3 \text{ mol H}_2 \text{SO}_4 \end{array} \left(\begin{array}{c} 3 \text{ mol H}_2 \\ \hline 3 \text{ mol H}_2 \text{SO}_4 \end{array} \right) = 3 \text{ mol H}_2 \\ (\text{less, limited to 3 mol}) \end{array}$

c. What would be the limiting reactant if you had 5 moles of Al and 7 moles of sulfuric acid? H₂SO₄

 $5 \text{ mol Al} \left(\frac{3 \text{ mol H}_2}{2 \text{ mol Al}}\right) = 7.5 \text{ mol H}_2 \qquad 7 \text{ mol H}_2 \text{SO}_4 \left(\frac{3 \text{ mol H}_2}{3 \text{ mol H}_2 \text{SO}_4}\right) = 7 \text{ mol H}_2 \text{(less, limited to 7 mol)}$

d. What would be the limiting reactant if you had 0.50 moles of Al and 0.75 moles of sulfuric acid? Niether

0.5 mol Al $\left(\frac{3 \text{ mol } \text{H}_2}{2 \text{ mol } \text{Al}}\right) = 0.75 \text{ mol } \text{H}_2$ 0.75 mol H₂SO₄ $\left(\frac{3 \text{ mol } \text{H}_2}{3 \text{ mol } \text{H}_2\text{SO}_4}\right) = 0.75 \text{ mol } \text{H}_2$

Neither Al nor H_2SO_4 is the limiting reactant! Both will get used up completely to make 0.75 mol H_2 .

4. Ammonia can be synthesized by the following reaction.

$$2 \text{ NO } (g) + 5 \text{ H}_2 (g) \rightarrow 2 \text{ NH}_3 (g) + 2 \text{ H}_2 0 (g)$$

a. If 10. moles of NO and 20. moles of hydrogen react, what is the theoretical yield (in moles) of ammonia?

 $\begin{array}{c} 10 \text{ moles NO} \\ \underset{(\text{will be leftover})}{\text{Excess reactant}} & \left(\begin{array}{c} 2 \text{ mol NH}_3 \\ \hline 2 \text{ mol NO} \end{array} \right) = 10 \text{ mol NH}_3 \end{array} \begin{array}{c} 20 \text{ mol H}_2 \\ \hline \left(\begin{array}{c} 2 \text{ mol NH}_3 \\ \hline 5 \text{ mol H}_2 \end{array} \right) \end{array} \begin{array}{c} \underbrace{\$ \text{ mol NH}_3}_{(\text{lower amount = theoretical yield)}} \end{array}$

b. If 0.50 moles of NO and 1.5 moles of hydrogen react, what is the theoretical yield (in moles) of ammonia?

 $0.5 \text{ mol NO} \left(\frac{2 \text{ mol NH}_3}{2 \text{ mol NO}}\right) \in \underbrace{0.5 \text{ mol NH}_3}_{\text{(lower amount = NO is the limiting reactant!}} 1.5 \text{ mol H}_2 \left(\frac{2 \text{ mol NH}_3}{5 \text{ mol H}_2}\right) = 0.6 \text{ mol NH}_3$

166 Unit 10

CHEM& 121 WORKBOOK

UNIT 11 SOLUTIONS



Solutions and Solubility



Polar

cipitation

- 1. What is a saturated, unsaturated, and supersaturated solution? homogeneous mixture
- 2. Solution = solute (s) + solvent
- 3. Give a solute and solvent for each of these common household products:
 - a. Orange juice Solute: Vitamin C Solvent: Water
 - b. Ammonia cleaner Solute: Ammonia (NH₃) Solvent: Water
 - Solute: Sodium chloride (NaCl) Solvent: Water c. Salt water
- 4. What does the phrase "like dissolves like" mean in terms of predicting the formation of solutions?

Polar solutes dissolve well in polar solvents.

Non-polar solutes dissolve well in non-polar solvents.

- 5. Which of the following compounds are soluble in water? Look up structures as needed.
 - a. C_5H_{12} Non-polar
- e. Ethanol, CH₃CH₂OH NH₃ Polar

Dissolution

- b. CaCl₂ lonic d. Glucose, $C_6H_{12}O_6$ Polar f. $C_{18}H_{36}O_2$ (stearic acid) Look at structure (lots of O—H bonds)
- 6. Which of the following pairs of compounds form a solution? Which of the following pairs of compounds are soluble in water? Look up structures as needed.
 - a. KCI and CCI₄ K⁺ CI[−] and C[−] C[−] CI lonic Non-polar CI b. 1-propanol and water Look up 1-propanol c. Cyclodecanone ($C_{10}H_{18}O$) and water Non-polar Pentane and hexane Non-polar Non-polar OH (polar)
- 7. A solid substance is added to water until a precipitate is formed. After some time, a dynamic equilibrium is established.
 - a. Label each arrow in the beaker by the process it represents: dissolution, precipitation
 - b. What is the rate of dissolution compared to the rate of precipitation at equilibrium? Equal rates
 - c. Use the concept of "dynamic equilibrium" to explain why the amount of precipitate stays constant. Since the rate of dissolution and precipitation are equal, the amount of precipitation stays the same.
 - d. Is this solution saturated or unsaturated? Explain.

Saturated. It is in equilibrium; rate of dissolving and precipitation is equal. If it were unsaturated, you would see no precipitate. Rate of dissolution > rate of precipitation.



Electrolytes



- 1. NaCl dissolves in water. **Ionic compounds; ions present in solution.**
 - a. Is the resulting solution an electrolyte / nonelectrolyte?
 - b. Draw a molecular-scale drawing of NaCl dissolved in water.



2. Methanol (CH₃OH) dissolves in water.

Covalent (molecular) compound

- a. Is the resulting solution an electrolyte / nonelectrolyte?
- b. Draw a molecular-scale drawing of methanol dissolved in water.



- 3. For the following substances,
 - a. Determine the type of compound—ionic or covalent.
 - b. What solute particles (what molecules or ions) will be found in aqueous solutions of the substance?
 - i. LiCl lonic/ions iv. K₂SO₄ lonic/ions
 - ii. Glucose, C₆H₁₂O₆ Covalent/molecules v. NaOH lonic/ions
 - iii. Ethanol, CH₃CH₂OH Covalent/molecules vi. CO₂ Covalent/molecules
- 4. Label each solution as a strong electrolyte, weak electrolyte, or nonelectrolyte.



SOLUBILITY

Factors affecting solubility:

- Temperature
- Pressure
- Physical agitation

How does temperature affect solubility in liquids (sugar syrup) vs. gases (soda)?

Exo Endo

Tutorial: Percent Concentration

Concentration refers to how much solute is in the total solution, often as a percentage:

% concentration =
$$\frac{solute}{solution} \times 100\%$$

- 1. What are the units of mass commonly used for everyday solutions? **q** Volume? mL
- 2. Write out the variations for the equation above for % (m/m), % (v/v), and % (m/v) including units.

 $\% (m/m) = \frac{g \text{ solute}}{g \text{ solution}} \times 100\% \qquad \% (m/v) = \frac{g \text{ solute}}{mL \text{ of solution}} \times 100\% \qquad \% (v/v) = \frac{mL \text{ solute}}{mL \text{ solution}}$ – × 100%

- 3. For each of the following, give an interpretation of the percent concentration as a fraction (solute / solution): mL/mL g/mL
 - a. Beer = Approx 5% (v/v) ethanol c. Whole Milk = 4% (m/m) fat 5 mL ethanol 100 mL beer g/mL
- 4 q fat 100 mL milk
 - b. Saline = 0.8 % (m/v) NaCl 0.8 g NaCl 100 mL saline
- 4. What is the % (m/m) for salt water (aqueous NaCl) in which 3.5 grams of salt has been added to 21.5 grams of water? (HINT: Masses are additive!)

% (m/m) = $\frac{\text{g NaCl}}{\text{g salt water}} \times 100\% = \frac{3.5 \text{ g NaCl}}{(3.5 \text{ g} + 21.5 \text{ g}) \text{ salt water}} \times 100\% = 14.\% \text{ m/m NaCl}$

Instructions: For these problems, write out the % as a fraction, and use dimensional analvsis to solve.

5. What is the mass (grams) of glucose in 250. mL of a 5.00 % (m/v) solution of glucose?

Rewrite as:	🔴 5.00 g glucose 🗋		250. mL solution	5.00 g glucose	
	100 mL solution) ٦		\rightarrow 100 mL solution	= 12.5 g glucose

6. What is the volume of wine required to ingest 1.00 mL of pure ethanol molecules? Wine is about 14.5% (v/v) ethanol in water.



7. A patient needs 100. grams glucose in 12 hours. How many liters of a 5.00% (m/v) glucose solution must be administered?

Rewrite as: $\left(\frac{5.00 \text{ g glucose}}{100. \text{ mL solution}}\right)$ = 2,00 mL = 2,00 mL = 2.00 L

Percent Concentration 1



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.



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?

7.6 g sucrose	× 100% = 8.35% (<i>m/m</i>)			
(7.6 g + 83.4 g) solution				

2. Calculate the volume percent, % (v/v), of solute in the following solution: 20.0 mL of methyl alcohol in enough water to give 475 mL of solution.

20.0 mL methyl alcohol × 100% (4.2% (v/v) 475 mL solution

- solute
- solution 3. You have 125 g of potassium sulfate and 325.6 L of solution. What is the % (m/v) concentration of your solution? % (m/v) is g/mL × 100 (Need to convert 325.6 L to mL!)



4. If 162.35 g aluminum hydroxide are dissolved in 6750 g of solution, what is the % (*m*/*m*) concentration of the solution?

 $\frac{162.35 \text{ g}}{(162.35 \text{ g} + 6750 \text{ g})} \times 100\% = 2.35\% (m/m)$

Need g of solute

5. Calculate the concentration in % (m/v) of a 450.0 mL solution containing 0.0762 moles of iodine (I_2) .

```
\frac{0.0762 \text{ mol}}{\left(\frac{253.8 \text{ g } \text{ I}_2}{1 \text{ mol } \text{ I}_2}\right)} = 19.3 \text{ g } \text{ I}_2
\frac{19.3 \text{ g } \text{ l}_2}{450.0 \text{ mL solution}} \times 100\% = 4.30\% (m/v)
```
Percent Concentration 2



When given a %, you can solve using the equations or you can use dimensional analysis if you think of the percentage as a conversion factor. Practice both ways and see which makes more sense to you.

Practice solving these using the % equations:

1. Normal saline solution that is used to dissolve drugs for intravenous use is 0.92% (m/v) NaCl in water. How many grams of NaCl are required to prepare 35.0 mL of normal saline solution?



2. Calculate the mass, in g, of glucose present in 400.0 mL of a solution that is 8.0% (m/v) glucose.



Practice solving these problems using the dimensional analysis method. You'll need to rewrite the percentage as a conversion factor.

3. How many grams of a 0.600% (*m*/*m*) penicillin solution should be given to a patient who requires 1.0 g of penicillin?

0.600% =	0.600 g penicillin	1.0 g penicillin (100. g solution	-166.7 a solution -170 a solution (2.61)
l	100. g solution		0.600 g penicillin	

4. A 500.-mL bottle of rubbing alcohol is labeled as having 70.% isopropanol by volume. Calculate how many milliliters of rubbing alcohol solution contain 5.0 mL of isopropanol.



Tutorial: Molarity



Concentration units can be expressed as a percentage.

Chemists often use concentration in units of molarity, called a molar (M)

Molarity (M) = <u>mol solute</u> 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.

1. Calculate the molarity of each solution:



 Calculate the molarity of a solution made by putting 15.5 g NaCl into a beaker and adding water to make 1.50 L of NaCl solution.
 (1) a → mol

15.5 g (í 1 mol	0.005	mol NaCl	0.265 mol	0 477 88
	58.44 g	= 0.265 mot Nact	L solution	= 1.50 L	

Molarity can be used as a conversion factor, as with percents. Try using molarity as the conversion factor in these problems and solve by dimensional analysis rather than solving from the equation.

3. An ethanol (C_2H_6O) solution has a concentration of 0.500 M. What **volume (in L)** of this solution would contain 8.0 **moles** of ethanol?



4. An ethanol (C_2H_6O) solution has a concentration of 0.500 M. What **volume (in mL)** of this solution would contain 8.0 grams of ethanol? Convert g mol

$$(1) \text{ Think of as:} \left(\begin{array}{c} 0.500 \text{ mol solute} \\ \hline 1000 \text{ mL} \end{array} \right) (2) 8.0 \text{ g } \text{C}_2 \text{H}_6 \text{O} \\ \hline 46.08 \text{ g } \text{C}_2 \text{H}_6 \text{O} \\ \hline \end{array} \right) = 0.174 \text{ mol} \begin{array}{c} (3) 0.174 \text{ mol} \\ \hline (0.500 \text{ mol} \\ \hline 0.500 \text{ mol} \\ \hline \end{array} \right) = 347 \text{ mL} \\ \text{solution} \\ \hline \end{array}$$

5. Determine how many grams of sucrose $(C_{12}H_{22}O_{11})$ are contained in 1.72 L of 0.758 M sucrose solution.

$$\underbrace{1 - \underbrace{0.758 \text{ mol}}_{L} (2) 1.72 \text{ L}}_{L} \left(\underbrace{\frac{0.758 \text{ mol}}{1 \text{ L}}}_{1 \text{ L}} \right) = 1.30 \text{ mol } C_{12}H_{22}O_{11} \xrightarrow{(3)} 1.30 \text{ mol } C_{12}H_{22}O_{11} \left(\underbrace{\frac{342.34 \text{ g}}{1 \text{ mol } C_{12}H_{22}O_{11}}}_{1 \text{ mol } C_{12}H_{22}O_{11}} \right) \in \underbrace{445. \text{ g } C_{12}H_{22}O_{11}}_{1 \text{ mol } C_{12}H_{22}O_{11}}$$

Molarity



Answers given in parentheses.

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? (0.75 M)



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? (2.0 M)

150 mL	(- 0 150 I	0.300 mol NaOH
	1000 mL	= 0.130 L	0.15 L 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? (0.500 M)

 $2.00 \text{ NaOH} \left(\frac{1 \text{ mol NaOH}}{40.00 \text{ g NaOH}}\right) = 0.0500 \text{ mol NaOH}$

 How many milliliters of 0.250 M NaOH are needed to provide 0.200 moles of NaOH? (800. mL)

$$\frac{0.250 \text{ mol NaOH}}{1000 \text{ mL NaOH}} \stackrel{\text{or solve for L and}}{\text{convert to mL}} M = \frac{\text{mol}}{\text{L}} \text{ or L} = \frac{\text{mol}}{\text{M}} = \frac{0.200 \text{ mol}}{0.250 \text{ mol/L}} 0.800 \text{ L} \qquad 0.200 \text{ mol} \left(\frac{1000 \text{ mL}}{0.250 \text{ mol}}\right) \in 800. \text{ mL}$$

5. How many grams of NaOH are there in 80.0 mL of 0.400 M NaOH solution? (1.28 g)

 $\begin{array}{c} 0.400 \text{ mol} \\ \hline 1000 \text{ mL} \end{array} \qquad \begin{array}{c} 80.0 \text{ mL} \\ \hline 1000 \text{ mL} \end{array} \left(\begin{array}{c} 0.400 \text{ mol} \\ \hline 1000 \text{ mL} \end{array} \right) \left(\begin{array}{c} 40.00 \text{ g} \\ \hline 1 \text{ mol} \end{array} \right) \left(\begin{array}{c} 1.28 \text{ g} \\ \hline \end{array} \right)$

6. How many moles of NaOH are contained in 150 mL of 3.00 M NaOH solution? (0.45 mol)

 3.00 MOL
 150 mL
 3.00 mol

 1000 mL
 1000 mL
 0.45 mol NaOH

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? (2.70 M)

50.0 mL $\left(\frac{1 \text{ L}}{1000 \text{ mL}}\right) = 0.0500 \text{ L}$ M = $\left(\frac{0.135 \text{ mol}}{0.0500 \text{ L}}\right) = 2.70 \text{ M}$

8. How many grams of KMnO₄ would you need to prepare 520 mL of 0.58 M solution? (48 g)

0.58 mol	520 mL ((0.58 mol)(158.04 g KMnO ₄	
1000 mL		1000 mL	1 mol	

9. How many milliliters of 5.3 M KBr do you need to get 0.112 moles of KBr? (21 mL)

5.3 mol	0.112 mol KBr	(1000 mL)	21.1 ml
1000 mL		5.3 mol	

Molarity: Applications



- 1. Consider a solution of MgCl₂ in water.
 - a. What is the solvent? Water
 - b. What is the solute? MgCl₂
 - 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.
 (and H₂0) Mg²⁺, Cl⁻—solutes
 - 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?

```
\frac{1.0 \text{ mol MgCl}_2}{1 \text{ L}} (1 \text{ L}) = 1.0 \text{ mol MgCl}_2 = \frac{1.0 \text{ mol Mg}^{2+}}{2.0 \text{ mol Cl}^- (\text{since there are 2 Cl}^- \text{ for every Mg}^{2+})}
```

- 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 Na₃PO₄? <u>1 mol Na₃PO₄</u> = 3.00 mol Na⁺ <u>3 mol Na⁺ ions</u>
 - b. What is the molarity of sodium in this beverage?

0.0020 mol Na ₃ PO ₄	(3 mol Na ⁺ `	M
1 L	1 mol Na ₃ PO ₄	

- 3. If 75 mL of 0.211 M NaOH is diluted to a final volume of 125 mL.
 - a. How many moles of NaOH are there?

 $\frac{75 \text{ mL}}{1000 \text{ mL}} = 0.016 \text{ mol NaOH}$

b. What is the final molarity of NaOH?

```
\frac{0.016 \text{ mol}}{125 \text{ mL}} \left( \begin{array}{c} 1000 \text{ mL} \\ \hline 1 \text{ L} \end{array} \right) = 0.13 \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?
 0.800 L (1.2 mol)
 €0.96 mol glucose
 - b. What volume of water should you add to dilute the concentration to 1.0 M?

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Tutorial: Dilutions

If you have made orange juice from frozen concentrate, you have performed a dilution. A dilution is the process of adding water to a solution, which results in a solution of lower concentration than the original solution (sometimes called a "stock" solution).

Add Water More particles (moles) for a given volume (more concentrated. less dilute). Less particles for a given volume (less concentrated, more dilute).

- 1. a. Draw in solute particles in the graduated cylinder. When poured into the volumetric flask and after water is added, does the spacing between the particles increase, decrease, or remain the same? Increase
 - b. Does this mean the concentration of the solution has increased, decreased, or remained the same after dilution? **Decreased**

The original solution is often called the "stock solution," and several dilutions can be made from it by removing a portion of the solution and adding water. The concentration of the stock solution is C_1 . The volume of stock solution used is V_1 . The final volume after water is added is V_2 . The resulting concentration (diluted solution) is C_2 .

Dilution: $C_1V_1 = C_2V_2$

Use the equation to solve the following dilution problems. Sometimes it is helpful to draw a diagram of graduated cylinder/volumetric flask to keep track of what is happening.

2. 100. mL of a 10.00 M solution was diluted to 250. mL. What is the molarity of the new solution? $C_1V_1 = C_2V_2$ $(100 \text{ mL})(10.0 \text{ M}) = C_2 (250. \text{ mL})$ $V_2 = 250. mL$ V₁ = 100 mL $C_2 = (100 \text{ mL})(10.0 \text{ M})$ $C_2 = ?$ 4.0 N (250. mL)



Dilutions

- V₂ C₂
 How do we prepare 200. mL of a 3.5 M solution of acetic acid if we have a bottle of 6.0 M acetic acid? Alternatively this question can be asked: How many mL or what volume of 6.0 M acetic acid would we measure to make 200. mL of a 3.5 M solution of acetic acid? (117 mL rounded to 120 mL)

 $V_1 = (200 \text{ mL})(3.5 \text{ M})$ = 117 mL

Take 117 mL of 6.0 M solution and add H_2O to a final total volume of 200. mL. The resulting solution will be 3.5 M.

How would you prepare 500. mL of a 0.30 M HCl solution from a concentrated solution of hydrochloric acid? Concentrated hydrochloric acid is 12¹M.

 $V_1 = (500 \text{ mL})(0.30 \text{ M})$ = 12.5 mL (12 M)

Take 12.5 mL of 12 M solution and add H_2O to a final volume of 500 mL. Result is 0.30 M HCl.

We have 500. mL of 0.24 M NaOH. This solution is diluted to 800. mL. What is the molarity of the resulting solution? (0.15 M)

 $C_2 = (500 \text{ mL})(0.24 \text{ M}) = 0.15 \text{ M}$ (800 mL)

 Given a 15% solution of NaCl, how would you make 250 mL of a 0.90% solution of NaCl? Show the calculation and describe the process. Assume solutions are % by mass/ volume.

 $V_1 = (250 \text{ mL})(0.90\%)$ = 15 mL

Take 15 mL of 15% NaCl and add H_2O to a final volume of 250 mL. The result will be 0.9% NaCl solution.

5. What is the final volume of 125 mL of a 4.00 M solution diluted to 2.50 M? $V_2 = (125 \text{ mL})(4.00 \text{ M})$ (2.50 M)

Review: Solution Concentrations

1. A typical intravenous solution used at the hospital consists of 5.0% (m/v) dextrose (also called glucose). How many grams of dextrose is in 500. mL? (25 g)

```
500. mL \left(\frac{5.0 \text{ g}}{100. \text{ mL}}\right) = 25 \text{ g}
```

2. Whole milk (homogenized) contains 3.6% (*m/m*) butterfat. How many grams of fat is in a cold glass of milk (8.00 fluid oz.)? The density of whole milk at 4 °C is 1.033 g/mL and 1 mL = 0.034 fl. oz. (8.7 g)

8.00 fl. oz. $\left(\frac{1 \text{ mL}}{0.034 \text{ fl. oz.}}\right) \left(\frac{1.033 \text{ g}}{1 \text{ mL}}\right) \left(\frac{3.6 \text{ g fat}}{100 \text{ g milk}}\right) = 8.8 \text{ g fat}$

3. a. If you make a solution by using 45 g sucrose and add enough water to make 250. mL of solution, what is the % (m/v) of this solution? (18%)

```
\frac{45 \text{ g}}{250 \text{ mL}} \times 100\% = 18\% (m/v)
```

- b. What is the molarity of this solution? Sucrose is $C_{12}H_{22}O_{11}$. (0.53 M) $45 \text{ g } C_{12}H_{22}O_{11} \left(\begin{array}{c} 1 \text{ mol} \\ \hline 342.34 \text{ g} \end{array} \right) = 0.131 \text{ mol} \quad \begin{array}{c} 0.131 \text{ mol} \\ \hline 0.250 \text{ L} \end{array} = 0.53 \text{ M}$
- 4. A bottle of cold medicine contains 25% (v/v) ethyl alcohol (ethanol) in water. In 125 mL of medicine, how much ethyl alcohol is there (in mL)? (31 mL)

```
125 mL medicine ( 25 mL ethyl alcohol )
                   100 mL medicine
```

```
= 31.25 mL \Rightarrow (31 mL ethyl alcohol (2 s.f.))
```

1 mol

(= 1.28 g

5. What is the molarity of a solution when 0.135 moles of LiOH are dissolved in water to give 50.0 mL of solution. (2.70 M)

```
M = \left(\frac{0.135 \text{ mol}}{0.0500 \text{ L}}\right) = 2.70 \text{ M LiOH}
```

6. How many grams of NaOH are there in 80.0 mL of 0.400M NaOH solution? (1.28 g)

0.400 mol	- 0.022 mol
1 L	= 0.032 1101

0.0800 L

0.032 mol (

- 7. How many milliliters of 0.250M NaOH are needed to provide 0.200 moles of NaOH? (800. mL)

```
0.200 mol
            1000 \text{ mL} = 800. mL
```

8. We have 500. mL of 0.24 M NaOH. This solution is diluted to a final volume of 800 mL. What is the molarity of the resulting solution? (0.15 M)

```
(500 mL)(0.24 M)
                  (= 0.15 M)
    (800 mL)
```

9. Given a 15% solution of NaCl, how would you make 250 mL of a 0.90% solution of NaCl? Show the calculation and describe the process. (Assume solutions are % by mass/ volume.)

```
(15\%) V_1 = (250 \text{ mL})(0.90\%)
```

```
V_1 = 15 \text{ mL}
Take 15 mL of 15% (m/v) solution and add H<sub>2</sub>O to it until the final volume is 250 mL. The concentration
of this solution will be 0.90\% (m/v) NaCl.
                                                                                                                Unit 11
```

Osmosis

Osmosis is a process by which molecules of a solvent tend to pass through a semipermeable membrane from a less concentrated solution into a more concentrated one.

- 1. In the process of osmosis, water moves through a semi-permeable membrane from a solution of lower / higher solute concentration to a solution with a lower / higher solute concentration.
- 2. In biology, students are often taught that water moves through a semi-permeable membrane from a higher water concentration to a lower water concentration. Is this true or false?
- 3. a. In each case below, predict whether the egg will expand, shrink, or stay the same volume based on its contents and the solution it is in.

Egg expands

Inside: 5% (m/v) Sucrose

Outside: Distilled Water

Egg shrinks

Higher conc. outside

Inside: 0.8% (m/v) Saline Outside: Syrup (saturated sugar)

b. Label each case above as containing a solution that is "hypertonic," "hypotonic," or "isotonic" to the egg.

Hypotonic

Hypertonic

Hypertonic

- 4. Check the correct box according to whether the solution described is iso/hypo/hypertonic: Isotonic Hypotonic Hypertonic
 - a. The concentration of the solute in this solution is lower than the concentration inside the cell. Х
 - b. When a cell is placed in this solution, water will enter the cell by osmosis causing it to swell. Х
 - c. When this solution is injected into the body, no cell disruption occurs because no net osmosis occurs. Х
 - d. Putting a plant in this solution will result in water loss and cause the plant to wilt.
 - e. When a cell is placed in this type of solution, there will be equal amounts of water moving in and out of the cell at equal rates. Χ

Higher conc. inside lower conc. outside)



Higher conc. outside

X

Inside: Distilled Water Outside: 5% (m/v) NaCl

Egg shrinks



5. If you drink lots of seawater, it can cause dehydration. Use the concept of osmosis to explain why that is.

Seawater		Seawater is hypertonic to our cells.
(hypertonic)		Water moves out of the cells—cells shrink.
	Cell	Dehydration is the loss of water.

6. Dialysis is used to treat patients with kidney failure. Do a search (in the book or on the internet) to read more about how dialysis works. Explain how small waste particles can be removed from blood without losing blood cells, which get returned to the patient.

Cells are too large to pass through the membrane. Waste products move to lower solute concentration. CHEM& 121 WORKBOOK

UNIT 12 CHEMICAL EQUILIBRIUM



Notes: Chemical Equilibrium



Many reactions are reversible.

Reversible reactions operate in both "directions."

Equilibrium is reached when:



WRITING AN EQUILIBRIUM CONSTANT EXPRESSION FOR A REACTION

Reversible reactions reach equilibrium no matter what the initial concentrations of reactants or products were. When mathematically manipulated in a certain way, the values of the concentrations of reactants and products at equilibrium yield a constant value (a specific number) at the stated temperature. This constant value is called the **equilibrium constant** for the reaction.

General Format	reactants	<u> </u>	nroducts	K – _	[products]
General i officia.	reactants		products	K -	[reactants]
Specific Format:	aA + bB		cC + dD	K =	[C] ^c [D] ^d [A] ^a [B] ^b

[] means concentration in molarity (mol/L)

EXAMPLE

- Nitrogen gas reacts with hydrogen gas to produce ammonia gas.
- Write balanced chemical equation: $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$
- Write equilibrium expression: $K = \frac{[NH_3]^2}{[N_2][H_2]^3}$

MAGNITUDE OF K DESCRIBES "EXTENT" OF REACTION

lf K = 1,		The reaction favors: neither reactants nor products
lf [products] >> [reactants]: K >> I	\rightarrow	The reaction favors: products
lf [reactants] >> [products]: K << I	\rightarrow	The reaction favors: reactants

LE CHATELIER'S PRINCIPLE

A reaction already at equilibrium will react to relieve applied "stress," which will "shift" the reaction in a certain direction. Let's consider this system that is already at equilibrium: $H_2CO_3 + H_2O \rightleftharpoons H_3O^+ + HCO_3^-$. Reaction will shift to the left.

What happens if more H_30^+ is added? What will it react with? What will be produced?

Notes: Acids and Bases



ACIDS AND BASES IN EVERYDAY LIFE

Typical properties of acids: sour tasting, burn skin, dissolve metals, react with bases

Typical properties of bases: bitter tasting, slippery feeling on skin, react with acids

Common acids and their uses: hydrochloric acid (stomach acid), hydrofluoric acid (etches glass), sulfuric acid (battery acid), phosphoric acid (colas), nitric acid (fertilizers), acetic acid (vinegar), citric acid (citrus fruit).

Common bases and their uses: sodium hydroxide (lye), lithium hydroxide (batteries), calcium hydroxide (cement), magnesium hydroxide (antacid), sodium bicarbonate (baking soda), sodium carbonate (baking powder), sodium hypochlorite (bleach), ammonia (cleaning agent)

BRØNSTED-LOWRY ACIDS AND BASES

Covalent molecules involve the sharing of electrons between atoms to make covalent bonds and therefore do not consist of ions, so why do some covalent molecules act as electrolytes?

Acids donate (give away) protons.
 Bases accept protons.

EXAMPLES

 $\begin{array}{l} \textbf{H}_{acid} \textbf{C} \textbf{I} + \textbf{H}_2 \textbf{0} \rightarrow \textbf{C} \textbf{I}^- + \textbf{H}_3 \textbf{0}^+ \\ \textbf{dissociate} \\ \text{Na} \textbf{0} \textbf{H} + \textbf{H}_2 \textbf{0} \\ \textbf{Na}^+ + \textbf{0} \textbf{H}^- + \textbf{H}_2 \textbf{0} \rightarrow \textbf{H}_2 \textbf{0} + \textbf{0} \textbf{H}^- \\ \textbf{base} \end{array} \qquad \qquad \begin{array}{l} \textbf{H}_{acid} \textbf{F} + \textbf{H}_2 \textbf{0} \rightarrow \textbf{F}^- + \textbf{H}_3 \textbf{0}^+ \\ \textbf{acid} \textbf{H}_3 + \textbf{H}_2 \textbf{0} \rightarrow \textbf{N} \textbf{H}_4^+ + \textbf{0} \textbf{H}^- \\ \textbf{h}_{acid} \textbf{H}_3 + \textbf{H}_2 \textbf{0} \rightarrow \textbf{N} \textbf{H}_4^+ + \textbf{0} \textbf{H}^- \end{array}$

Double-headed arrows indicate that both the forward and reverse reaction take place.



An acid loses a proton and becomes the conjugate base (CB). A base gains a proton and becomes the conjugate acid (CA).

The acid and base are located on the left-hand side of the reaction arrow, and the conjugate acid and conjugate base are located on the right-hand side of the reaction arrow.



REACTIONS OF ACIDS AND BASES WITH WATER: INTRO TO STRONG VS. WEAK

ACIDS

Covalent compounds that donate H^{+} ions to water molecules.

Strong Acids
 Example:

nitric acid



There are six strong acids: hydrochloric acid (HCl), hydrobromic acid (HBr), hydroiodic acid (HI), nitric acid (HNO₃), sulfuric acid (H₂SO₄), and perchloric acid (HClO₄). You need to be familiar with these and recognize them as strong acids. The list will be provided on the exams (do not need to memorize). Example: acetic acid (ethanoic acid, in vinegar)

Weak Acids



Any acid (substance that donates H^+) that is not one of the 6 strong acids is a weak acid!

BASES

Compounds that act as acids accept H⁺ ions from water molecules.

Strong Bases

Example: KOH

к⁺[:ё—н]

These are actually not covalent compounds, but instead are ionic compounds that contain the hydroxide ion, OH⁻, along with a metal cation.



Covalent compounds that act as bases by accepting $H^{\rm +}$ ions from water molecules.

Amines and amides are organic functional groups that contain a lone pair on a nitrogen atom that can bond to H⁺ ions.

MONOPROTIC VS. DIPROTIC ACIDS



Instructions for this section:

- Draw in the correct type of reaction arrow in each chemical equation.
- Draw in the particles that are present in the solution in each beaker.

Strong acids and strong bases completely ionize or dissociate in aqueous solution.



Weak acids and weak bases only partially ionize or dissociate in aqueous solution.



If you have already covered the concept of equilibrium, which of these would have large equilibrium constants? Small *K*? **Strong acids/bases in H**₂**O (large K)**. **Weak acids/bases (small K)**.

Brønsted-Lowry Acids and Bases

- 1. A Brønsted-Lowry **acid** donates H⁺ and a Brønsted-Lowry **base** accepts H⁺. (remove H⁺ from acid)
- 2. Give the conjugate base for each of the following Brønsted-Lowry acids:
 - a. HI I⁻ b. H_2SO_4 HSO₄⁻ c. HCO_3^- CO₃²⁻ d. H_2O OH⁻ (add H⁺ to conjugate base)
- 3. Give the conjugate acid for each of the following Brønsted-Lowry bases:
 - a. $H_2PO_4^- H_3PO_4$ b. $NH_3 NH_4^+$ c. $OH^- H_2O$ d. $H_2O H_3O^+$
- 4. Complete the following table:

Acid	Conjugate Base
CH ₃ COOH	CH₃COO ⁻
HCIO ₄	CIO ₄ ⁻
HNO ₃	NO ₃ ⁻
HCN	CN [−]

5. Complete the following table:

Base	Conjugate Acid
HCO ₃ ⁻	H ₂ CO ₃
F	HF
C0 ₃ ²⁻	HCO ₃ ⁻
H ₂ S	HS⁻

- 6. For the following reaction: $NH_3 + H_2^0 \rightleftharpoons NH_4^+ + 0H^$
 - a. Identify which reactant is the acid and which is the base.
 - b. Identify the conjugate acid and conjugate base.
 - c. Identify the conjugate acid-base pairs using connecting lines.
- 7. For the following reaction: $H_2^{\text{base}} + H_2^{\text{conj. acid}} = H_3^{\text{conj. base}} + H_3^{\text{conj. base}}$
 - a. Identify which reactant is the acid and which is the base.
 - b. Identify the conjugate acid and conjugate base.
 - c. Identify the conjugate acid-base pairs using connecting lines.
- Since H₂O can act as either a weak acid or a weak base, it can react with itself. Write a balanced chemical equation for the reaction. The term for something that behaves as a weak acid or weak base is *"amphoteric."*

conj. acid conj. base $H_20 + H_20 \rightarrow H_30^+ + 0H^-$

Notes: *K*_w and the pH Scale



Is water an acid, a base, either, or neither? Look back over the reactions presented so far and see what water was acting as.

Water is **amphoteric**, which means that water can act as either an acid or a base.

Write a chemical equation for the reaction of two water molecules, where one acts as the acid and the other as the base.

HINTS

- Water is not one of the 6 strong acids. Which type of reaction arrow is best to use?
- Perform a proton transfer. Move H⁺ from one water to the other. Figure out the new chemical formulas after gain or loss of H⁺. (When adding H⁺, add an H and increase charge by 1. When losing H⁺, remove an H and decrease charge by 1.)

$$H_2O + H_2O \rightleftharpoons H_3O^+ + OH^-$$

IONIZATION CONSTANT FOR WATER, Kw

The concentration hydronium ion and hydroxide ion in water are controlled by the equilibrium of the auto-ionization of water. Let's write an equilibrium expression for the auto-ionization of water.

Pure substances are not included in equilibrium expressions, so leave [H₂0] out!

$$K_{\rm w} = \frac{[{\rm H}_3 0^+][0{\rm H}^-]}{[{\rm H}_2 0]^2 \,({\rm leave out})} = 1.0 \times 10^{-14}$$

The equilibrium constant for water at room temperature is 1.0×10^{-14} .

The mathematical product of the concentration of hydronium ion times the concentration of hydroxide ion will always equal 1.0×10^{-14} .

If you know the concentration of H_3O^+ , you can solve for OH^- , and vice versa.

EXAMPLE

A solution has a hydroxide concentration of 7.0×10^{-11} M. What is the concentration of hydronium ion? $\frac{[H_30^+][0H^-]}{[0H^-]} = \frac{1.0 \times 10^{-4}}{[0H^-]} = \frac{1.0 \times 10^{-14}}{7.0 \times 10^{-11}} = 1.4 \times 10^{-4}$

NEUTRAL, ACIDIC, OR BASIC SOLUTIONS

When $[H_30^+] = [OH^-]$, the solution is **neutral**.

When $[H_30^+] > [OH^-]$, the solution is **acidic**.

When $[H_30^+] < [OH^-]$, the solution is **basic**.



Would a solution with a hydroxide concentration of 7.0×10^{-11} be acidic, basic, or neutral? $[OH^{-}] = 7.0 \times 10^{-11}$ M $[H_3O^+] = 1.4 \times 10^{-4}$ M \leftarrow greater than 1×10^{-7} , so acidic.

THE pH SCALE

Consider a solution of 0.00100 M HCl; this concentration can also be written as 0.00100 mol/L HCl. Because HCl is a strong acid and it completely ionizes in solution, there are 0.00100 mol/L H_30^+ or 1×10^{-3} mol/L H_30^+ . Hey, wait chemists always have shorthand notations for situations that repeatedly deal with exponents; what's the shorthand here? The pH scale.

The pH scale is used to express the acidity, or the related basicity, of solutions. The pH of a solution is a measure of the amount of hydronium ions (or protons) present, or the related amount of hydroxide ions present.

The "p" means *power* and is the mathematical operation of a negative logarithm, and the "H" refers to hydronium ions. A logarithmic scale counts each change in a factor of 10 as a whole number integer. The pH values for solutions are typically between 0 to 14, nice whole numbers to work with rather than powers of ten.

$$pH = -\log [H_3O^+]$$



Your calculator should have a "log" (logarithm base 10) button. Make sure to use the negative sign, not the subtraction button.

Taking the antilog of this equation lets you solve for hydronium ion concentration given a pH.

$$[H_30^+] = 10^{-pH}$$

On your calculator, enter 10 raised to the negative of the pH. Type in "10", then "^", then "-", then the pH value!

EXAMPLES

$[H_30^+] = 1 \times 10^{-6} M$	pH = 8
pH = 6	[H ₃ 0 ⁺] = 10 ^{−pH} = 1 × 10 ^{−8} M
pH = –log (1 × 10 ⁻⁶)	
$[H_30^+] = 7.5 \times 10^{-6} M$	pH = 8.75
pH = 5.12	[H ₃ 0 ⁺] = 10 ^{-8.75} = 1.78 × 10 ⁻⁹ M
–log (7.5 × 10 ^{–6})	

pH Concept



1. Write a balanced chemical equation for self-ionization of water. Write the equilibrium expression for this reaction and include its value (K_w). $H_2O(I) + H_2O \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

 $n_{\rm W} = [11_30] [[011]] = 1.0 \times 10$

- What is pH? What is the range of pH values? pH is a measure of acidity (based on a logarithmic scale of H⁺ concentration). The range of pH values are 0–14.
- 3. What is the pH of
 - a. a neutral solution? 7 c. a basic solution? > 7
 - b. an acidic solution? < 7
- 4. a. What is the pH of a solution with $[H^+] = 1.0 \times 10^{-10}$ M? pH = 10
 - b. What is the hydroxide concentration in this solution? Is it acidic or basic? pH > 7
 - c. What is the pH of a solution with $[H^+] = 1.0 \times 10^{-3}$ M? Is it acidic or basic? pH < 7 pH = 3
- 5. Use the words high or low to fill in the following blanks.
 Acidic solutions have high H⁺ ion concentration and low pH values.
 Basic solutions have low H⁺ ion concentration and high pH values.
 Acidic solutions have high H⁺ ion concentration and low OH⁻ ion concentration.
 Basic solutions have low H⁺ ion concentration and high OH⁻ ion concentration.

pH Calculations





1. An aqueous solution has a hydronium ion concentration of 1×10^{-9} M. Is the solution acidic, basic, or neutral?

 $[H_30^+] = 1 \times 10^{-9} M$ (pH = 9)

2. An aqueous solution has a hydroxide ion concentration of 4.8×10^{-3} M. What is the concentration of hydronium ion? $[OH^-] = 4.8 \times 10^{-3}$ M

$$H_30^+] = \frac{1.0 \times 10^{-14}}{4.8 \times 10^{-3}} = 2.08 \times 10^{-12} \text{ M}$$

- 3. An decrease of 2 pH units is an increase of hydronium ion concentration by a factor of 100. Two powers of $10 = 10^2 = 100$ times.
- 4. A solution has a pH of 3.9. Is the solution acidic, basic, or neutral?

 $[H_{3}0^{+}]$

- 5. A solution has a hydronium concentration of 7.5×10^{-4} . What is the pH? pH = -log [H₃0⁺] = -log (7.5 × 10⁻⁴) = (3.12)
- 6. A solution has a pH of 5.5. What is the concentration of hydronium ion? $[H_30^+] = 10^{-5.5} = 3.16 \times 10^{-6} M$
- 7. A solution has a hydroxide ion concentration of 2.2×10^{-2} M. What is the pH? $[H_30^+] = \frac{1.0 \times 10^{-14}}{2.2 \times 10^{-2}} = 4.5 \times 10^{-13} \text{ M} \qquad \text{pH} = -\log(4.5 \times 10^{-13}) = 12.3$
- 8. Fill in the missing information. For example, $[H_3O^+] = 1.0 \times 10^{-6}$ needs to be related to the pH, the concentration of hydroxide ion $[OH^-]$, the pOH, as well as specifying whether the solution is acidic or basic.

рН	[H ₃ 0 ⁺]	[OH ⁻]	Acidic, Basic, or Neutral?
6	1.0×10^{-6}	1.0 × 10 ⁻⁸	acidic
10	10 ⁻¹⁰	10 ⁻⁴	basic
12.6	2.70 × 10⁻¹³	3.7×10^{-2}	basic
8.59	$10^{-8.59} = 2.6 \times 10^{-9}$	3.89 × 10⁻⁶	basic
3.60	2.5×10^{-4}	4 × 10 ⁻¹¹	acidic
7.00	1.0 × 10 ⁻⁷	1.0 × 10 ⁻⁷	neutral

From: The Open Course Library Project, http://www.openwa.org/open-course-library/

Notes: Buffer Solutions



Buffer

HΑ

Δ

Will neutralize bases added to buffer.

HA

HA

Conj. base of 1

Weak acid

- Add H_30^+

Will neutralize

acids added to

the buffer.

or OH

WHAT IS A BUFFER SOLUTION (SYSTEM)?

A buffer solution resists changes in pH by neutralizing small amounts of added acid or base. A buffer can do this because it contains both weak acids, which can react with added bases, and weak bases, which can react with added acids.

WHAT ARE THE COMPONENTS OF A BUFFER SYSTEM?

- Ionic compound a weak acid and its conjugate base (in the form of a "salt") Ex. H₂CO₃/HCO₃⁻ (or KHCO₃/NaHCO₃) ... or ... 🔼
- a weak base and its conjugate acid (in the form of a "salt") Ex. NH₃/NH₄⁺ NH₄CI

HOW DOES A BUFFER WORK?

- The weak acid in the buffer can react with base $(OH^{-} ions)$ added to the solution. **HA** (in buffer) + OH⁻ (added to buffer) \rightarrow H₂O (water is formed) + *I*
- The conjugate base in the buffer can react with any acid (H_30^+) added to the solution. **A**⁻ (in buffer) + H₃0⁺ (added to buffer) \rightarrow HA + H₂0 (water is formed)

PRACTICE QUESTION

Which of these pairs can act as buffer systems?

a. HCI and NaOH **b.** HF and KF c. KCI and NaCI Strong Weak acid and conj. base No acid/base

d. NaHCO₃ and H₂CO₃ Weak acid and conj. base

BUFFER AMOUNTS

TYPICAL CONCENTRATIONS

The concentrations of acid and conjugate base (or base and conjugate acid) are typically equal (or similar) so that the buffer system can neutralize added acid or base equally well.

BUFFER CAPACITY

The higher the concentration of the buffer components, the more added acid or base the buffer can neutralize. As the buffer concentrations increase, the buffer capacity increases.

EXAMPLE

Consider 1 L of 0.1 M CH₃COOH/0.1 M CH₃COO $^{\ominus}$ (write in the other buffer component). How many moles of NaOH can the buffer neutralize before it is exhausted?

= 0.1 mol can be neutralized

NEUTRALIZATION REACTIONS IN BUFFERS

If small amounts of acid or base are added to a buffer system (much less than the buffer capacity), then the relative concentrations of the buffer components will change after neutralization occurs. When one buffer component reacts to neutralize added acid or base, it is consumed and the other buffer component is generated as a product.

- When acid (H₃0⁺) is added to a buffer system, which buffer component will neutralize it? The weak acid or the conjugate base?
- When base (OH⁻) is added to a buffer system, which buffer component will neutralize it? The weak acid or the conjugate base?



Let's practice writing chemical equations to represent the neutralization reactions that take place in a buffer.

EXAMPLE

Consider the buffer system of $HNO_2/NaNO_2$. Write equations for the action of the buffer when the following are added to the buffer.

acid is added:	H ₃ O ⁺ (added)	+	NO ₂ ⁻ (in buffer)	\rightarrow HNO ₂ + H ₂ O
base is added:	OH ⁻ (added)	+	HNO ₂ ⁻ (in buffer)	\rightarrow NO ₂ ⁻ + H ₂ O More coni, base is formed.

Neutralization Reactions and Buffers



The products of a neutralization reaction include water and a salt (ionic compound).

Sometimes if the base is a bicarbonate (HCO_3^{-1}) or carbonate (CO_3^{2-}), CO_2 will also be produced as a decomposition product of carbonic acid H_2CO_3 .

- 1. The neutralization reaction between $AI(OH)_3$ and HNO_3 produces the salt with the formula $AI(NO_3)_3$.
- 2. Complete and balance the following reactions:
 - a. $HNO_3 + NaOH \rightarrow H_2O + NaNO_3$
 - b. **2** HCl + Ca(OH)₂ \rightarrow **2** H₂O + CaCl₂
 - c. $H_2SO_4 + 2 KOH \rightarrow K_2SO_4 + 2 H_2O$
 - d. $KHCO_3 + HCI \rightarrow KCI + H_2O + CO_2$

(NOTE: Bicarbonate is involved!)

- e. CH_3COOH (also known as $HC_2H_3O_2$) + NaOH $\rightarrow CH_3COONa$ (or $NaC_2H_3O_2$) + H_2O
- 3. A buffer can be made by dissolving H₃PO₄ and NaH₂PO₄ in water. A buffer solution is one that resists changes to pH due to its ability to neutralize an acid or base being added to it. It is often composed of a weak acid and its conjugate base, or a weak base and its conjugate acid.
 - a. Write the equation that shows how the buffer would neutralize a small amount of added HCI (remember HCI + $H_2O \rightarrow H_3O^+ + CI^-$ so use H_3O^+ in the equation):

 $\begin{array}{c} \text{H}_2\text{PO}_4^- + \text{HCI} \rightarrow \text{H}_3\text{PO}_4 + \text{CI}^- \\ \text{weak acid} \end{array} \end{array}$

b. Write the equation that shows how the buffer would neutralize a small amount of added NaOH (remember NaOH \rightarrow Na⁺ + OH⁻ so use OH⁻ in the equation):

 $H_3PO_4 + OH^- \rightarrow H_2O + H_2PO_4^$ weak base

Biological Buffers



Normal blood pH must be maintained within a narrow range of 7.35–7.45 to ensure the proper functioning of metabolic processes and the delivery of the right amount of oxygen to tissues. The blood contains a carbonic acid/bicarbonate (H_2CO_3/HCO_3^-) buffer system as one of several buffers in the body used to maintain blood pH within normal values.

1. Write an equation for how the blood buffers a strong acid, such as HCI (use H30⁺ as a reactant):

 $H_30^+ + HCO_3^- \longrightarrow H_2CO_3 + H_2O$

2. Write an equation for how the blood buffers a strong base, such as NaOH (use OH⁻ as a reactant):

 $OH^- + H_2CO_3 \longrightarrow HCO_3^- + H_2O$

The amount of carbonic acid in blood (and thus, blood pH) can be affected by levels of dissolved CO₂:

Equation 1 $H_2O + CO_2 \rightleftharpoons H_2CO_3$ Equation 2 $H_2CO_3 \rightleftharpoons HCO_3^- + H^+$

According to Equation 1, what happens to the levels of H₂CO₃ if MORE CO₂ were present in blood? In other words, does adding more CO₂ increase or decrease the amount of H₂CO₃?

 CO_2 (reactant) is present, the more H_2CO_3 can be formed (via Equation 1).

4. What happens to blood pH if this happens: does it increase or decrease? (HINT: How does [H⁺] change with H₂CO₃, see Equation 2?) Does the blood become more acidic or less acidic?

If H_2CO_3 increases, more H^+ can be formed (via Equation 2).

Increase in H⁺ makes the blood more acidic.

5. If blood becomes too acidic, do you want to increase or decrease levels of CO₂?

Decrease CO_2 to make less H_2CO_3 . Then less H^+ will be formed.

Respiratory **acidosis** (when blood pH is too low) or **alkalosis** (when blood pH is too high) can be caused by the following (see table below). Notice that any disease or condition that affects the lungs, kidneys, metabolism, or breathing has the potential to cause acidosis or alkalosis.

Respiratory Acidosis	Respiratory Alkalosis	
Decreased breathing rate (respiratory drive) due to drugs or central nervous system disorders	Hyperventilation due to anxiety, pain, shock	
Impaired breathing and lung movement (respiratory mechanics) due, for example, to trauma or abnormal presence of air between the lung and the wall of the chest (pneumothorax)	Drugs (aspirin)	
Respiratory muscle/nerve disease (myasthenia gravis, botulism, amyotrophic lateral sclerosis (ALS), Guillain-Barre syndrome)	Pneumonia, pulmonary (lung) congestion, or embolism	
Airway obstruction (food or foreign object)	Exercise, fever	
Lung disease	Central nervous system tumor, trauma, infection (meningitis, encephalitis)	

Taken from http://labtestsonline.org/understanding/conditions/acidosis/start/1

Metabolism also generates large quantities of acids that must be neutralized and/or eliminated to maintain pH balance. Most of the acid is carbonic acid, which is created from carbon dioxide (CO_2) and water. Lesser quantities of lactic acid, ketoacids, and other organic acids are also produced.

6. A person who is stressed out or under trauma can hyperventilate which results in rapid breathing and loss of CO₂. Use a chemical equation and words to explain why the blood pH increases during hyperventilation. If blood pH is increased, is it acidosis or alkalosis?

In hyperventilation, CO_2 is lost. This means less H_2CO_3 will be produced, which means less H⁺ than needed. This result in alkaline blood (basic, high pH).

7. One way to treat alkalosis as a result of hyperventilation is to exhale into a paper bag, and then breathe CO₂ back into the lungs. Does this increase or decrease the blood pH? Explain how it does that.

Increase in CO_2 increases production of H_2CO_3 , which dissociates to CO_2 . This helps reduce blood pH (makes it more acidic).

During starvation, the body makes acidic compounds which can cause metabolic acidosis. Will the blood pH increase or decrease as a result of acidosis? If you could, would you resist this by increasing CO₂ levels or decreasing CO₂ levels?
 Acidic compounds will make blood more acidic (decrease the blood pH).

To raise the blood pH, decrease [H⁺] by decreasing H_2CO_3 production. Decrease H_2CO_3 production by decreasing CO_2 levels in the blood.

pH and Buffers Review Worksheet



1. Since H_2O can act as either a weak acid or a weak base, it can react with itself. Write a balanced chemical equation for the reaction.

 $\begin{array}{c} \text{H}_2 0 + \text{H}_2 0 \\ \text{acid} \\ \text{base} \\ \end{array} \rightarrow \begin{array}{c} \text{H}_3 0^+ + 0 \text{H}^- \\ \text{conj. acid conj. base} \end{array}$

- 2. What is the pH of a solution that has a H_3O^+ concentration of 3.4×10^{-5} M? pH = -log (3.4×10^{-5} M) pH = 4.47
- 3. What is the $[H_3O^+]$ in a solution that has a pH of 3.456? $[H_3O^+] = 10^{-pH} = 10^{-3.45} = 3.55 \times 10^{-4}$

or [H₃0⁺]

4. What is the H⁺ concentration of a solution that has a $[OH]^-$ of 4.87 × 10⁻⁸?

$$\frac{[\mathrm{H}^+]}{4.87 \times 10^{-8}} = 2.1 \times 10^{-7} \mathrm{M}$$

5. What is the [OH⁻] of a solution that has a pH of 3? $[H_{3}O^{+}] = 10^{-3}$ $[OH^{-}] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{-11} \text{ M}$ 6. What is the pH of a solution that has an OH⁻ concentration of 8.2×10^{-9} M? [OH⁻] = 8.2×10^{-9} M [H⁺] = $\frac{1.0 \times 10^{-14}}{8.2 \times 10^{-9}}$ = 1.2×10^{-6} M

$$pH = -log (1.2 \times 10^{-6})$$

 $pH = 5.92$

7. Indicate whether each of the following solutions is acidic, basic, or neutral:

a.
$$pH = 7$$
 Neutral c. $[H_3O^+] = 1.0 \times 10^{-8}$ M pH = 8 (basic)

b.
$$pH = 9$$
 Basic
d. $[OH^-] = 10^{-3}$ $[H_3O^+] = 10^{-11}$
 $pH = 11$ (basic)

- 8. Which of the following pairs constitutes a buffer solution?
 - a. HCl, NaCl c. NaOH, NH₃ HCl – strong acid NaOH – strong base
 - b. HF, H₂SO₄
 HF weak acid
 H₂SO₄ but not conj. base of HF

d HC₂H₃O₂, NaC₂H₃O₂ HC₂H₃O₂ - weak acid + its conj. base

- 9. A buffer can be made by dissolving acetic acid/acetate ($HC_2H_3O_2$ and $C_2H_3O_2^-$) in water.
 - a. Write the equation that shows how the buffer would neutralize a small amount of added H_3O+ (from a strong acid):

```
Yes, HC_2H_3O_2 is a weak acid and C_2H_3O_2 is the conj. base.
H_3O^+ (added to buffer) + C_2H_3O_2^- (buffer) \rightarrow HC_2H_3O_2 + H_2O_2^-
```

b. Write the equation that shows how the buffer would neutralize a small amount of added OH⁻ (from a strong base):

 OH^- (added to buffer) + $HC_2H_3O_2 \rightarrow H_2O + C_2H_3O_2^-$

Laboratory Equipment: Names and Pictures



Pasteur Pipet and Bulb

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Lab Safety Quiz (10 Points)



Name: ______

Section: _____

- 1. Indicate location of exits, eye-wash, hoods, safety shower, fire extinguisher, and firstaid box. (You can draw a box and indicate positions or list them and make a comment on the position.)
- 2. What is the correct procedure for combining acid and water?
- 3. When in laboratory, you should wear_____
- 4. What happens if you heat a closed test-tube?
- 5. In case of fire in the laboratory, notify the teacher at once and then prepare to

6. If you are not sure what to do in an emergency then ______.

- 7. Why is eating and smoking not permitted in the laboratory?
- 8. What is the correct technique for checking the odor of a substance?
- 9. Draw an Erlenmeyer flask, beaker, and a graduated cylinder.

CHEM& 121WORKBOOK

EXPERIMENT 1 MEASUREMENTS



Measurements are essential to experimental sciences such as chemistry, physics, biology, and geology. The measurements are usually made using the metric system units. This experiment is intended to give practice in making measurements to the greatest precision possible using metric 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.

PROCEDURE

MASS

Your instructor will give directions on how to use the balance. Some general rules to follow when using the balances follow:

- The "tare" button will zero the balance.
- 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 on the balance.

- 1. In order to gain practice with the balance, measure the mass of one of the weights in the weight sets. (Choose any size weight you like.)
- 2. Weigh a coin to the nearest 0.01 gram. Record your answer on the data page.
- 3. Weigh an empty 150-mL beaker to the nearest 0.01 grams. Record your measurement.

VOLUME MEASUREMENTS

There are 10-mL, 25-mL, and 100-mL graduated cylinders available. Use the size most appropriate for the measurements.

- 1. Fill a test tube to the **brim** (top edge) with water. Measure the volume of water.
- Fill a 125-mL Erlenmeyer flask to the brim with water. Measure the volume of water to the nearest 1-mL.



TEMPERATURE

- Temperature measurements are made using mercury thermometers, thermocouples, gas filled thermometers, alcohol thermometers, etc. We will use a digital 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:
 - Position the thermometer probe away from the walls of the container. Be sure the liquid is thoroughly mixed.
 - Allow the thermometer to be in contact with the liquid for enough time so that the thermometer reaches equilibrium with 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.

COLD TAP WATER

- 1. Half fill a 400-mL beaker with cold tap water.
- 2. Place the thermometer bulb in the water.
- 3. Allow the thermometer to reach thermal equilibrium.
- 4. Keep the thermometer in the middle of the liquid away from the glass.
- 5. Read and record the temperature of the tap water.
BOILING WATER

- 1. Half fill a 250-mL beaker with tap water.
- 2. Add one or two boiling chips to the water, using the arrangement shown in the figure.
- 3. Heat the water to boiling.
- 4. Read and record the temperature being sure to keep the thermometer bulb away from the bottom of the beaker.

ICE WATER

- 1. Place a hand-full of crushed ice in a 250-mL beaker.
- 2. Add approximately 50 to 60 mL of cold tap water.
- 3. Add more ice if necessary so that there is ice mixed with the water.
- Without stirring place the thermometer in the water. Wait for thermal equilibrium, then read and record the temperature. Repeat this measurement with stirring.

ICE WATER PLUS SALT

- 1. Weigh out approximately 4 to 6 grams of sodium chloride, NaCl.
- 2. Add the salt to the ice-water mixture.
- 3. Stir for few minutes and add more ice if necessary.
- 4. Read and record the temperature of the mixture.

DISTANCE/LENGTH MEASUREMENTS

- 1. Measure the external height of a 400-mL beaker in both centimeters and inches.
- 2. Measure the length of a test tube in both centimeters and inches.

DENSITY DETERMINATION

The density of a substance is an intensive physical property. It can be used to help identify a material. The formula for calculating density follows:

Density $(d) = \frac{\text{mass in grams}}{\text{volume in mL}}$



DENSITY OF WATER

- 1. Weigh a clean, dry 25-mL graduated cylinder to the nearest 0.01 gram. Record the weight.
- 2. Add tap water to the graduated cylinder bringing the water level to the 10-mL mark. Record the volume of water.
- 3. Weigh the graduated cylinder and water.
- 4. Calculate the density from the mass of water (obtained from the difference in mass of the empty and full graduated cylinder) and the measured volume.

DENSITY OF A SOLID OBJECT

- Obtain a solid object. Record the vial number. (The unknown vial and the inside of the vial must be dry before you begin the experiment. If it is not, pour the metal onto a paper towel and allow it to dry. Use a paper towel to dry the inside of the vial.)
- 2. Weigh the solid object **and** the vial on the balance. Record the mass.
- 3. Add enough tap water to a 25-mL graduated cylinder to bring the water level to the 10-mL mark. Record the volume of water.
- Carefully place the unknown solid object (all of it) into the graduated cylinder. The water level will rise.
- 5. Read and record the new position of the water level.



Graduated cylinder before and after addition of unknown solid.

- 6. Weigh the empty vial and subtract to find the mass of the unknown solid.
- 7. Calculate the density of the solid object.
- 8. Dry the unknown solid and the vial as described in Step 1.

REPORT SHEETS Experiment 1: Measurements



Name:	Section:
Lab Partner:	

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

MASS MEASUREMENT

1.	Mass of the gram weight from the weight sets	 g
2.	Mass of the coin (identify)	 g
3.	Mass of 150-mL beaker	 g

VOLUME MEASUREMENTS

1.	Test tube	 mL
2.	125-mL flask	 mL

TEMPERATURE

1.	Cold tap water	°C
2.	Boiling water	°C
3.	a. Ice water without stirring	°C
	b. Ice water with stirring	°C
4.	Ice water and salt	°C

DISTANCE/LENGTH MEASUREMENTS

1.	Height of 400-mL beaker	cm	in
2.	Length of test tube	cm	in

DENSITY DETERMINATION FOR WATER

1.	Weight of empty 25-mL graduated cylinder	g
2.	Weight of 25-mL graduated cylinder plus water	g
3.	Weight of water	g
4.	Volume of water	mL
5.	Density calculation	g/mL

Show your calculations:

DENSITY DETERMINATION OF SOLID OBJECT

Via	al number	
1.	Mass of unknown solid and vial	g
2.	Mass of empty vial	g
3.	Mass of solid object	g
4.	Initial volume of water in 25-mL cylinder	mL
5.	Volume of water and solid object	mL
6.	Volume of solid object	mL
7.	Density of solid object	g/mL

Show your calculations:

PRE-LAB Experiment 1: Measurements



Name:

Section:

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. Write the full names of each of the following units:
 - a. mL _____
 - b. mg _____
 - c. km _____
 - d. °C

2. Write the abbreviation for each of the following units:

- a. Liter
- b. micrometer
- c. kilogram
- d. milligram
- 3. Complete the following blanks:
 - a. 1 L = _____ dL
 - b. 1 km = _____m
 - c. 1 mg = _____ g
- d. 1 m = _____ cm
- e. 1 mL = _____ L f. 1 cm³ = _____ mL
- 4. If an unknown metal weighs 12.35 g and occupies a volume of 4.57 mL, what is the density of this metal? Show your calculations and include units in your answer.

CHEM& 121 WORKBOOK

EXPERIMENT 2 CALORIC CONTENT OF FOODS



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INTRODUCTION

Food labels contain information about calories. How is this measured? Our bodies "burn" calories, but we can also burn calories (literally) by setting food on fire! It is difficult to measure the energy released directly from burning food, so this energy will be transferred to a can of water (as "heat"). The change in temperature of the water due to the heat transfer will be used to estimate the amount of heat that was released by each food item. Since the food items are different sizes and masses, a comparison will be done by measuring the heat released from each and calculating the number of Calories **per gram** of each food item, instead of total Calories.

OBJECTIVES

In this experiment, you will

- Observe and measure the amount of energy released from different food items.
- Calculate the amount of heat transferred in units of Calories.
- Gain exposure to the concept of experimental error and discuss sources of error.
- Practice making mass and volume measurements.
- Suggest nutritional explanations for experimental results.

HAZARDS

Open flames can be dangerous. Make sure long hair is tied back. Be ready to extinguish a flame that gets out of control. Do not toss partially lit or glowing items in the trash. Wash your hands with soap when you are done with the experiment.

PROCEDURE

- Weigh each food item by taring a weigh paper on the balance. (Use the tare/zero function). Record its precise mass. (Recommendation: Cheese puffs are big ... use 1 g or less.)
- 2. Make a stand for the food item using a paper clip. See Figure 1.
- 3. Weigh an empty graduated cylinder. Record its mass. Add 100 mL of tap water to the cylinder and weigh it again. Record its mass. Enter this into the data table. Place this into a clean aluminum can.
- Place the can on a clay triangle on a ring stand. Place the food item in an empty tuna or cat food can. Place underneath the soda can.
 Bring the level of the can down so it is close to the food item.
- 5. Clamp a thermometer so it is submerged in the water in the can *but not touching the can.*
- Clay Triangle Ring Clamp Peanut on Fire Paper Clip Tuna Can

Thermometer

Figure 1. Set-up for a burning peanut.

- 6. Record the initial temperature of the water.
- 7. Use a flame to catch the food on fire. Matches, a lighter, or candles will be provided.
- 8. Note the highest temperature that is achieved by the energy from the burning food.
- 9. Examine the food item after burning. Did it burn all the way through?
- 10. Repeat for other food items. Feel free to add any other food items you'd like to test.

CALCULATIONS AND HOW TO FILL OUT YOUR DATA TABLE

Please refer to your data table rows.

- 1. Enter the food item you used.
- 2. Enter the weight of your food item.

- 3, 4, 5. Enter the weight of your empty graduated cylinder, cylinder plus water, and calculate the difference.
- 6, 7, 8. Enter the Initial (starting) temperature, final temperature, and calculate the difference.
- 9. This requires a heat calculation. We will use this equation and solve for q:



For the specific heat of water, use 0.001 Cal/g °C.

EXAMPLE

If your mass of water is 100 grams and the temperature change was 8.5 °C, your heat calculation is:

 $q = (100 \text{ g water} \times 0.001 \text{ Cal/g }^{\circ}\text{C} \times 8.5 \,^{\circ}\text{C}) = 0.85 \text{ Calories}$

This means your water absorbed 0.85 calories from your food item.

- 10. Assuming the energy absorbed by water was released from the burning food item, what is the amount of energy that your food item released? It's the same value as the previous column.
- 11. Divide your Cal of food item (column 10) with the mass of your food item (column 1).

EXAMPLE

If the food item burned was 2.5 grams, then the caloric content is 0.85 Cal/2.5 g = 0.34 Cal/g.

REPORT SHEETS Experiment 2: Caloric Content of Foods



Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PREDICTION

I predict the ______ will have the highest number of Calories per gram.

DATA

1. Food Item		
2. Mass of Food Item (g)		
3. Mass of Empty Cylinder (g)		
4. Mass of Cylinder Plus Water (g)		
5. Mass of Water (g)		
6. Initial Temp (°C)		
7. Final Temp (°C)		
8. Change in Temp, ∆ (°C)		
9. Energy Absorbed by Water (Cal) (See Sample Calculation Below)		
10. Energy Released From Food Item (Cal)		
11. Cal per gram of Food Item		
12. % Error		

POST-LAB QUESTIONS

- 1. Based on your experiment, which food item resulted in the most Calories per gram?
- 2. Are you surprised? Can you give a nutritional explanation for why that food item provided the most energy per gram (what is in the food that contains so many calories)?
- Look up the calories and the grams per serving on the food label and calculate Cal/gram *for each food item* (see sample calculation, right).

Write the results of the calculations here, to answer the next question.

Nutrition Facts Serving Size 1¼ cups (57 g)		
Amount Cereal with Per Serving Cereal ½ Cup Milk		
Calories Calories from Fat	230 30	270 30

Sample Calculation for Cal/g from the Food Label:

230 calories 57 grams = 4.0 Cal/g

4. a. Compare this value to your experimental value (obtained by this experiment) by calculating a percent error for each of your food items and enter this in the last column of your data table. The closer you are to theirs, the higher the accuracy, and the lower the % error.

% error = (your value – their value) their value × 100

- b. If you get a negative % error, why is that?
- 5. Your results may not be as accurate as the laboratories that compute this information for food labels. Can you give at least **TWO** reasons why your results may not be as accurate? (Do NOT say that you miscalculated or had errors in your measurements think about the inherent flaws in this experiment.)

PRE-LAB Experiment 2: Caloric Content of Foods

Name: ______ Section: _____

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

NOTE: If it is convenient for you to do so, please bring a 12-fluid-ounce aluminum soda can to lab for the experiment (one is needed for every two students). There will be some extras in case you and your lab partner do not have one.

- 1. Without looking at anything else, describe the term "heat" as you know it from your daily experiences in 1–2 sentences.
- 2. Ok, now you can look! Using a textbook or the internet, give a definition for "heat" and cite the source (give textbook title/author/edition/pg or website):
- In this experiment, you will use various foods: a peanut, a Cheeto, and a mini-marshmallow. Make a prediction for which food item will result in the greatest number of Calories per gram of food item. Explain why you think this food item will have the highest Cal/g (1–2 sentences max).
- 4. One part of the procedure states that you should bring the level of the can down close to the food item. Why is this important?
- 5. A peanut is much smaller than a cheese puff. This means that a cheese puff might give off more energy simply because of its size, rather than its composition. How will you take this into account in your results so size is not a factor?

CHEM& 121 WORKBOOK

EXPERIMENT 3 POPCORN: Water in a Mixture



PURPOSE

Apply the law of conservation of mass to determine the percentage-by-mass water in popcorn.

Percentage-by-Mass: (Sometimes called parts per hundred). The number of grams of a part in 100 grams of the whole.

DISCUSSION

Foods are composed of carbohydrates, fats, and proteins along with small amounts of vitamins, minerals, and variable amounts of water. Popcorn kernels are corn seeds consisting of these compounds encased in a hard shell. The main component of the mixture is the carbohydrate called starch. Starch is in the form of granules that are impregnated with water. This experiment is designed to find the percentage-by-mass of water in popcorn kernels.

Within a kernel of popcorn, a small amount of water is intimately distributed throughout the starch granules. When a kernel is heated, the temperature of the water rises. As the temperature exceeds 100 °C, the water boils and turns to steam. A kernel which does not pop (a "dud") often has a crack or leak in the shell so that the steam escapes without expanding the starch.

To determine the percentage-by-mass of water in popcorn: First, a sample of kernels is weighed. The sample is heated to pop the kernels. The popped corn is weighed. The mass of water driven off can be found by subtracting the mass of the popped corn from the mass of the unpopped corn. The percentage of water is found by dividing the mass of water by the mass of the sample of unpopped kernels and multiplying by 100.

SAFETY ALERT: DO NOT EAT YOUR PRODUCT.

LABORATORY PROCEDURE

- Weigh a 50-mL beaker. Place 20 kernels of unpopped corn in the weighed beaker. Weigh the beaker with the 20 kernels. Pour these kernels into a hot air popper. *Repeat this process two more times with different kernels.* When the class has added all their kernels to the hot air popper, the popcorn will be popped.
- 2. Weigh a 250-mL beaker. Count out 20 kernels of *popped* corn and place into the weighed beaker. Weigh the beaker with the 20 kernels of *popped* corn. *Repeat this process two more times with different popped kernels.*

Calculate the average mass of an unpopped kernel and the average weight of a popped kernel. The average mass of water driven off from the 20 kernels of popped corn can be found by subtraction of these averages. Use the data to calculate the average **percentage-by-mass** of the water in the popcorn by dividing the mass of the water by the mass of the unpopped kernels and multiplying by 100.

REPORT SHEETS Experiment 3: Popcorn: Water in a Mixture

Name:	 Section:
Lab Partner:	

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

	TRIAL #1	TRIAL #2	TRIAL #3
Mass of 50-mL Beaker			
Mass of 50-mL Beaker and 20 Kernels of Unpopped Corn			
Mass of 20 Kernels of Unpopped Corn			
Average Mass of 20 Kernels of Unpopped Corn			
Mass of 250-mL Beaker			
Mass of 250-mL Beaker and 20 Kernels of Popped Corn			
Mass of 20 Kernels of <i>Popped</i> Corn			
Average Mass of 20 Kernels of Popped Corn			
Average Mass of Water in 20 Kernels			
Average Percentage-by-Mass of Water in Popcorn			

POST-LAB QUESTIONS

- 1. a. What is the law of conservation of mass?
 - b. How did you use the law of conservation of mass in this lab? What did it help you accomplish?
- 2. Why does unpopped popcorn pop when heated?
- 3. a. Find the average mass of one kernel of unpopped popcorn.
 - b. Using this answer, calculate the number of kernels in a one pound jar of popcorn.
- 4. Using the data for the grams of water obtained in your experiment, calculate the number of grams of water per kernel of popcorn.
- 5. Determine the number of grams of water that is contained in 1.0 pound of unpopped corn.

CHEM& 121 WORKBOOK

EXPERIMENT 4 PENNIUM

PART A: AVERAGE ATOMIC MASS

Until 1982, US pennies had been made from pure copper. But in 1982, for economic reasons, the composition was changed and pennies were made from zinc, with a thin coating of copper on the exterior to preserve their appearance. As a result, although all pennies are the same size, there is a significant difference in weight between pre-82 pennies and post-82 pennies.

If pennies were a chemical element (let's say "pennium"), it would have two isotopes (pre-82 and post-82). Both would look the same and behave in the same way, but would differ in their weight, just like ordinary isotopes. In this case, both isotopes would be stable. As with all atomic elements, we would need to know its atomic mass, which is an averaging of all the available isotopes.

Your first task is to determine the "atomic mass" of pennium. In the course of this determination, it is important to use good scientific method. For example, weighing a single penny will give an erroneous result since the true atomic mass of pennium is an average of two isotopes. Furthermore, weighing one pre-82 penny and one post-82 penny will also give a false result since it assumes that 50% of all pennies are pre-82, and 50% are post-82, which may not be the case. Talk this over with your lab partner, and devise a method to get an accurate and realistic atomic mass of pennium. As a secondary task, try to determine the relative abundance of each isotope.

PART B: ALCHEMY

Please use the post-82 pennies for this part of the experiment

Select three of your cleanest pennies and put them aside. If you can't find some relatively shiny pennies, you can also clean your pennies prior to the experiment by putting them in a salt/vinegar mixture. Rinse them with water before proceeding with the experiment. One of these pennies will serve as a control. Prepare a cauldron of approximately 3 mL of 1.5 M sodium hydroxide (NaOH). Into this place a sprinkling of zinc powder and two of your pennies. Heat the cauldron over a Bunsen burner as demonstrated by the instructor (turn the Bunsen burner off if the solution starts to boil vigorously). Occasionally turn the pennies over with a pair of tongs. After a couple of minutes, they will turn silver. When the transformation is complete, remove the pennies with a set of tongs, and place them in a beaker of ordinary tap water to wash them.

Take one of the silver pennies and hold it (by its edge) with a set of tongs in a Bunsen burner flame. (*Be careful; if you heat strongly, you may melt your penny!*) The silver penny will gradually turn to gold. When this is complete, dunk the now gold penny in your beaker of ordinary tap water to cool it to room temperature. Of your three pennies, one should now be copper, one silver, and one gold. Dispose of the NaOH in the provided waste jar.

If you want to take your pennies home and you have small children in your family, please enclose all three of them in contact paper and keep in a secure place.

REPORT SHEETS Experiment 4: Pennium



Name:_____

Section:

Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PART A: AVERAGE ATOMIC MASS

- 1. Describe your method:
- 2. Data collected and calculations:
- 3. Atomic mass of pennium: _____
- 4. Relative abundance of pre-82 penny: _____
- 5. Relative abundance of post-82 penny: _____

PART B: ALCHEMY

- 1. Review the terms **chemical** and **physical** changes. Does the first step (copper to zinc) accomplish a chemical change or a physical change? Explain.
- 2. Does the second step (zinc to brass) appear to be a chemical or physical change? Describe what you think might be happening at the atomic level.

CHEM& 121 WORKBOOK

EXPERIMENT 5 CHEMICAL CHANGES



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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 provide evidence that chemical changes have taken place.

Simple examples of evidence of chemical change include the following: 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 any-thing 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.
- Set up water bath in the beginning of the lab period.

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 (CaCl₂) 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 (KNO₃, 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 NaHCO₃ (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 lons

- a. Obtain a test tube and add about 2–3 mL of 5% (*m*/*v*) CuCl₂•2H₂O (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 5: 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.

Depatient	Observations		
Reactions	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 lons:

a. Based on your observation and chemicals used, what substance do you think deposited on the aluminum wire? What observations explain your choice?

PRE-LAB Experiment 5: Chemical Changes **B B B**

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.

CHEM& 121 WORKBOOK

EXPERIMENT 6 SPECTROSCOPY

INTRODUCTION TO LIGHT

Light is a form of energy called electromagnetic radiation. A chart of the electromagnetic spectrum is shown in Figure 1.



Figure 1. Types of 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 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.

Some other units for wavelength include

- a micron (a micrometer, or 10⁻⁶ meters). It's often used in biology since it is the appropriate size of cells and bacteria. You will encounter microns when using most microscopes.
- an Angstrom (Å). One Angstrom is 10⁻¹⁰ meters and is used for describing the diameter of atoms.



Figure 2. Terminology of a wave.

The second variable used to describe waves is called **frequency**, 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).

Shorter \downarrow Wavelength = Higher \uparrow Frequency

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!

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.

Higher \uparrow Frequency = Higher \uparrow Energy

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 × 10¹⁴) 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, **Y**iolet). An interesting bit of trivia—the British use the following mnemonic: **R**ichard **O**f **Y**ork **G**ave **B**attle **I**n **Y**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.) Electrons that are in the lowest energy orbits such as level 1 (n = 1) are in 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 4. Figure 4a shows 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. Figure 4b shows 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.

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 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.

OBJECTIVES

In this experiment you will

- Learn about light and the relationship between energy, wavelength and frequency.
- Observe colors from flame emission of various elements and identify an unknown based on the flame color.
- Observe continuous vs. line spectra resulting from the emission from gas discharge tubes and identify the unknown element in fluorescent light bulbs.

NOTE: This lab requires observations of colors. Students will work in pairs, so if you happen to be color-blind, you will work with another student to identify the colors.

HAZARDS

- 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.

PROCEDURE

PART 1: FLAME TESTS

(All of the samples must be finely ground powders!)

- 1. Light a Bunsen burner. Adjust the air vent to obtain a blue flame.
- 2. Obtain two small beakers filled about 1/3 full with tap water. One will be used to moisten a wooden splint and the other to extinguish your wooden splint.
- 3. 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 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.
- 4. Place the splint into the clear blue flame and observe/record the color of the flame for that element. Dip the splint into the other beaker of water to douse the flame.
- 5. 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 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 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. Which element is used to make fluorescent light bulbs? Justify your answer.
REPORT SHEETS Experiment 6: Spectroscopy



Name: _____

Section: _____

Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

DATA AND QUESTIONS

PART 1: FLAME TESTS FOR KNOWN ELEMENTS

Metal Ion	Color of Flame	Metal Ion	Color of Flame
Lithium		Copper	
Calcium		Potassium	
Sodium		Iron	
Strontium		Barium	

PART 2: IDENTIFICATION OF SOLID UNKNOWNS

Solid Unknown #	Color of Flame	Element

PART 3: CONTINUOUS AND LINE SPECTRA

(Draw what you see through the spectroscope. Include colors and wavelengths.)

Incandescent light bulb:

Fluorescent light bulb:

- 1. How are these spectra different?
- 2. Which is a line spectrum?

PART 4: IDENTIFICATION OF ELEMENT IN FLUORESCENT LIGHT BULBS

Draw what you see through the spectroscope when you examine each of the following gas discharge tubes. Include colors (use colored pencils). Also, describe the color of the glowing discharge tube.

Helium	Color of the glowing discharge tube:
Spectrum	
Neon Spectrum	Color of the glowing discharge tube:
Mercury Spectrum	Color of the glowing discharge tube:
Hydrogen Spectrum	Color of the glowing discharge tube:

Compare the emission spectrum of these elements with the emission spectrum of the fluorescent light bulb. Based on your results, which element is most likely present in fluorescent light bulbs? (**NOTE:** Other elements not listed here may also be present in the bulbs.) Explain your reasoning.

PRE-LAB **Experiment 6: Spectroscopy**



Name:

Section:

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 3. 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?
- 4. The following diagrams of waves are labeled λ_1 , λ_2 , and λ_3 . Fill in each blank with the best choice.



- 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? _____
- 5. What is the difference between a continuous spectrum versus a line spectrum?

CHEM& 121 WORKBOOK

EXPERIMENT 7 NUCLEAR CHEMISTRY



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INTRODUCTION

Radiation is all around us. There are two main types of radiation: ionizing and non-ionizing. In this lab, we will focus on radioactivity or ionizing radiation (though non-ionizing radiation, such as ultraviolet, can also be harmful.) Radioactivity results from unstable nuclei decaying to more stable isotopes. The Sun produces cosmic rays, there are radioactive isotopes in the walls and air around us, and even people are radioactive. This natural radioactivity is called background radiation, and life on Earth has adapted to it. The variety of life probably comes from the mutations caused by radioactivity.

Nuclear reactions are different from chemical reactions primarily in the fact that the *chemical identity* of atoms can change. Nuclear reactions involve the spontaneous emission of nuclear particles or high energy, in the form of alpha, beta, or gamma radiation. These emissions can be detected by means of a Geiger counter, which counts how many particles it detects per minute (in units of counts per minute, cpm). The Geiger counter won't be able to detect all the particles emitted however. As you can see from the diagram below, radiation is dispersed in all directions—much of it will not get to the detector. Radiation will also collide and get absorbed with surrounding air molecules. Thus, the distance you hold the Geiger counter away from your sample will affect your measurement.



Individual atoms may decay at any time (it is a random process), but with a large enough sample, it can be stated that in a certain period of time, about half of the sample will decay. The amount of time it takes for half the sample to decay is called the half life, and is characteristic of each radioisotope. For example, iodine-131 is used to treat thyroid cancer and has a half life of eight days. It doesn't mean exactly 50% of iodine-131 days after 8 days, but it's approximately half. Many radioisotopes used for medical applications have short half lives and decay very quickly. Many naturally occurring radioisotopes have very long half lives, even thousands or millions of years, and can be useful in dating very old objects (for example, ¹⁴C and ²³⁸U).

We will model half life decay to observe how this happens and learn about sample size and experimental error bars.

OBJECTIVES

In this experiment, you will

- Detect and measure radioactivity with a Geiger counter.
- Measure the local intensity of background radiation.
- Determine the difference in penetrating power between different radioactive particles.
- Simulate the process of half life using tossed pennies.
- Graph your results and gain an understanding of probability and error analysis.

HAZARDS

No hazardous chemicals or glassware are to be used, so goggles are not required. Nuclear samples should not be handled with bare skin. Use the forceps in your drawers. The quantity of nuclear material is very small and within acceptable exposure limits. If you have any questions or health concerns, let the instructor know.

PROCEDURE

NOTE: This lab works well when half the class starts with Part A and the other half starts with Part B, due to the supply of radiation monitors (6 max).

PART A: MEASURING RADIOACTIVITY

1. Use of the Geiger Counter (radiation monitor) and Radioactive Samples

You will connect the radiation monitor to the Lab Quest Pro which will give you an activity reading called "counts per minute (cpm)." Ask your instructor if you have trouble with the setup and use of the Geiger counter or Lab Quest Pro. Use the instructions sheet which comes with the Geiger counter.

Record the isotope for each source in the data table. Using tongs remove one sample at a time from the sample case and place it about a yard from the other sources to cut down on interference. Record cpm with the radiation monitor 1 cm away from the source, and again when it is 10 cm away from the source. **Make sure the window of the radiation monitor is facing the sample!**

2. Background Radiation

Move the radiation monitor away from the sample case, perhaps to the window or take it out to the window in the hallway. Record the cpm. Repeat two more times and average the data.

The actual radiation of a sample is measured by using the observed radiation for the sample minus the background radiation. Does this make sense?

3. Shielding (design your own scientific experiment!)

The greater the penetrating power of the radiation, the more likely the radiation is to go through a material. Which materials are better shields (less penetration)? Select ONE of the radiation sources to use (alpha, beta, or gamma). Select at least three different types of materials provided by your instructor (or ones you have, such as clothing or paper) to devise your own experiment to answer this question. Determine what your control variables are (what is constant in your experiment?), what you are changing (the independent variable), and what you are measuring (the dependent variable). Record your method and data on the report sheet.

PART B: DEMONSTRATION OF HALF LIFE

- 1. Obtain 100 pennies from your instructor and a tub provided in the classroom. Arrange the pennies in a single layer on the bottom of the tub. Carefully lift the tub and give it one shake straight up and down so that some of the pennies jump, but none are lost out of the tub. (Practice a few times before you start counting.)
- 2. Place the pennies all facing heads up in the bottom of the tub. Give the tub one shake. Remove all the pennies that are not heads up. This is one "flip." Record the number of heads left in the tub on the report sheet for this flip. Make sure this number decreases with each flip. Shake again and remove the tails for one more flip. Record the data for each successive flip going down the column. A total of 10 flips will be done for that sample. Repeat the whole process for a total of three samples of 100 pennies each.
- 3. **Graph the data.** Calculate the average value of heads up coins at each flip for all of your samples combined. Graph the average number of heads versus flip # (use graphing area provided on report sheet or use a spread sheet in the lab computers).

Guidelines for a "good" graph:

- Use as much area of the graph paper as possible: Count all the squares in the xand y-axes, and use these totals to estimate the values you should assign for each square along the axes.
- Label these divisions in appropriate increments considering your data (e.g., 0, 5, 10, etc.)
- Include units on your axes.
- Include a title for your graph which describes what your graph conveys.
- Don't make your data points too small or too big. For example, or .
- Don't connect the data points! Use a best fit line where appropriate (see below).
- In this case, the relationship is not linear, so approximate a best-fit curve.

ERROR ANALYSIS

Record the **ranges** of values from **highest number to lowest number** for each flip. (Use the blank space in your data table.) Using the highest and lowest data from each trial, show the **ranges** you found for each flip with **vertical** (up and down) error bars through your average points. See sample to the right.

In a different color, and/or a different style of line, plot the *expected points* for half-life for each trial. (Trial 1 is the first half-life. How many pennies *should* be left?) Explain **best-fit** line.



REPORT SHEETS Experiment 7: Nuclear Chemistry

Name: ______ Section: _____

Lab Partner:

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PART A DATA

1. Geiger Counter and Samples

	a Source	β Source
Element (isotope)		
Counts per Minute at 1 cm		
Counts per Minute at 10 cm		

2. a. Record the background counts for 60 seconds. (Do this three times.)

_____ Average background cpm = _____ **cpm**

b. Use this average background count to report the ACTUAL counts per minute for each sample.

	a Source	β Source
Element (isotope)		
Counts per Minute at 1 cm		
Counts per Minute at 10 cm		

- 3. Identify the variables in your shielding experiment.
 - a. Independent variable: ______ c. Controlled variables: _____
 - b. Dependent variable: _____
- 4. Describe your shielding experiment in 1–2 sentences. Create your own data table below for the shielding experiments you devise.

PART B DATA: HALF LIFE DEMONSTRATION WITH PENNIES

	Sample 1	Sample 2	Sample 3	Average	Range (lowest to highest)
Flip 1					
Flip 2					
Flip 3					
Flip 4					
Flip 5					
Flip 6					
Flip 7					
Flip 8					
Flip 9					
Flip 10					

GRAPHING: AVERAGE WITH ERROR RANGES

Refer to the guidelines on how to prepare a good graph under the "Graph the data" section in the procedure. Do the error analysis for the actual plot (include the error bars), and include the theoretical plot as well. (See the instructions under "Error analysis" in the procedure.)



POST-LAB QUESTIONS

PART A: GEIGER COUNTER AND SAMPLES

- 1. a. Which of the samples gave the most counts (alpha or beta)?
 - b. Based on your answer, which particle appears to have a greater penetration power through air?
- 2. a. Using your most active sample (alpha or beta), describe how the number of counts is affected by the distance of the sample from the counter. **Why does this make sense?**
 - b. What are two ways you might protect yourself from a very active sample?
- Complete the nuclear equations for the spontaneous decay of each sample you observed. (Use the conservation of mass/charge and periodic table to figure this out. Remember, the top number is the mass number and the bottom number is the number of protons or charge.)

alpha emitter: ${}^{210}_{84}$ Po \rightarrow _____ + ${}^{4}_{2}$ He (alpha particle) beta emitter: ${}^{90}_{38}$ Sr \rightarrow _____ + ${}^{0}_{-1}\beta$ (beta particle)

SHIELDING

4. State your conclusion from the shielding experiment based on your results. Is it what you expected? Explain briefly.

PART B: HALF-LIFE DEMONSTRATION

- 5. What does a "heads" penny represent and what does a "tails" penny represent?
- 6. a. **Based on your data**, how many flips were needed to get the number of pennies to 1/4 of the original? (**HINT:** How many half-lives have passed for 1/4 to remain?)
 - b. How many flips should it take to get to 1/4 of the original, ideally (theoretically)?
 - c. If you started with 1000 pennies instead of 100, how many flips would it take to get to ¼ of the original, **ideally** (theoretically)?

ERROR ANALYSIS

- 7. Compare your actual plot to your theoretical plot. Does the theoretical plot fall within the error bars on your actual plot? Does this mean your results match or do not match the expected (theoretical) values?
- 8. It is **unlikely** the theoretical curve on the graph will exactly match the experimental curve. What is it about the nature of radioactivity that results in a difference between the theoretical and experimental plots?
- 9. If you did only one sample instead of an average of three samples, what effect would this have on your experimental plot compared to the theoretical?

PRE-LAB Experiment 7: Nuclear Chemistry

Name:

Section:

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. Look up, in your textbook or other source, how a Geiger Counter works and describe it briefly below. Cite the source(s) you used.
- 2. Will all the radiation that is emitted from a sample get detected by the Geiger counter? If not, why not?
- 3. What is background radiation? Where does it come from? List three sources of background radiation.
- 4. Define half-life in terms of radioactive isotopes.
- If you have 100 radioactive atoms, will exactly 50 decay after one half-life? (HINT: If you have 100 coins and you toss them simultaneously, will exactly 50 flip over?) Answer "yes" or "no" and explain your choice. (HINT: Think about this realistically, not theoretically.)
- 6. Predict which of the following materials has the greatest shielding effect?

CHEM& 121 WORKBOOK

EXPERIMENT 8 MOLECULAR MODEL BUILDING



The first set of molecules we will examine contains only two atoms. These are considered to have a linear molecular shape. Since each molecule only consists of one bond, if the bond is polar the molecule is polar.

PART B

- 1. The molecules in this part contain more than two 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 (two for a double bond; three for a triple bond).
 - Use the following colored spheres:
 - C, Si = black
 - N, P, As = blue
 - H = white

- 0, S = red
- Cl, Br, I = use 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.	Hummer H Hummer H
Trigonal pyramidal A four-sided figure. Three sides are isosceles triangles. They are sitting on an equilateral triangular base.	HUMPEN H HUMPEN H
Bent A nonlinear arrangement of atoms.	H H H

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.	F B F B F F
Bent A nonlinear arrangement of atoms.	

REPORT SHEETS Experiment 8: Molecular Model Building



Name: ______ Section: _____

Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PART A—PRE-LAB

Species/ Total v.e.	Draw the Lewis Dot Structure	Name of Molecular Shape	Polar or Nonpolar
H ₂ 2 v.e.			
Br ₂			
HBr			
N ₂			
CO			

CONCLUSIONS

- 1. If only two atoms are bonded, the molecular shape will always be ______.
- 2. If the two bonded atoms are identical, the molecule is polar / nonpolar.
- 3. If the two atoms are different, the molecule is polar / nonpolar. The atom with the greater / smaller electronegativity will be partially negative (δ -).

PART B			
Polar or Nonpolar			
Name of Molecular Shape			

REPORT SHEETS | EXPERIMENT 8: MOLECULAR MODEL BUILDING

Polar or Nonpolar				
Name of Molecular Shape				
Name of Electron Geometry				
3-D Sketch				
Bond Angles				
# of Electron Groups				
Lewis Dot Structure				
Species/ Total v.e.	CF ₄ 32 v.e.	CHF ₃	0F ₂	H ₂ S

Polar or Nonpolar				
Name of Molecular Shape				
Name of Electron Geometry				
3-D Sketch				
Bond Angles				
# of Electron Groups				
Lewis Dot Structure				
Species/ Total v.e.	PCI3	CS ₂	HCN	COCI ₂

This page contains polyatomic ions. They have gained or lost valence electrons to become charged.

Polar or Nonpolar	*	*	*
Name of Molecular Shape			
Name of Electron Geometry			
3-D Sketch			
Bond Angles			
# of Electron Groups			
Lewis Dot Structure			
Species/ Total v.e.	NO ₂ -	C0 ₃ ^{2–}	NH4 ⁺

* Even if the bond dipoles cancel, polyatomic ions have a net charge and can be considered polar.

**Nitrite (NO $_2^-$) and carbonate (CO $_3^{2-}$) have resonance structures—different but equivalent Lewis structures.

CHEM& 121 WORKBOOK

EXPERIMENT 9 INTERMOLECULAR FORCES



PART A: SURFACE TENSION

Surface tension is the tendency of a substance to attract itself and minimize its surface area.

SUPPLIES NEEDED

1 penny
 a pipet
 some water

PROCEDURE

Use the pipet as a dropper to deliver drops on top of a penny. As you do that, you will see water bead up due to its surface tension. Count how many water droplets it takes before the bead breaks!

PART B: MILK RAINBOW

Food coloring is an **aqueous** solution (hydrophilic) while milk contains **non-polar** fat molecules (hydrophobic). It will help to review the structure of soap molecules.



"Polar Head" (hydrophilic)

"Non-Polar Head" (hydrophobic)

A surfactant is a substance that reduces surface tension. Soap is a great example of a surfactant.

SUPPLIES NEEDED

- 2% milk (or you can use milk with a higher fat content)
- Food coloring
- Liquid dish soap

Small plate or saucer

PROCEDURE

Pour enough milk into a saucer to cover the bottom. Add 4 to 8 drops of different colors of food coloring to the milk, placing the drops on separate areas of the milk.

REPORT SHEETS Experiment 9: Intermolecular Forces



Name: _____ Section: _____ Lab Partner:

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PART A: SURFACE TENSION

1. How many droplets did it take before the bead broke? _____ Sketch what the bead looked like:

2. Why does water have such a high surface tension? What are the intermolecular forces of attraction in water called?

3. Does this lead you to believe polarity plays a factor in surface tension? Which would tend to have a higher surface tension: polar molecules or non-polar molecules?

4. Draw two water molecules (and their dipoles) and show the intermolecular forces between them in the correct orientation.

PART B: MILK RAINBOW

1. Describe what it looks like, especially what the food dye does (before adding soap or after adding color):

2. Add 1 drop of soap to the saucer and wait for several seconds. Describe what happened:

3. Milk is a mixture that contains mostly water and some fat. Is soap attracted to water, fat, or both? (If you're not sure, try adding some food coloring to water. Does it dissolve? Think about "like dissolves like.") **Explain.**

4. A surfactant is a substance that reduces surface tension. Soap is a great example of a surfactant. Consider that milk is made mostly of water. How might soap break the intermolecular forces between water molecules? (Look at the structure of a soap molecule provided at the beginning of this activity, and use the terms polar and nonpolar OR hydrophilic and hydrophobic in your answer.)

5. Is soap necessary to remove salt (NaCI) from your hands? **Explain.** (**HINT**: Is salt hydrophilic or hydrophobic or both?) CHEM& 121 WORKBOOK

EXPERIMENT 10 EXPLORING GASES



Adapted from Experiments 6 and 7, *Chemistry with Vernier* © Vernier Software & Technology

Note to Students: Check with your instructor to see which parts of this lab (Parts A, B, or C) you will complete.

INTRODUCTION

Gases are made up of molecules that are in constant motion and exert pressure when they collide with the walls of their container. The velocity (speed) and the number of collisions of these molecules is affected when the temperature of the gas increases or decreases, or if the volume of gas changes as it is compressed or expanded.

OBJECTIVES

In this experiment, you will

- Investigate what pressure is and how various substances change their properties with a changes in pressure.
- Determine the relationship between the temperature of a gas sample and its pressure.
- Determine the relationship between the volume of a gas sample and its pressure.
- Practice graphing data to determine if variables are inversely or directly proportional.
- Describe how the macroscopic changes that are observed result from a model of how molecules move at the submicroscopic level.

HAZARDS

Even though dangerous chemicals are not used, glassware is used so make sure to wear your goggles this entire period.

PROCEDURE

Parts A, B, and C can be done in any order. However, it helps to get the warm water needed for Part B started on the hot plate earlier in the period.

PART A: EXPLORING PRESSURE

- 1. Place a flask (200–300-mL) with about 100 mL of water on a hot plate and cover its opening with a balloon. Turn on the hot plate. While waiting, predict what you expect to happen on the report sheet. Then describe what you actually see and draw a before and after picture of what you think the molecules are doing.
- 2. Place a similar flask with 100 mL of water on a hot plate. Bring the water to a boil. Remove from the hot plate. Be very careful and cover the opening with a balloon. (CAUTION: The flask can be very hot!) Wait for the flask and the water to cool. While waiting, predict what you expect to happen on the report sheet. Then describe what you actually see and draw a before and after picture of what you think the molecules are doing.

OPTIONAL: Your instructor may have a few other demonstrations to show you. You can experiment with various substances/objects on your own—these objects will be placed at various stations in the lab. Students will rotate to various stations in groups of four. Your job is to predict what will happen at each station, and explain how it works with words and drawings. Have fun!

PART B: TEMPERATURE VS. PRESSURE

Using the apparatus shown (right), you will place an Erlenmeyer flask containing an air sample in water baths of varying temperature. Pressure will be monitored with a Pressure Sensor and temperature will be monitored using a Temperature Probe. The volume of the gas sample and the number of molecules it contains will be kept constant (the same).



- 1. Put about 800 mL of hot tap water into a 1-L beaker and place it on a hot plate. Turn the hot plate to a high setting but keep an eye on it—you want hot water to be between 50 °C and 70 °C for this experiment.
- 2. Prepare a second 1-L beaker with about 700 mL of cold tap water and ice to make 800 mL.
- 3. Put about 800 mL of room-temperature water into a third 1-L beaker.
- 4. Obtain a Lab Quest Pro, a temperature probe, and a pressure sensor.
- 5. Set up the pressure sensor:
 - a. Plug the Pressure Sensor into a channel in the Lab Quest.
 - b. Plug the Temperature Probe into a channel in the Lab Quest.
 - c. Check that a rubber-stopper assembly with a piece of heavy-wall plastic tubing connected to one of its two valves is attached to the Pressure Sensor. Leave its two-way valve on the rubber stopper open (lined up with the valve stem as shown in the figure) until Step 6b.
- 6. Prepare the gas sample in the flask.
 - a. Insert the rubber-stopper assembly into a 125-mL Erlenmeyer flask.
 Important: Twist the stopper into the neck of the flask to ensure a tight fit. MAKE SURE THERE ARE NO LEAKS!!!
 - b. Close the 2-way valve above the rubber stopper—do this by turning the valve handle so it is perpendicular with the valve stem itself (as shown in the figure below). The air sample to be studied is now confined in the flask.



- c. Do NOT open this valve for the rest of the experiment!
- 7. Collect pressure *vs.* temperature data for your gas sample. **DO NOT ZERO** THE PRESSURE READINGS!
 - a. Place the flask into the ice-water bath. Make sure the entire flask is covered (see figure above). Stir.
 - b. Place the temperature probe into the ice-water bath. When the temperature stabilizes, record it in your data table.
- 8. Repeat Step 7 using the room-temperature bath.

- 9. a. Check the temperature of the hot water bath. If it is above 70 °C, the stopper will pop out! Add cold water or ice to bring the temperature below 70 °C.
 - b. Repeat Step 7 using the hot-water bath: Use a utility clamp to hold the flask and temperature probe in the boiling-water bath. To keep from burning your hand, the tubing, or the wire, hold the tubing, the probe wires, and the clamp together in one hand using a glove or a cloth. After the temperature and pressure have stabilized, repeat Step 7. CAUTION: Do not burn yourself or the probe wires with the hot plate.

PART C: VOLUME VS. PRESSURE

The gas we use will be air, and it will be confined in a syringe connected to a Pressure Sensor. When the volume of the syringe is changed by moving the piston, a change occurs in the pressure exerted by the confined gas. This pressure change will be monitored using a Pressure Sensor. It is assumed that temperature will be constant throughout the experiment.

- 1. Prepare the Pressure Sensor and an air sample for data collection.
 - a. Plug the Pressure Sensor into a channel of the Lab Quest. You don't need a temperature probe for this one.
 - b. With the 20-mL syringe disconnected from the Pressure Sensor, move the piston of the syringe until the front edge of the inside black ring is positioned at the 10.0 mL mark, as shown in the figure.
 - c. Attach the 20-mL syringe directly to the white valve stem of the Gas Pressure Sensor with a gentle half-turn, as shown in the figure.



- 2. Collect the pressure *vs.* volume data.
 - a. Move the piston to position the front edge of the inside black ring at the 5.0-mL line on the syringe. Hold the piston firmly in this position until the pressure value stabilizes but avoid placing your hand around the barrel of the syringe.
 - b. When the pressure reading has stabilized, record it in your data table.
 - c. Continue the procedure for volumes of 10.0, 15.0, and 20.0 mL.

REPORT SHEETS Experiment 10: Exploring Gases **B**

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 PART A

1. Place a flask (200–300-mL) with about 100 mL of water on a hot plate and cover its ending with a balloon. Turn on the hot plate. While waiting, **predict what you expect to happen.**

Describe what you actually see and draw a before and after picture of what you think the molecules are doing. To show movement of molecules, use long arrows \longrightarrow for faster molecules and shorter arrows \rightarrow for slower molecules.

Before	After

2. Place a similar flask with 100 mL of water on a hot plate. Bring the water to a boil. Remove from the hot plate. Cover the ending with a balloon. Wait for the flask and the water to cool. While waiting, **predict what you expect to happen**.

Describe what you actually see and draw a before and after picture of what you think the molecules are doing.

Before	After

DATA AND QUESTIONS FOR PART B

Water Bath	Temperature (°C)	Pressure (kPa)
lce		
Room Temperature		
Hot Water		

Plot your data below. Make sure you provide axes labels and a best-fit line. Is the relationship linear or non-linear? Directly proportional or inversely proportional?



- 1. Based on the data and graph that you obtained for this experiment, express in words the relationship between gas pressure and temperature.
- 2. Describe what the gas molecules are doing when the temperature increases. How does that affect the pressure of the gas in the flask?
- 3. If the gas was held in a balloon rather than in a flask, and the temperature increased significantly, how would that affect the pressure of the gas? Explain.

Volume (mL)	Pressure (kPa) (to 3 sig figs)	P × V (multiply P and V)	<i>P/V</i> (divide <i>P</i> and <i>V</i>)
5.0			
10.0			
15.0			
20.0			

DATA AND QUESTIONS FOR PART C

Plot your data below. Do not draw a best-fit line until you answer the next question. It might be difficult to tell if the relationship is directly or inversely proportional.



- 1. Check your data table above. If $P \times V$ is nearly constant for all data pairs, then the relationship is *inversely* proportional. If P/V is nearly constant, the relationship is *directly* proportional. After you determine which it is, draw the best-fit line appropriate for this plot (a straight line for a direct plot; a curve for an inverse plot).
- 2. Based on the data and graph that you obtained for this experiment, express in words the relationship between gas volume and pressure.
- 3. Describe what the gas molecules are doing when the volume increases. How does that affect the pressure of the gas in the syringe?

PRE-LAB Experiment 10: Exploring Gases **•**•

Name:

Section:

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. Define "pressure" and look up at least three units of pressure you may encounter.
- 2. Define "temperature" and look up at least three units of temperature.
- 3. Circle the correct answer for each statement below.
 - a. As gas molecules get colder (decrease in temperature), do they move faster or slower? Will the pressure exerted by these molecules on the container increase or decrease?
 - b. As gas molecules are compressed (decrease in volume), do they collide more or less? Will the pressure exerted by these molecules on the container increase or decrease?
- 4. If two variables increase or decrease together and with a constant ratio, the variables are said to be directly proportional to each other. If two variables oppose one other (one increases, the other decreases), and their product is a constant, the variables are inversely proportional to each other.

Based on this description, label each graph by plotting two different sets of variables as "directly proportional" or "inversely proportional."




EXPERIMENT 11 COUNTING PARTICLES



Count the number of items in a bag of dry goods without actually counting each individual item.

RULES

- Each team selects a bag of dry goods (grains or beans).
- Your team is prohibited from counting each individual item in the bag; however, items may be removed from the bag in multiples of 23. Each group of 23 items will be referred to as a "unit."
- The entire collection of items may be taken out of the bag, as long as the individual items are not counted (unless in units of 23) and all of the items are eventually returned to the bag.

REPORT

- 1. Describe the process that your team used to determine how many items were in your team's bag.
- 2. Use dimensional analysis to show how your group calculated the number of items in your team's bag.
- Determine how many units were in your team's bag. (See above for the definition of a unit.)

- 4. Create a table of dry goods.
 - a. Each group needs to come up with a symbol using letters to represent their dry good and be ready to report the mass of 1 unit of their dry good. (See above for the definition of a unit.)
 - b. Collect info from all groups.
 - c. Number the dry goods using whole numbers starting with 1, in order of increasing unit mass.

Use dimensional analysis to perform the following calculations:

- 5. How many bags would be equivalent to one million of your group's item?
- 6. How many bags would be equivalent to one billion of your group's item?
- 7. How many bags would be equivalent to 6.022×10^{23} of your group's item?
- 8. Flatten out the bag into a rectangular shape and measure in centimeters the length of the longest side of the bag. If the number of bags calculated in #7 were laid end-to-end, how many times the distance from Earth to the sun (9.3×10^6 miles) would the bags span?

CONNECTION TO CHEMISTRY

The chemists' counting unit is called the mole. The mole is defined as the number of atoms in exactly 12 g of carbon-12 atoms, which turns out to be 6.022×10^{23} , which is referred to as Avogadro's number.

1 mole = 6.022×10^{23} objects

Because the atomic mass unit (amu) is also based on carbon-12, we can extend the concept of the mole to any element. The mass written below each element on the Periodic Table not only represents the mass of an average atom of the element in amu, but also represents the mass in grams of 1 mole of atoms (6.022×10^{23} atoms) of that element.

EXAMPLE USING HYDROGEN

1 mole of H atoms = 1.01 g and 1 mole H atoms = 6.022×10^{23} H atoms

Thus...

 6.022×10^{23} H atoms = 1.01 g

What would a sample of 1.004×10^{23} aluminum atoms look like? (**HINT**: What's the ratio of this number to Avogadro's number?)

REPORT SHEETS Experiment 11: Counting Particles

Name:	 Section:	
Lab Partner:		

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

1. Describe the process that your team used to determine how many items were in your team's bag.

2. Use dimensional analysis to show how your group calculated the number of items in your team's bag. Solve for number of units.

Total g =
$$\left(\frac{\text{unit}}{g}\right)$$
 = # of units

3. Determine how many units were in your team's bag. (See above for the definition of a unit.) _____

of units =
$$\left(\frac{23 \text{ items}}{1 \text{ unit}}\right)$$
 = _____ items

4. Create a table of dry goods.

Use dimensional analysis to perform the following calculations:

5. How many bags would be equivalent to one million of your group's item?

6. How many bags would be equivalent to one billion of your group's item?

7. How many bags would be equivalent to 6.022×10^{23} of your group's item?

8. Flatten out the bag into a rectangular shape and measure in centimeters the length of the longest side of the bag. If the number of bags calculated in #7 were laid end-to-end, how many times the distance from Earth to the sun (9.3×10^6 miles) would the bags span? **HINT:** How many miles correspond to the answer in Question 7?

CHEM& 121 WORKBOOK

EXPERIMENT 12 IRON IN FORTIFIED CEREAL

INTRODUCTION

Iron is one of the minerals needed for good health. In your body iron is part of a molecular unit called heme which is critical to the transport of oxygen by hemoglobin and to the temporary storage of oxygen in myoglobin. The oxygen is carried by the single Fe⁺² in the center of the heme molecule. Iron is also an important part of some proteins involved in cellular respiration, enzymes, and bacterial growth.

Iron occurs naturally in red meats, liver, eggs, raisins, and leafy green vegetables (such as spinach). Iron has also been added to many cereals. We will examine some of those cereals in this experiment.

PRE-LAB (OPTIONAL)

If you have on-hand* a cereal which is iron-fortified, bring to class a sandwich bag containing one serving of the cereal. **Record information about your cereal on the report sheets.** (*Please do not purchase a box of cereal for this experiment. Hopefully, many of you have this kind of product in your home. If you don't have cereal at home, come to class and we can provide you with a sample.)

PROCEDURE

- 1. After recording the Pre-Lab information about your cereal, place your sample in a mortar and pestle. Grind the cereal to a fine-textured mixture.
- 2. Pour the cereal into a large beaker and add approximately 500 mL of water. Drop into the mixture a magnetic stirring bar and place the beaker on an automatic stirring plate. Turn on the stirrer (NO HEAT) and make sure the stirring bar is turning smoothly. Continue to stir for 60 minutes. Add more water if your mixture is very thick. (HINT: Your cereal may be difficult to stir. The slowest speed usually works best.)
- 3. After 60 minutes, carefully pour the water-cereal mixture into a waste bucket making very sure the stirring bar is not thrown away. If your stirring bar has a great amount of cereal stuck to it, add a small amount of water to the beaker with the magnetic stirring bar , return the beaker to the automatic stirring plate, and stir for an additional minute or two. Handle your stirring bar carefully as you don't want to disturb the iron on the bar.
- 4. Place the stirring bar on a piece of filter paper. Write the name of your cereal on the filter paper, the Percent of the Daily Value, and the form of iron identified on the list of ingredients on your cereal box. Place your sample on a table to share with other students in the class.
- 5. Your instructor has also placed an iron tablet in a beaker of water and has performed the same experiment as a control.
- 6. Compare your magnetic stirring bar with the magnets used with the other cereal samples in the class. Also, examine the magnet that was used with the iron tablet. As a class, organize the samples based on the amount of iron on the magnet. Use a scale from 0–2 with 2 representing the most iron and 0 representing no iron.

REPORT SHEETS Experiment 12: Iron in Fortified Cereal



Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PRE-LAB

Read Introduction section for instructions.

Name of cereal ______ Serving size _____

Percent of the Daily Value for iron _____

(For best results in this experiment, choose a cereal with at least 45% of the Daily Value.)

Look at the list of ingredients. What form of iron is in your cereal?

(**HINT**: Look for terms such as mineral iron, reduced iron, or a ferrous-or ferric-containing compound.)

DATA

List the samples you have identified as "2." Identify the form of iron in each sample and the % of the Daily Value.

Sample	% of the Daily Value	Form of Iron

List the samples you have identified as "1." Identify the form of iron in each sample and the % of the Daily Value.

Sample	% of the Daily Value	Form of Iron

List the samples you have identified as "0." Identify the form of iron in each sample and the % of the Daily Value.

Sample	% of the Daily Value	Form of Iron

Check the label.

POST-LAB QUESTIONS

1. What generalizations can you make based on these observations?

2. Please use your text book or the internet to find the formula and charge of

- a. the ferrous ion: ______ c. the phosphate ion: _____
- b. the ferric ion:______ d. the sulfate ion: _____
- 3. Use your knowledge of ionic nomenclature to determine the chemical formula of these compounds:
 - a. Ferrous chloride:
 - b. Ferric phosphate:
 - c. Ferrous sulfate:
 - d. Ferric oxide:

CHEM& 121 WORKBOOK

EXPERIMENT 13 REACTION OF HYDROGEN AND OXYGEN



INTRODUCTION

Hydrogen is a clear, colorless gas which is said to be "combustible," meaning that it can burn quite readily. Oxygen is also a clear, colorless gas that is said to "support combustion," meaning that it must be present for combustible materials to burn. In this lab, you will be generating, collecting, and testing hydrogen and oxygen gas.

Hydrochloric acid is reacted with zinc to generate hydrogen. In general, any strong acid and many different metals react to produce hydrogen gas. Hydrogen peroxide is added to manganese dioxide to generate the oxygen. Hydrogen peroxide decomposes by itself to produce water and oxygen gas at a slow, imperceptible rate; the manganese dioxide acts as a catalyst to speed up this reaction. By collecting and pop-testing (igniting) different hydrogen/oxygen mixtures, you will audibly compare them to determine the most reactive (loudest) mixture, which will be the optimal stoichiometric ratio for reaction.

OBJECTIVES

In this lab you will:

- Write chemical equations to represent chemical reactions.
- Manipulate the ratio of reactants to investigate the stoichiometric relationship of a chemical reaction.
- Determine the limiting reactant in a chemical reaction.

HAZARDS

Since this lab is performed on the microscale level, the explosions, though potentially loud, are safe; however, sometimes the explosions can occur in the pipet bulb. This can be avoided by moving the bulb away from the flame immediately after squeezing the bulb with-out allowing gases back into the bulb.

The two solutions used in this lab, hydrochloric acid (HCl) and hydrogen peroxide (H_2O_2) , can cause serious damage should they come in contact with your eyes. Use them with caution!

PROCEDURE

- 1. Fill a 250-mL beaker with tap water. This will act as a test tube holder, a temperature regulator, and a water reserve during the experiment.
- 2. Label two test tubes; one is a hydrogen generator, and the other an oxygen generator.
 - To make the hydrogen generator: place approximately six pieces of zinc shot into the appropriately labeled test tube and add enough 3 M HCl to fill the test tube to within 4 cm from the top. Top with a one-hole stopper.
 - To make the oxygen generator: place a small spatula full of MnO₂ into the appropriately labeled test tube. Add about 3 mL of 7.5% hydrogen peroxide. Top with a one-hole stopper. When oxygen production slows down, add another 3 mL of 7.5% hydrogen peroxide. Repeat as needed; when tube gets to within 4 cm from the top, start a new tube.
 - Place both generators into the 250-mL beaker.
- 3. Cut off the long end of a polyethylene pipet leaving approximately a one-centimeter tip. Using a push pin, pierce the pipet near the end of the tip. (Alternately, this may be done for you ahead of time.) Mark the polyethylene pipet with a permanent marker to show approximately six equal-volume increments. This pipet will be referred to as the "collection bulb."





4. Fill the collection bulb with water. Do this by submerging it in water (open end up) and gently squeezing the bulb.

5. Insert the tip of the collection bulb into the one-hole stopper on the top of the hydrogen generator. Begin collecting hydrogen gas. You can sometimes speed up this collection process by using your finger to put slight pressure on the top of the bulb. This helps create a better seal between the bulb and the one hole stopper. (Do not squeeze the bulb too much, as this may introduce air into the bulb.)

Read these instructions fully before performing the next part:

After you have collected a bulb of hydrogen gas, hold the bulb with the narrow end down to prevent losing hydrogen gas. (There may be a few drops of water, but this is ok. The water will act as a seal to prevent the gas from escaping into the room.) Keeping the bulb vertical, move it toward the flame.



Quickly rotate it to a horizontal position with its mouth roughly 3–4 cm from the mid-section of the flame. Gently squeeze the contents of the bulb into the flame and observe. (If nothing happens, try squeezing the bulb again.)

Avoid putting the bulb directly into the flame. It will melt and possibly burn. Should this happen, quench the tip in a beaker of water. Repeat this step until you obtain reproducible results.

6. Repeat the above, but this time collect oxygen gas.

What seems different this time?

- 7. Repeat the above but this time collect and test all different possible ratios of hydrogen and oxygen gas in the *same* bulb. Do this by transferring the bulb from one generator to the other. During this transfer, hold the pipet with the open end "down."
- Put the liquid wastes into the proper waste container in the hood. Note that there are two separate waste containers so that the zinc and MnO₂ may be collected for reuse. The pipet bulbs may also be collected for reuse.

REPORT SHEETS Experiment 13: Reaction of Hydrogen and Oxygen

Name:	Section:	
l ah Partner		

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

1. Use this template to create a bar graph showing the relative loudness of each of the samples that you tested (including the pure hydrogen and oxygen).



2. Write a balanced equation for the reaction taking place inside the hydrogen generator (the test tube containing Zn and HCI).

Solid zinc and aqueous hydrochloric acid react to generate hydrogen gas and aqueous zinc chloride.

3. Write a balanced equation for the reaction taking place inside the oxygen generator (the test tube containing H_2O_2). **HINT:** MnO_2 is a catalyst for this decomposition reaction. It is neither a reactant nor a product. You do NOT need to display it in the chemical equation.

Aqueous hydrogen peroxide decomposes to create liquid water and oxygen gas.

4. Based on your results, what ratio of hydrogen to oxygen (H₂:O₂) produced the most explosive mixture? ____: ____

Compare your data with others. Give two possible reasons why the results may vary among different teams.

What ratio was the most explosive for the class overall? ____: ____

5. Write a balanced equation for the reaction of hydrogen and oxygen to make water.

Based on the balanced equation, what should be the most explosive ratio of hydrogen to oxygen? **Explain. NOTE:** The ratio of the reacting volumes of gases is equal to the mole ratios of the reacting gases.

- 6. Did you find any reaction bulbs that produced no explosion at all? Explain how that could happen?
- 7. Why don't the hydrogen and oxygen in the collection bulb react as soon as they mix? What role does the flame play?
- For which ratios of H₂:O₂ was hydrogen the limiting reactant? ______
 Show one calculation to support your answer.

For which ratios of H₂:O₂ was oxygen the limiting reactant?

Show one calculation to support your answer.

PRE-LAB Experiment 13: Reaction of Hydrogen and Oxygen



Name: ______ Section: _____

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

1. Write a balanced equation for the reaction of hydrogen gas and oxygen gas to form liquid water. Include physical states.

2. List all the ways you could classify the above reaction. (What type(s) of reaction?)

3. When collecting the hydrogen and oxygen gases, the collection bulb is initially filled with water. What happens to the water as the gas is collected? (Review the procedure.)

4. Challenge question: If one mole of hydrogen gas is reacted with one mole of oxygen gas, which one will run out first (called the limiting reactant)? Explain.

HINT: It might help to think of this at the molecular level. Draw molecular level diagrams and think of it in terms of molecules instead of in moles (If one *molecule* of hydrogen gas is reacted with one *molecule* of oxygen gas ...)

CHEM& 121 WORKBOOK

EXPERIMENT 14 STOICHIOMETRIC ANALYSIS OF AN ANTACID



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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 ba2lanced equations.

- A. $CH_3COOH(aq) + NaHCO_3(s) \rightarrow 2 CO_2(g) + CH_2O(aq) + Na^+(aq) + 3H^+(aq)$
- B. $CH_3COOH(aq) + NaHCO_3(s) \rightarrow CO_2(g) + H_2O(l) + CH_3COO^-Na^+(aq)$
- C. $CH_3COOH(aq) + 2 NaHCO_3(s) \rightarrow CO_2(g) + Na_2CO_3(aq) + H_2O(I) + 2 CH_2O(aq)$

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 (CH₂O), but Reaction B doesn't. The products of Reaction A are acidic (H⁺ is produced); those of B and C are basic (CH₃COO⁻Na⁺ and Na₂CO₃ 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 NaHCO₃ produces 2 moles CO₂
- In Reaction B, 1 mole NaHCO₃ produces 1 mole CO₂
- In Reaction C, 2 mole NaHCO₃ produces 1 mole CO₂

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:

mass of $CO_2 = \begin{pmatrix} mass of reaction \\ mixture before reaction \end{pmatrix} - \begin{pmatrix} mass of reaction \\ mixture after reaction \end{pmatrix}$

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 (NaHCO₃) 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. Record the mass of NaHCO₃ indicated on the package on your Lab Report.

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 14: Stoichiometric Analysis of an Antacid

Name: _____ Section: _____ Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

PART A: REACTION STOICHIOMETRY

	Tablet 1	Tablet 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 NaHCO₃, calculate the number of moles of NaHCO₃.
- 4. Compare the calculated number of average moles of CO₂ and NaHCO₃ 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.

	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		

PART B: ANALYSIS OF AN ANTACID

*Calculate by adding 0.0040 moles to the "Moles of CO₂ Gas Released."

**Use the balanced chemical equation from Part A for the mole-to-mole ratio to convert moles of CO_2 (corrected) to moles of NaHCO₃.

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 NaHCO₃ in the balanced chemical equation that you determined in Part A, calculate the moles of NaHCO₃ present in the tablet.

4. Convert the moles of NaHCO₃ in the tablet to mass of NaHCO₃ in the tablet.

5. Calculate for the average mass of NaHCO₃ 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.

% error (experimental value – actual value) actual value × 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. How would this affect the perceived mass of CO₂? Would your final calculated mass of sodium bicarbonate in the tablet be *artificially high* or *artificially low* as a result of this splashing? Choose one, and explain why.

PRE-LAB Experiment 14: Stoichiometric Analysis of an Antacid

Name: ______ Section: _____

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. When acetic acid and sodium bicarbonate react, why does total mass decrease?
- Suppose you are running a trial of the reaction in Part A. You use 2.80 g of NaHCO₃, 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 NaHCO₃ to moles using the molar mass of NaHCO₃.
 - 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 NaHCO₃ and CO₂. (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 NaHCO₃ and CO₂ in the three balanced chemical equations given in the Introduction. Which of the three possible reactions discussed is consistent with these results? **WRITE THE EQUATION HERE**.

NOTE: The result of reaction C in the Introduction is not necessarily the correct answer for the lab. It is a hypothetical situation for practicing the calculations required in this lab. Do not use this actual result in the lab!

CHEM& 121 WORKBOOK

EXPERIMENT 15 IONIC SOLUTIONS (ELECTROLYTE SOLUTIONS)



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INTRODUCTION

Molecular compounds are made up of molecules, while ionic compounds are made up of ions.

Many ionic compounds dissolve in water; many do not. If an ionic compound dissolves in water, it separates into individual charged ions. For example, pictured below is a solution of sodium carbonate (N_2CO_3) dissolved in water. The resulting solution is composed of separate sodium ions and carbonate ions surrounded by water molecules.

The following chemical equation communicates how the soluble ionic compound, sodium carbonate, separates into sodium ions and carbonate ions. The notation "(*aq*)" means "aqueous" or that the ion is dissolved in water. Note that water is not written as a reactant, but over the reaction arrow.



$$\operatorname{Na_2CO_3}(s) \xrightarrow{\operatorname{H_2O}} 2\operatorname{Na^+}(aq) + \operatorname{CO_3^{2-}}(aq)$$

Once ionic compounds are dissolved, the ions in solution may undergo further chemical reactions with other substances, including neutralization, precipitation, oxidation-reduction, and other reactions.

One technique that can be used to detect the presence of ions is conductivity since charges in motion conduct electricity. Soluble ionic compounds form solutions containing mobile ions that conduct electricity and are therefore referred to as **electrolytes**. In contrast, insoluble ionic compounds do not conduct electricity and are called **nonelectrolytes** because no separate ions are formed in solution.

Beyond being used to classify electrolytes and nonelectrolytes, conductivity is proportional to the concentration of ions, so it can also be used to determine the actual concentration of ionic compounds in water. Conductivity testing is simple, sensitive, and rugged/ inexpensive equipment can be used. For these reasons it is used for a wide variety of field and industrial analyses.

Molecular compounds are not made up of charged particles; therefore, they cannot conduct electricity and are nonelectrolytes. However, there is an important class of molecular compounds that can form ions via a chemical reaction when they dissolve in water. If each molecule separates into ions, the compound is called a "strong electrolyte", but if the molecules of a compound produce only a few ions, it is called a "weak electrolyte." Soluble ionic compounds are also considered "strong electrolytes."

For electrolytes, conductivity depends on concentration. In this lab you will measure the conductivity of a solution with some initial concentration, and then you will dilute the solution by adding solvent. The concentration of the original solution and diluted solution is determined by the following equations:

Original solution: the initial concentration, C =

$$C_{i} = \frac{\text{mass of dry NaCl (in grams)}}{\text{volume of solution (in mL)}} \times 100\%$$

Diluted solution: the final concentration,

$$C_f V_f = C_i V_i$$
 therefore, $C_f = C_i \frac{V_i}{V_f}$

where C_f and C_i are the final and initial concentrations, V_i is the initial volume, and V_f is the final volume. (*Notice that the units for* V will cancel.)

OBJECTIVES

In this experiment, you will

- Classify substances as strong, weak, or non-electrolytes.
- Use conductivity to observe the process of dissolving an ionic compound.
- Learn and practice the technique of dilution.
- Observe the relationship between concentration of an ionic substance and conductivity.

HAZARDS

Hydrochloric acid can cause chemical burns on the skin and damage eyes. Wear goggles and wash your hands after using.

PROCEDURE

PART A: IONIC SOLUTIONS

- 1. Obtain a large beaker and label it "Rinse." All subsequent rinses from Part A and B are to be collected in this rinse beaker, and then later emptied into the sink.
- 2. Obtain a LABQuest and conductivity probe (also used in Part B). Set up the LABQuest to recognize the conductivity probe.
- 3. Make sure the probe is set to read from 0–20,000 μS/cm (μS/cm = microSiemens per centimeter, a unit of conductivity). If your values are not in this range, the post-lab questions will be difficult to answer!
- 4. Thoroughly rinse the conductivity probe with distilled water into your waste beaker. Use a Kimwipe or paper towel to pat it dry. Hold the probe in the air and tap the display until you get an option to "zero." Select it. The reading in air should now read zero. This is called calibration and it **should only be done once** for each experiment!
- 5. There are six solutions circulating about the room. The order you use them does not matter. You will read the conductivity of each substance in its vial by submerging the probe into the liquid in the vial, and slightly stirring it until you get a stable reading. Make sure you rinse the probe with distilled water and pat it dry before/after every measurement so you avoid contamination of the sample vials. (Avoid getting water into the vial, as it dilutes the samples!) Rinsings go into the waste beaker.
- 6. Record the conductivity values you obtain for each of the solutions in your data table.
- 7. Rinse the probe thoroughly when finished. The rinses can be poured into the sink.

PART B: CONDUCTIVITY ANALYSES

All wastes from this part **only** may be emptied into the laboratory sink. You may find it convenient to use a waste beaker, and then to empty this into the sink.

- 1. Obtain a LABQuest and conductivity probe. Set up the LABQuest to recognize the conductivity probe. Thoroughly rinse the conductivity probe with distilled water into a large waste beaker.
- 2. Place a dry, 250-mL beaker on a magnetic stirring plate and add a magnetic stirring bar. Clamp the conductivity probe so that it is near the wall of the beaker and lowered almost to the bottom of the beaker.

- 3. Carefully measure 100 mL of deionized water using a 100-mL graduated cylinder. Record the **actual** volume to the closest 0.1 mL. Pour this into the 250-mL beaker.
- 4. Place the probe into the above beaker. (Using a 0–20,000 μS/cm range, it will likely read between 0 and 100 μS.) Calibrate the probe as you did earlier by setting this to zero. Record this value (0 μS/cm) in the data table for Part B. Do not re-zero it again after this point!
- 5. Obtain a piece of weigh paper. Use your spatula to weigh out between 0.100 and 0.150 g sodium chloride (NaCl) onto the weigh paper. Record the **actual** mass. Do NOT return any NaCl to the original container to prevent contamination; put extra material into your waste container or give it to a classmate.
- 6. Slowly turn on the magnetic stirrer (NOT the heater). Make sure the stir bar does not hit the probe while it is stirring and set the speed so a small vortex can be seen in the distilled water.
- 7. On the LABQuest, click on the graphical display. Press "Start." (Ask your instructor or your lab neighbors if you cannot find this.) Add the NaCl to the DI water in the beaker and watch the trace on the screen while the NaCl dissolves. Sketch this trace in the data section of the report sheet. Do not close the window yet and do not press any buttons on the lab quest.
- 8. This is now your **original NaCl solution**. When the conductivity becomes almost constant, record the final conductivity value in the data table as the original NaCl solution.
- 9. Click "Stop," remove the probe, and rinse it with distilled water into the waste beaker.
- 10. Remove your NaCl solution from the stirring plate, remove the spin bar with tweezers, and rinse the bar with distilled water. **SAVE your solution** for the following steps.
- 11. Pour between 20 and 25 mL of the original NaCl solution into a 50 mL graduated cylinder. Read and record this volume as the "initial" volume (V_i) of the original NaCl solution *to the nearest 0.1 mL*. Then add deionized water to the cylinder to a total volume of between 40 and 45 mL. Record the "final" volume (V_f), *to the nearest 0.1 mL*. Pour the diluted solution into a new, dry 100 mL beaker. (Why a dry beaker?)

This is now your diluted NaCl solution.

- 12. Immerse the conductivity probe in the diluted solution and record the displayed conductivity value in the data table as the diluted NaCl solution.
- 13. Discard the NaCl solutions and rinses in the sink. Rinse all of the glassware and the conductivity probe and put the equipment away. Return the magnetic stir bar to your instructor.

REPORT SHEETS Experiment 15: Ionic Solutions (Electrolyte Solutions)



Name: _____

Section:

Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

DATA PART A: IONIC SOLUTIONS

	Compound	Conductivity Values (µS) (Record data for all trials.)	Strong, Weak, or Nonelectrolyte?	Many, Few or No Ions Produced in Water?
1	NaCl Sodium Chloride			
2	CaCl ₂ Calcium Chloride			
3	HCI Hydrochloric Acid			
4	CH ₃ COOH (C ₂ H ₄ O ₂) Acetic Acid			
5	HOCH ₂ CH ₂ OH (C ₂ H ₆ O ₂) Ethylene Glycol			
6	CH ₃ OH Methanol			

DATA PART B: CONDUCTIVITY ANALYSES

For original and diluted NaCl solutions use sig figs!

Total Volume of DI Water for Original NaCl Solution, in mL	mL
Mass of Dry NaCl Used for Original NaCl Solution	g
Calculated Concentration of Original NaCl Solution (C_i)	
C _i = $\frac{\text{mass of dry NaCl (in grams)}}{\text{volume of solution (in mL)}} \times 100\%$	%(<i>m/v</i>)
Volume of Original NaCl Solution Used ("initial" volume, V _i)	mL
Volume of Diluted NaCl Solution Obtained ("final" volume, V _f)	mL
Calculated Concentration of Diluted NaCl Solution (C _f)	
Since $\mathbf{C}_f \mathbf{V}_f = \mathbf{C}_i \mathbf{V}_i \longrightarrow \mathbf{C}_f = \mathbf{C}_i \frac{\mathbf{V}_i}{\mathbf{V}_f}$	%(<i>m/v</i>)

(Record conductivity data in the table below.)

Reproduce the conductivity trace observed on the LabQuest when the NaCl was dissolved in the DI water. A rough sketch is sufficient, but **label the** *x***- and** *y***-axes correctly.**

Describe in words what this graph implies about the NaCl particles on a molecular level.

CONDUCTIVITY DATA

Sample	Concentration (% <i>m/v</i> NaCl)	Conductivity (µS/cm)
Deionized Water	0	
Original NaCl Solution (C _i)		
Diluted NaCl Solution (C _f)		

CALIBRATION PLOT

Make a graph of the concentration (*x*-axis) and conductivity (*y*-axis) for distilled water and the two NaCl solutions. On your graph, make sure to **label the axes correctly and include units**.

Complete the calibration plot by drawing a single straight line that best fits the points. (Do not connect the dots!)

NOTE: This graph should result in a straight line!

POST-LAB QUESTIONS

- 1. HCl is a covalent (molecular) compound in the gas phase. Does your data indicate that HCl behaves as molecules or ions when dissolved in water? Explain and include data to support your conclusion.
- 2. Write a chemical equation that communicates what solid $CaCl_2$ forms when dissolved in water. Note that water is not a reactant in this process. Use physical states in your equation as appropriate: (*s*), (*I*), (*aq*).

3. The following are beakers of water. Water molecules are already drawn in the beaker. Fill in the ions or molecules present when each of the following substances is dissolved in water. Use spheres to represent atoms/ions/molecules, and include a legend or labels with their chemical formula. (A similar example in the Introduction shows chemical formulas instead of spheres.) Use at least 4–5 spheres (molecules or ions) for each drawing.



- 4. Calculate the concentration of a 2.4-L solution that has 1.5 g of NaCl dissolved in it. Using your graph, estimate the conductivity.
 - a. Calculated % *m*/*v* concentration: ______ Include units!
 - b. Estimated conductivity: _____ Include units!
 - c. Describe how you obtained the result.
- 5. a. If you measured the conductivity of pond water and found it was 2000 µS/cm, what % *m/v* concentration of NaCl would you expect for the pond water?
 - b. Estimated concentration based on your graph: _____ Include units!
 - c. Would the pond water sample above (2000 μS/cm) taste "salty", assuming the conductivity of the solution was due to salt content? (What's salty? Let's say a solution of 2.5g NaCl in 250 mL water is considered "salty.")

State "yes" or "no," and **EXPLAIN**. Show your work if it involves a calculation.
PRE-LAB Experiment 15: Ionic Solutions (Electrolyte Solutions)

Name:

Section: _____

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. Define the following terms (in your own words). Make sure you cite any sources used (provide author and title and page number or website).
 - a. nonelectrolyte
 - b. strong electrolyte
 - c. weak electrolyte
- 2. Predict whether each of the following is ionic or covalent (molecular). Circle your answers.

a.	Water, H ₂ O	ionic compound	covalent compound
b.	Sodium chloride, NaCl	ionic compound	covalent compound
C.	Calcium carbonate, CaCO ₃	ionic compound	covalent compound
d.	Hydrogen chloride, HCl	ionic compound	covalent compound
e.	Glycerol, $C_3H_8O_3$	ionic compound	covalent compound

- 3. Nitrate is a polyatomic ion with a charge of 1–. Its ionic formula is NO₃⁻. Strontium is a Group 2 atom that forms a cation with a charge of 2+. Its ionic formula is Sr²⁺.
 - a. Write the correct chemical formula for the ionic compound strontium nitrate.
 - b. The dissociation of sodium carbonate is written as this equation (see Introduction):

 $Na_2CO_3(s) \rightarrow 2 Na^+(aq) + CO_3^{2-}(aq)$

Write a chemical equation for the dissociation of strontium nitrate in water. (**HINT**: *Water is not a reactant but you can write it over the chemical arrow.*)

- 4. Calcium chloride is CaCl₂. Which equation best describes calcium chloride when it dissociates?
 - a. $\operatorname{CaCl}_{2}(s) \rightarrow \operatorname{Ca}^{2+}(aq) + \operatorname{Cl}_{2}^{2-}(aq)$ c. $\operatorname{CaCl}_{2}(s) \rightarrow \operatorname{Ca}^{2+}(aq) + 2\operatorname{Cl}^{-}(aq)$
 - b. $\operatorname{CaCl}_2(s) \to \operatorname{Ca}^{2+}(aq) + \operatorname{Cl}_2^{-}(aq)$

CHEM& 121 WORKBOOK

EXPERIMENT 16 ACIDS, BASES, AND BUFFERS



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INTRODUCTION

The pH scale is a measure of the relative acidity (or basicity) of a solution.

pH is defined as – $log[H^+]$ where $[H^+]$ is the molar concentration (or molarity) of the hydrogen ion, H^+ .

The following scale shows the relationship between pH and $[H^+]$:



NOTE: The lower the pH, the higher the [H⁺], and the higher the pH, the lower the [H⁺]. Solutions of low pH are more acidic and those of high pH are more basic.

Litmus and red cabbage juice can give some idea of the pH value. Red litmus paper turns blue in basic solutions (think: "Bases turn blue") and blue litmus paper turns red in acidic solutions.

Red cabbage juice contains pigments which change color at various pHs.

These colors and approximate pH values follow:

Approximate pH	2	4	6	8	10	12
Color of Extract	red	purple	violet	blue	blue-green	green

In this experiment, you will use litmus and red cabbage juice to determine the pH values of household substances over the entire pH range. Then you will learn how to measure the pH more quantitatively using a pH probe, and investigate the nature of buffer solutions.

OBJECTIVES

In this experiment, you will

- Determine the pH values of household substances using two different methods.
- Add acid and base to a buffer solution and determine the effect on the pH of the solution as an exercise in scientific inquiry.

HAZARDS

The ammonia solution is toxic. Its liquid and vapor are extremely irritating, especially to eyes. Drain cleaner solution is corrosive. Handle these solutions with care. Do not allow any of the solutions to contact your skin or clothing. Wear goggles at all times, as any of these substances can irritate your eyes. Notify your instructor immediately in the event of an accident. *Do not eat or drink in the laboratory.*

PROCEDURE

PART A: pH AND COLOR

- 1. Obtain and wear goggles.
- 2. Label 8 test tubes with the numbers 1–8.
- 3. Before you move on, make sure you have a prediction for the acidity of each solution. Fill these into your data table.
- 4. Use a Pasteur pipette to transfer a sample of each substance in the data table (1–7) into each labeled test tube. The volumes are not precise—you just need about a pipette full (squeeze the bulb, draw up liquid, and expel it into the test tube. It will be about 1–2 mL in volume.
- 5. In vial 8 prepare your sample that you brought from home. If your sample is a solid, dissolve it in 4 mL of deionized (DI) water. If your sample is a liquid, it may be used as is.

LITMUS TESTS

- 6. Use a glass stirring rod to transfer one drop of each sample onto a small piece of blue litmus paper on a paper towel. Transfer one drop to a piece of red litmus paper on a paper towel. Record the results. Set up a "waste" beaker for collecting the rinse solutions. Rinse the stirring rod with distilled water and dry the stirring rod each time.
- 7. **NOTE:** You may use one piece of litmus paper multiple times—at least three samples may be tested using one litmus paper. Be sure to rinse and dry the stirring rod each time.

RED CABBAGE JUICE INDICATOR

- 8. *After* you have finished the litmus tests, add a Pasteur pipette full of red cabbage juice indicator to each of the eight samples and mix using a glass rod. Record your observations. Place these solutions in the waste container. (Do not put paper towels or litmus paper into the waste container.)
- 9. Using these data, determine whether each sample is acidic, basic, or neutral and enter the information on your Report Sheet.

PART B: BUFFERS

pH TESTS

For this part of the lab, you will need to measure pH more precisely, using a pH sensor. The sensors are expensive, so please be careful. The bulb on the end is particularly fragile and prone to break so please be gentle with the equipment!

TO SET UP A pH SENSOR

- Connect the pH Sensor to the LabQuest and DO NOT set the pH readings to zero. You should see a pH value (that is non-zero) displayed on the screen. If you have any questions, ask your instructor.
- 2. Obtain about 20 mL of distilled water in a clean 50-mL beaker.
- Use a wash bottle filled with deionized water to thoroughly rinse the tip of the sensor as demonstrated by your instructor. Catch the rinse water in a 250-mL beaker labeled "waste."
- 4. Use the sensor to swirl the solution gently. When the pH reading stabilizes, record the pH value in your data table.
- 5. Prepare the pH Sensor for reuse by rinsing it with deionized water from a wash bottle (rinsings go into the waste beaker).

- 6. Add a drop of 0.10 M HCl to the 50 mL beaker containing sample. Swirl the solution gently and record the pH when it stabilizes. Repeat for each drop of HCl added until you have completed the data table.
- 7. Repeat Steps 4–6 but use 20 mL of a pH 4 or pH 10 buffer solution in a clean 50-mL beaker.
- 8. Repeat Steps 4–6 but use 20 mL of the same buffer solution and use 0.10 M NaOH in Step 6 instead of HCl.
- 9. When you are done, rinse the tip of the sensor with deionized water and return it to the sensor soaking solution. **Place the probe back in the box in an upright position.**

Pour all liquid waste into the labeled waste container. (Do not put paper towels or litmus paper into the waste container.)

TROUBLESHOOTING: WHAT IF I HAVE TROUBLE READING pH VALUES?

My pH probe seems to read strange values!

In case you accidentally zero your pH probe, you will need to perform a calibration. Follow these steps and/or ask your instructor for guidance:

- a. Obtain a pH 4 solution. Place the pH probe in the solution and tap on the red box on the display and choose "calibrate." When the voltage reading stabilizes, enter "4" for pH and click "keep."
- b. Rinse the pH probe. Obtain a pH 7 solution. When the voltage reading stabilizes, enter "7" for pH and click "keep."

NOTE: If you plan to take measurements at the upper end of the pH scale, then use pH 10 buffer instead of pH 4.

REPORT SHEETS Experiment 16: Acids, Bases, and Buffers



Name: _____

Section: _____

Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

DATA

PART A

Test Tube	Solution	Prediction: Acid, Base, Neutral?	Blue Litmus (color)	Red Litmus (color)	Red Cabbage Juice (color)	Acid, Base, or Neutral?
1	Vinegar					
2	Ammonia					
3	Lemon Juice					
4	Soft Drink					
5	Drain Cleaner					
6	Detergent					
7	Baking Soda					
8						

PART B

		1	2	3	4	5
Total Numl	ber of Drops of HCI (<i>aq</i>) Added					
nH	DI Water					
рп	Buffered Solution					
Total Number of Drops of NaOH (<i>aq</i>) Added						
рН	DI Water					
	Buffered Solution					

POST-LAB QUESTIONS

PART A

- 1. Which of the household solutions tested are acids? Based on your tests, what do they have in common?
- 2. Which of the solutions are bases? Based on your tests, what do they have in common?
- 3. What color(s) is red cabbage juice indicator in acids? In bases?
- 4. State at least one advantage of using red cabbage juice over litmus paper.

PART B

- 1. Briefly explain any differences in pH changes observed for DI water compared to a buffered solution when HCI (*aq*) was added to each.
- 2. Briefly explain any differences in pH changes observed for DI water compared to a buffered solution when NaOH (*aq*) was added to each.
- 3. Did you use a pH 4 or pH 10 buffer? _____ Find someone in the lab who used a different buffer from you. Do you see similarities in how their buffer behaved compared to yours? Explain.

- 4. Based on your results, which statement is most accurate about buffer solutions?
 - a. A buffer solution is a solution in which the pH remains constant when base or acid is added.
 - b. A buffer solution is a solution which resists significant changes in pH when a base or acid is added.
 - c. A buffer is a neutral solution at pH 7 after a base or acid is added.
 - d. A buffer solution can neutralize an infinite amount of acid or base added to it.
- 5. The pH of our blood buffer system (bicarbonate) is 7.4. What do you think will happen to the pH of blood if one or two drops of 0.10 M HCl is added to it?

- 6. Consider a buffer solution that consists of the blood buffer system, H_2CO_3 / HCO_3^{-} .
 - a. Write the reaction that occurs when a strong acid (H⁺ or H_3O^+) is added to the buffered solution.

b. Write the reaction that occurs when a strong base (OH^{-1}) is added to the buffered solution.

PRE-LAB Experiment 16: Acids, Bases, and Buffers

Name: _____

Section: _____

Complete the following questions **BEFORE** class. Refer to your textbook or the web if needed.

- 1. Of the 7 compounds you will be testing in this lab (see Part A data table), which do you predict are
 - a. acids?
 - b. bases?
- 2. A solution has a pH of 9. Is the solution acidic, basic, or neutral?
- 3. The pH sensor will be stored in water. Why is it inappropriate to zero the pH sensor in this solution?
- 4. Refer to the Acids and Bases Chapter (buffer section) of your textbook or do an internet search to answer these questions:
 - a. What is the normal pH range of blood? _____
 - b. At what minimum and maximum pH levels do cells start to die?
 - c. What is the buffer system that regulates blood pH? (weak acid = ?; anion = ?)
 - d. What do the terms acidosis and alkalosis mean with respect to blood pH?
- 5. If you would like to test a household solution, please bring at least 1 sample of a colorless liquid (10 mL) and/or solids (size of a grape). Check to make sure that the substances will mix with water. Do *not* bring flammables or alcoholic beverages. Juices, vinegar, household cleaners, detergents, or aspirin are some possibilities. You will test your sample in this lab.

CHEM& 121 WORKBOOK

EXPERIMENT 17 CHEMICAL DETECTIVE

You have now learned a great deal about chemicals and how they behave. To use this knowledge, we would like you to take on the role of a chemical detective. This is very similar to what happens in crime labs around the world. You will have several tests that you can perform to determine what chemicals are present in your sample. You will have a chance to test all the possible suspect chemicals, then you will decide what procedure is best to determine the identity of the compounds in your unknown sample vial.

Here is a list of the "suspect" chemicals:



The testing materials you will have to work with follow:

A conductivity meter

- Iodine solution
- Litmus paper (both blue and red)
- Benedict's Reagent

Vinegar

Each test will give a definite response to each of the suspect chemicals. You will need to make good observations to know what to look for in each test.

General Procedures: These tests can be done in any order.

HAZARDS

- This experiment involves many reagents. Wear goggles until all chemicals are put away. lodine and Benedict's reagent are corrosive—avoid skin contact and ingestion.
- When heating test tubes of solutions, beware for superheating—if it bubbles over, it can cause burns.

PROCEDURE

You will need to test the reactions of each known suspect chemical with the testing materials. You can simultaneously analyze *two* unknown vials. **There are two suspect chemicals in each unknown vial.** The following are general notes on how each process is performed. You will probably want to do the same test to all five known chemicals at one time to see how they are different.

CONDUCTIVITY METER TEST

Take a small amount of your sample, about the size of a match head, and place it in a 50 mL beaker. Add about 30 mL of distilled water and stir until it is dissolved. Place both electrodes into the beaker until they are covered with the liquid by about a centimeter. Push and hold the small button on the meter to turn it on. Test first on low and if you get a reading of 10, then test on high. Think of low as a scale from 1–10 and high as a scale from 11–20.

LITMUS PAPER TEST

Litmus paper comes in two colors, red and blue. It is used for detecting acids and bases. (Hint to remember colors: Bases Become Blue). If the red paper turns blue, the solution is basic. If the blue paper turns red, the solution is acid. If neither paper changes color, the solution is neutral. To test this, dissolve a small amount of your sample in water, or use your sample from the conductivity test, and dip the litmus paper into the solution. The change should be obvious. Be sure to recap the container the paper is stored in to prevent reaction with the air.

VINEGAR TEST

Take a small amount of your sample, about the size of a match head, and place it on a spot plate. Add a few drops of vinegar to it from the dropper bottle. Observe any reaction.

IODINE SOLUTION TEST

The iodine solution tests for large carbohydrates. Take a small amount of your sample, about the size of a match head, and place it on a spot plate. Add a few drops of iodine solution to it from the dropper bottle. Observe any reaction.

BENEDICT'S REAGENT

Benedict's Reagent is used to identify small carbohydrates. Set up a hot water bath with a hot plate and beaker filled with water. Take a small amount of your sample, about the size of a match head, and place it in a test tube. Add 1 mL (one full dropper) of Benedict's Reagent to the test tube. Heat the mixture for 2–3 minutes. Observe any color changes. You may want to heat all of the samples at the same time. Be sure each test tube is labeled. **(NOTE: All samples with Benedict's solution MUST be placed in a waste jar.)**

Obtain two unknown vials and analyze them using the tests described above. EACH VIAL CONTAINS TWO SUSPECT CHEMICALS.

REPORT SHEETS Experiment 17: Chemical Detective

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Name: ______ Section: _____ Lab Partner: _____

Turn in these Report Sheets only as your lab report. You do not need to include the previous pages.

Prepare a data chart for your experiment.

CONCLUSIONS

Vial#	contains	and	
Vial#	contains	and _	



















POLYATOMIC IONS

ammonium	NH_4^+	sulfite	\$03 ²⁻
hydronium	H ₃ 0+	hydrogen sulfate (bisulfate)	HSO ₄ ⁻
hydroxide	OH-	hydrogen sulfite (bisulfite)	HSO3 ⁻
carbonate	CO3 ²⁻	perchlorate	CI04 ⁻
hydrogen carbonate (bicarbonate)	HCO3 ⁻	chlorate	CI0 ₃ ⁻
nitrate	NO_3^-	chlorite	CI02 ⁻
nitrite	N0 ₂ ⁻	hypochlorite	C10 ⁻
phosphate	P04 ³⁻	cyanide	CN ⁻
hydrogen phosphate	HPO4 ²⁻	permanganate	Mn0 ₄ ⁻
dihydrogen phosphate	H ₂ PO ₄ ⁻	thiosulfate	S ₂ O ₃ ²⁻
chromate	Cr04 ²⁻	oxalate	$C_2 O_4^{2-}$
dichromate	Cr ₂ 07 ²⁻	borate	B03 ³⁻
acetate	C ₂ H ₃ O ₂ ⁻ or CH ₃ COO ⁻	citrate	C ₆ H ₅ O ₇ ³⁻ or C ₃ H ₅ O(COO ⁻) ₃
sulfate	S04 ²⁻		

ELECTRONEGATIVITY VALUES

NUMBERS AND FORMULAS TO KNOW

