Workbook/Note-Guide
to accompany
College Chemistry I
Videos

Sandra Y. Etheridge, Ph.D.
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Introduction

College Chemistry I is a course that covers the topics addressed in most first semester college chemistry courses. Many programs of study, particularly certain engineering degrees, require one semester of college chemistry as opposed to a two semester course, hence the year long course has been split into two separate courses to accommodate those needs. The second semester course is available under the title, College Chemistry II.

The videos used for this course were made in the studios at Gulf Coast Community College by Dr. Sandra Etheridge and were designed to meet the needs of students taking chemistry by distance education. The course is referred to on the videos as CHM 1045 which is in accordance with the common course numbering system for Florida Universities and designates the course as being the college or university chemistry lecture course for majors in the field. The DVDs are copies of those distance education videos and are made available courtesy of Gulf Coast Community College.

The workbook/note guide written by Dr. Etheridge is designed to facilitate organized and complete note-taking to be used for study. The workbook matches the videos exactly since Dr. Etheridge wrote both the course and the workbook/note guide. If a mistake is made on the video (there are several), the error and correction are pointed out in the note guide. Further, tables and charts needed at certain points in the lecture are provided at those points in the note guide. Notations of appropriate points at which to pause the video for problem solving appear in the note guide, and utilizing those pauses as suggested facilitates learning. For those reasons it has been found that students benefit significantly from properly using the note guide. It is such an effective aid that students enrolling in the lecture course at the college frequently purchase the note guide to assist their note taking and learning.

The reference made to a syllabus during the first lesson refers to the syllabus for whatever college/university the student may be attending. It is always a good idea to have a syllabus for the course for which credit is sought. The study guide refers to a publication matching the textbook being used and published by that same company. Although a study guide can be handy, providing additional problem solving information, sample tests, and internet references, it is not critical to success in the course unless the instructor of record requires it.

It is suggested that individuals viewing these videos have had either a high school or college preparatory course in chemistry, as well as some background in algebra. Although it is possible to deal with the material without the chemistry or algebra background, it makes it more difficult for the student. It is also suggested that the viewer have access to a college chemistry textbook since that reference could be used to answer questions that might arise, as well as provide additional problems for practice and the very large reference tables needed for solving problems.

Hopefully you will enjoy these videos and note guide as much as Dr. Etheridge enjoyed making them for her students.
Points of clarification:

**Workbook/Note-guide** – this refers to this body of information written by the Dr. Etheridge, the professor on the video, and is designed to match the video you will be watching.

**Study Guide** – this is a publication usually provided by the company publishing the textbook you are using.

**Syllabus** – this document provides specific information regarding the course: credit hours, pre-requisites/co-requisites, student performance objectives or student outcomes, instructor information, dates/times of exams, etc. It is specific to the institution you may be attending. If you are enrolled in a chemistry course, you should have a syllabus for the course. Your chemistry department should be able to provide this for you.

This video course may be used with virtually any textbook you may be using and it is important that you have a textbook for this course. If you are using the videos and workbook for review or prep purposes and plan to acquire a textbook, any college level first year textbook should be fine. It should be a two semester text, not introductory, and be published no earlier than 2000 since terminology tends to change with passing time.

If your textbook has a study guide, that study guide could be quite helpful in that it may provide additional information, sample test questions, and suggestions for additional reading. However, many students progress well without having the accompanying study guide.

This **workbook/note-guide** will provide a detailed fit for the video you will be watching. The problems and questions will appear in this workbook, just as they appear on the video and ample room is left for your work and appropriate notes. This document guides your taking of notes, working of problems, and answering questions posed on the videos. Further, it provides the supplementary materials to match materials utilized in the videotapes/DVD’s, such as a table of oxidation numbers, activity series, etc.

There will be certain information you will need to study or prepare prior to beginning related sections of the course. The **workbook/note-guide** will direct you to those items of information or direct your preparative study, and more importantly will provide a guide for your note-taking.

All college chemistry courses are not organized in the same manner. The American Chemical Society provided recommended content information to members regarding these undergraduate courses, but did not specify the order in which topics were to be addressed. Although Professor Etheridge has included the topics she felt most critical to her students in their first semester of chemistry, she does not guarantee that her choice of topics will be the same as those of the professor directing chemistry study at the institution you may be attending. However, another course follows this one, College Chemistry II, and it may contain any additional topics you need covered.
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Appendix
During this unit, Dr. Etheridge will

1. discuss common terms such as substances, mixtures, intensive, and extensive properties. Except for those occasions in which the concept is noted as being for interest only, everything mentioned on the video or noted in this workbook to be important to your learning.

2. briefly go through the Metric system. Change among the various prefixes such as milli, centi, kilo, etc will be demonstrated, as well as the Angstrom, micro, and nano units. You should commit these prefixes and units to memory.

3. demonstrate conversions from English to Metric and Metric to English, including the temperature scales: Kelvin, Fahrenheit, and Celsius temperatures. Three conversions students should commit to memory are

   \[
   \begin{align*}
   2.54 \text{ cm} &= 1.00 \text{ in} \\
   946.3 \text{ mL} &= 1.00 \text{ qt} \\
   453.6 \text{ g} &= 1.00 \text{ lb}
   \end{align*}
   \]

   Many students are not familiar with the English system and should look in a good cookbook to learn the number of teaspoons in a tablespoon, tablespoons in a cup, cups in a quart, and quarts in a gallon. Additionally, students should learn there are 8 fluid ounces in a cup and 16 ounces (avoirdupois) in a pound (a volume ounce and a weight ounce are not the same), 12 inches in a foot, 3 feet in a yard, and 5280 feet in a mile.

4. discuss the difference between temperature and heat, and the units for each.

5. use scientific notation, but will not dwell on how to convert to and from scientific notation. Students are expected to understand scientific notation as a tool and utilize it on their calculators.

6. discuss density and specific gravity. Unless otherwise indicated, she will use as standards for specific gravity:

   - the density of water = 1.000 g/mL as the standard for solids and liquids
   - the density of air = 1.292 g/L at STP as the standard for gases

7. briefly cover the use of significant figures (sometimes referred to as significant digits), precision, and accuracy, and expects students to be proficient with percentages.

8. briefly discuss mass, force, weight, and their proper units.

9. address density and specific gravity for all three states of matter.
During this unit, we will discuss the following:

The Metric/English Systems

Temperature and Heat

Mass and Weight

Density and Specific Gravity

Lesson 1:

MEASUREMENTS: THE ENGLISH AND METRIC SYSTEMS

Fill this in before or during the course of the lecture

**The English System.**

How many

- inches in a foot? ______
- feet in a yard? ______
- feet in a mile? ______
- teaspoons in a tablespoon? ______
- tablespoons in a cup? ______
- cups in a pint? ______
- cups in a quart? ______
- quarts in a gallon? ______
The **Metric System**. (you should fill this in either during the video or prior to watching the video). The prefixes listed in **bold** should be committed to memory.

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Value</th>
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<tbody>
<tr>
<td>giga</td>
<td></td>
<td></td>
</tr>
<tr>
<td>mega</td>
<td></td>
<td></td>
</tr>
<tr>
<td>kilo</td>
<td></td>
<td></td>
</tr>
<tr>
<td>deci</td>
<td></td>
<td></td>
</tr>
<tr>
<td>centi</td>
<td></td>
<td></td>
</tr>
<tr>
<td>milli</td>
<td></td>
<td></td>
</tr>
<tr>
<td>micro</td>
<td></td>
<td></td>
</tr>
<tr>
<td>nano</td>
<td></td>
<td></td>
</tr>
<tr>
<td>pico</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Do you know what “SI” stands for? If not, look it up.

Metric and English Systems:

1.0000 inch = ______ cm

1.000 quart = ________ mL

1.000 pound (lb) = ________g
Calculate the volume of 1.8 ft³ in m³.

**solve:**

**TEMPERATURE AND HEAT:**

**Temperature**

What is temperature?

<table>
<thead>
<tr>
<th>Water Boils</th>
<th>Water Freezes</th>
</tr>
</thead>
<tbody>
<tr>
<td>°F</td>
<td></td>
</tr>
<tr>
<td>°C</td>
<td></td>
</tr>
<tr>
<td>K</td>
<td></td>
</tr>
</tbody>
</table>

What formula is used to convert:

- °C to °F
- °F to °C
- °C to K
Do you know?? What is significant about zero (0) on the

Fahrenheit scale

Celsius scale

Kelvin scale

Lesson 2:

Heat

Units of heat

Joule –

Calorie –

BTU -

Which way does heat flow?

Weight

What is weight?

What are the units of weight?
Mass

Define mass –

What is inertia?

REM: MASS IS CONSTANT!

The units of mass are:

Is weight constant? _______

Is mass constant? _______

Why?

DENSITY AND SPECIFIC GRAVITY

Density

Define “density”

The formula for finding density is
Remember this…
For a solid and/or liquid, the mass is measured in grams and the volume is measured in milliliters or cm³; but for a gas, the mass is measured in grams and the volume is measured in liters. The density of a gas when expressed in grams per milliliter is such a small value that it becomes awkward. It is more reasonable to measure it in terms of grams per liter.

There are two densities you are expected to know:

density of water at 4°C _____________

density of air at STP _____________

You have a 25.0 g piece of iron displacing 6.2 mL of water. Find the density of iron.
solve:

BE VERY CAREFUL WITH YOUR SIGNIFICANT FIGURES OR DIGITS.

Calculate the mass of 2.00 teaspoons (tsp) of water.
solve:

Specific Gravity

What is specific gravity?

What is the standard for solids/liquids? ____________________
for gases? ____________________
*If the density of a liquid alcohol is 0.80 g/mL, what is the specific gravity of the alcohol?*

What are the units for specific gravity?

*If the specific gravity of an iron bar is 7.88, what is the density of the iron bar?*

If the specific gravity of mercury is 13.6, what volume would 100.0 g Hg occupy?

**solve:**

If the specific gravity of carbon dioxide gas is 1.52, what volume would 100.0 g of the gas occupy?

**solve:**
Quick Quiz 1

1. \(2.54 \times 10^{-3} \text{ m}\) could be replaced with
   a. 2.54 cm     b. 2.54 mm     c. 2.54 dm     d. 2.54 \(\mu\text{m}\)     e. 2.54 km

2. When 66.0 °C is converted to Fahrenheit, the temperature is
   a. 18.9 °F     b. 46.7°F     c. 86.8 °F     d. 150.8 °F     e. 200.9 °F

3. An alcohol and an oil were carefully placed in a beaker containing water. A wooden toothpick was dropped in. The alcohol floated on the oil, and the toothpick came to rest on the water, below the oil. The substance with the highest specific gravity is the
   a. oil     b. water     c. alcohol     d. toothpick     e. insufficient information

4. If the distance between two nuclei is 206 pm, what would this distance be in meters?
   a. 2.06 \(\times 10^{-10}\)     b. 2.06 \(\times 10^{-12}\)     c. 2.06 \(\times 10^{-14}\)     d. 2.06 \(\times 10^{-15}\)     e. 2.06 \(\times 10^{-16}\)

5. If the density of alcohol is 0.782 g/mL, what volume would be occupied by 43.6 g?
   (Express your answer to the proper number of significant digits.)
   a. 34.09 mL     b. 34.1 mL     c. 55.75 mL     d. 55.8 mL     e. none of these

6. When 16.42 ft/s is expressed as km/min, the value should be
   a. 0.3003     b. 4.65 \(\times 10^{-2}\)     c. 21.5     d. 208.53     e. 898

7. Four old dimes were analyzed for silver content and the following data recorded. The correct silver content was 93.8%. Therefore, the data is best described as
   
<table>
<thead>
<tr>
<th>Analysis</th>
<th>% silver</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>89.7</td>
</tr>
<tr>
<td>2</td>
<td>96.6</td>
</tr>
<tr>
<td>3</td>
<td>99.3</td>
</tr>
<tr>
<td>4</td>
<td>89.4</td>
</tr>
</tbody>
</table>

   a. accurate and precise     b. accurate, but not precise
   c. inaccurate, but precise     d. neither accurate nor precise
During this unit, Dr. Etheridge will

1. suggest that you learn the names and symbols for the first 109 elements.

2. demonstrate the use of standard notation for indicating both atomic number and atomic mass number for atoms and ions.

3. discuss relative weights and give a brief history of relative weights.

4. show how to calculate the weighted average of an element’s atomic mass when given the individual atomic masses and the percent occurrence of naturally occurring isotopes.

5. expect students to calculate the percent abundance of specific isotopes when given masses of those isotopes or calculate the isotopic masses when given percent abundance of the isotopes.

6. give a brief description of how a basic mass spectrometer works.

7. give a general overview of the periodic table and point out where types of elements and families/groups of elements are found.

8. introduce the concept of the mole roadmap and demonstrate how to use this concept in chemical calculations.

9. go through the expanded rules of nomenclature.

10. discuss the chemical make up of selected common, household chemicals.

11. mention terms such as atoms, molecules, element, compound, mixture, isotope, ion, cation, anion, and allotropes which you are expected to know.

12. give very general introduction to atomic structure, reserving much of this topic for discussion at a later point in the course.

13. present a general and limited discussion of the history of atomic theory by introducing the work of Thomson, Millikan, Rutherford, and Chadwick.
In this unit we will discuss

- The Nuclear Atom (Symbols of Elements and Isotopic Notation)
- The Periodic Table (briefly)
- The Mole
- Nomenclature (Naming of compounds and writing of formulas)

**Lesson 3:**

**SYMBOLS OF ELEMENTS AND ISOTOPIC NOTATION**

**Symbols of Elements**

Symbols of elements may contain one or two letters. If the symbol is a single letter, i.e. N, it is ALWAYS capitalized. If the symbol contains two letters, the first is always capitalized and the second is ALWAYS lower case. Please note: a lower case letter is not a shrunken capital letter.

Give the isotopic notation for Na-23.

The atomic number of this isotope is _______

What does the atomic number tell us?

The atomic mass number of this isotope is _______

What does the atomic mass number tell us?

How many protons does it have? _______, neutrons? ________

What is the weight of sodium as recorded on the periodic table? ________________
Relative Weights

Why do we need “relative weights”?

What substance is the present standard for atomic weights? _____________

What is the atomic weight of this standard? _________________

It weighs that amount because ……

The units of these weights may be in amu (atomic mass units) or daltons, although no units are needed. Remember, they are RELATIVE weights.

Can an element have more than one atomic weight? Why?

What are isotopes?

How are they alike?

How do they differ?

Remember: When we are finding the weight used on the periodic table, we use only naturally occurring isotopes.
**Mass Spectrometer (Mass Spec)**

What is this device used for?

Sketch the device and label its parts:

What occurs in the oven?

What occurs in the ionizing chamber? Give a sample equation:

What is the purpose of the magnetic field?

What is meant by the magnetic field strength ratio?

Give the equation for finding isotopic atomic weight:
If the magnetic field strength ratio of Mg-24 is 1.41377, what is the relative weight of that isotope?

solve:

<table>
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<th>isotope</th>
<th>mag. field strength ratio</th>
<th>percent occurrence*</th>
<th>relative weight</th>
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<tbody>
<tr>
<td>$^{24}\text{Mg}$</td>
<td>1.41377</td>
<td>78.792</td>
<td></td>
</tr>
<tr>
<td>$^{25}\text{Mg}$</td>
<td>1.44297</td>
<td>10.148</td>
<td></td>
</tr>
<tr>
<td>$^{26}\text{Mg}$</td>
<td>1.47147</td>
<td>11.060</td>
<td></td>
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</tbody>
</table>

*these values are rounded off, thus the results may not be exactly that which appears on the periodic table.

Calculate the relative weights of each of the isotopes in the space below and place them in the column above.

$^{24}\text{Mg} =$

$^{25}\text{Mg} =$

$^{26}\text{Mg} =$

Find the weight that should be listed on the periodic table. (Pause the video.)
THE PERIODIC TABLE

Families or Groups

(Please refer to the periodic table which appears on the next page.)

There are three ways to number the families or groups. The notation we will use is the IA, IIA, IIB, IVB, VB, VIB, VIIB, VIII, IB, IIB, IIIA, IVA, VA, VIA, VIIA, and 0 or Noble Gas.

How are elements in a family/group alike?

The common names of the families/group are:

_____________________IA
_____________________IIA
_____________________VIIA
_____________________Noble (Inert) Gases

Other Family Names: Here an error is made as Dr. Etheridge misspoke herself: It is the manganese family, not the iron family. Then comes the iron group.

LOOK THIS UP: As the atomic number of elements increase, do atomic weights always increase? Cite examples to prove your answer.

Periods

There are _______ periods on the long form of the periodic table.

Later, we will discuss what these periods represent.
**PERIODIC TABLE OF THE ELEMENTS**

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<tr>
<th>1A</th>
<th></th>
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<tr>
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<td>H</td>
<td>10079</td>
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<tr>
<td>2</td>
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</tr>
<tr>
<td>3</td>
<td>Na</td>
<td>22.989</td>
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<td>4</td>
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<td>5</td>
<td>Rb</td>
<td>85.468</td>
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<tr>
<td>6</td>
<td>Cs</td>
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<tr>
<td>7</td>
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<table>
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**LANTHANIDES**

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**ACTINIDES**

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<td>259</td>
<td>201</td>
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</table>

**Notes:**
- Elements in bold are solid at room temperature.
- Elements in italics are gases at room temperature.
Lesson 4:

Properties of elements in a period:

Properties of elements in the same period gradually change across the period. The change in properties from one element to another depends on where the elements are located. Describe the extent to which properties of elements change as one moves from left to right across the periodic table. Compare this extent of change between the IA and IIA elements and between the IIA and IIIA elements? Describe the changes as we move from one transition element to another.

Metals/Non-Metals (A periodic table appears on the next page for your convenience and note-taking.)

List the properties of metals.

List the properties of non-metals.

Describe where metals are located on the periodic table.

Describe where the non-metals are located?

What terms are used to describe the elements that lie on the stair-step line separating metals from non-metals?

Describe properties of these elements:
# Periodic Table of the Elements

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<tr>
<th>Period</th>
<th>Group</th>
<th>Element</th>
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<th>Symbol</th>
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**Lanthanides**
- Ce (58): 140.12
- Pr (59): 140.91
- Nd (60): 144.24
- Pm (61): 145.00
- Sm (62): 150.36
- Eu (63): 151.97
- Gd (64): 157.25
- Tb (65): 158.92
- Dy (66): 162.50
- Ho (67): 164.93
- Er (68): 167.26
- Tm (69): 168.93
- Yb (70): 173.04
- Lu (71): 174.97

**Actinides**
- Th (90): 232.04
- Pa (91): 231.04
- U (92): 238.03
- Np (93): 237.05
- Pu (94): 244.09
- Am (95): 243.07
- Cm (96): 247.08
- Bk (97): 247.08
- Cf (98): 251.03
- Es (99): 252.07
- Fm (100): 257.04
- Md (101): 258.04
- No (102): 259.04
- Lr (103): 260.00
THE MOLE

This is probably the single most important concept you will study in chemistry. It is important that you have a thorough, working knowledge of the **mole**.

Define the mole:

as a number:

as a mass:

In terms of atomic weights, one average atom of hydrogen weighs ____________. The units are ____________ or ____________.

One “average” atom of sodium weighs ________________.

6.022 x 10^{23} atoms of sodium (or 1.0 mole) weighs ________________ grams.

One atom of iron weighs ____________.

6.022 x 10^{23} atoms of iron weigh ________________ grams.

What is the mass of 6.022 x 10^{23} carbon dioxide molecules? ________________

(OOPS, Dr. Etheridge goofed on the video. Please make the appropriate correction!)

1.0 mole of Neon contains ________________ atoms and weighs ________________
The Mole Roadmap:

<table>
<thead>
<tr>
<th>MOLE</th>
<th>gram atomic weight or gram molecular weight</th>
</tr>
</thead>
<tbody>
<tr>
<td>6.022 x 10^{23} things</td>
<td></td>
</tr>
</tbody>
</table>

1 atom of carbon weighs ________________

1 atom of hydrogen weighs ________________

**The point is:** \( 6.022 \times 10^{23} \text{ amu} = 1.000 \text{ gram} \)

Therefore: \( 6.022 \times 10^{23} \text{ H atoms weigh } 6.022 \times 10^{23} \text{ amu or daltons which is equivalent to } 1.000 \text{g} \)

(In truth, I’m not sure that is a legitimate way to make that point, but it seems to work well for students)

Determine the number of moles in 3.82 g Copper.

solve:

Calculate the number of moles in \( 2.9 \times 10^{21} \text{ atoms of lithium} \)

solve:
Dr. Etheridge made a calculation error. The answer is $4.04 \times 10^{22}$ atoms of Cu.

solve:

Lesson 5:

Calculate the number of atoms of any kind in 10.3 g water.

solve:

Calculate the number of atoms of carbon in 3.8 tsp of C$_2$H$_6$O (an alcohol) having a density of 0.79 g/mL.

solve:
The following is a wonderful type problem for a test because it covers so many areas we have studied at this point, all wrapped into a single multiple-choice question.

Which of the following contains the greatest number of atoms of any kind?

a. 2.00 tsp water
b. 150 mg lead
c. 1.2 L helium gas having a specific gravity of 0.139.
d. 23.6 mL alcohol (C\textsubscript{2}H\textsubscript{6}O) having a specific gravity of 0.797.

(The set-up and answer appear on the next page, but you should pause the video and try your hand at the problem, here.)

solve:
This is the set-up and solution for the problem on the previous page

<p>| | | | | | | |</p>
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<th></th>
<th></th>
<th></th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>a) 2.0 tsp water x 1 tbsp/3 tsp x 1 cup/16 tbsp x 1 qt/4 cups x 946.3 mL/qt x 1.00 g/mL x 1 mole/18.0 g x 6.022 x 10^23 molecules/mole x 3 atoms/molecule = 9.9 x 10^23 atoms</td>
<td></td>
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</tr>
<tr>
<td>b) 150 mg x 1 g/1000 mg x 1 mole/207.2 g x 6.022 x 10^23 atoms/mole = 4.36 x 10^20 atoms</td>
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<tr>
<td>c) 1.2 L x 1.292 g/L x 0.139 x 1 mole/4.002 g x 6.022 x 10^23 atoms/mole = 3.24 x 10^22 atoms</td>
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<tr>
<td>d) 23.4 mL x 0.797 g/mL x 1 mole/46.0 g x 6.022 x 10^23 molecules/mole x 9 atoms/molecule = 2.19 x 10^24 atoms</td>
<td></td>
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Therefore, the answer is d
NOMENCLATURE

I. Binary Nomenclature (Compounds of 2 elements)

A. 2 Non-Metals (No Hydrogen)

Rule:

Greek Prefixes:

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

More examples:

PBr₃

P₄O₁₀
B. Binary - Acids (or 2 Non-metals with Hydrogen)

Rule:

HCl  ______________________________
HBr  ______________________________
H₂S  ______________________________

Try this one:  H₂Se  ______________________________

Write the formula for hydroiodic acid:  _________________
Write the formula for hydrotelluric acid:  _________________

C. 1. A Metal and a Non-metal  (These are actually monovalent metals, metals having only one common oxidation state, i.e. sodium, magnesium, aluminum, etc.

1. Rule:

Name these:

NaCl
MgCl₂
Al₂O₃

Try writing the formulas for:

Sodium sulfide
Magnesium nitride
2. Multivalent metals (Multivalent metals are metals having more than one common oxidation state. i.e. copper (I) and copper (II), tin (II) and tin(IV) among others. (Note: A table of oxidation numbers appears on the next page, as well as at the end of this book.)

2. Rule:

Determine the names for these:

CuCl  _______________________________________________
Another name:  ______________________________________________
CuCl₂  _______________________________________________
Another name:  ______________________________________________
FeCl₂  _______________________________________________
Another name:  ______________________________________________
FeCl₃  _______________________________________________
Another name:  ______________________________________________

Practice Set

HI
Zn₃N₂
CrS
N₂O₃
MnCl₃
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**1+ radicals***

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<td>(Cu^{+}) (cupric)</td>
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**1- elements**

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<th>(O^2-) (oxygen)</th>
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**1- radicals***

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<th>(BO_3^{2-}) (borate)</th>
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<tr>
<td>(ClO_4^-) (perchlorate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(BrO_3^-) (bromate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(BrO_4^-) (perbromate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>(MnO_4^-) (permanganate)</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**1- radicals (cont)**

<table>
<thead>
<tr>
<th>(C_2) (carbon)</th>
<th>(O_2) (oxygen)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(CO_2) (carbon)</td>
<td>(O_2) (oxygen)</td>
</tr>
</tbody>
</table>

**2- radicals***

<table>
<thead>
<tr>
<th>(SiO_3^{2-}) (silicate)</th>
<th>(CrO_4^{2-}) (chromate)</th>
<th>(Cr_2O_7^{2-}) (dichromate)</th>
</tr>
</thead>
<tbody>
<tr>
<td>(SiO_3^{2-}) (silicate)</td>
<td>(CrO_4^{2-}) (chromate)</td>
<td>(Cr_2O_7^{2-}) (dichromate)</td>
</tr>
</tbody>
</table>

**2- radicals (cont)**

<table>
<thead>
<tr>
<th>(O_2) (oxygen)</th>
</tr>
</thead>
</table>

**3- radicals***

<table>
<thead>
<tr>
<th>(AsO_3^{3-}) (arsenite)</th>
</tr>
</thead>
</table>

**3- radicals (cont)**

<table>
<thead>
<tr>
<th>(AsO_3^{2-}) (arsenite)</th>
</tr>
</thead>
</table>

* The term “radicals” is an older term for polyatomic ions.
II. **Ternary Compounds**

How are ternary compounds defined?

A. Ternary Acids:

Rule:

radical ending  acid ending

Examples:

HNO₂
HNO₃

Ternary acids – Practice. Pause the video and try these; then start the video again to see the answers:

H₂CO₃
HClO₄
H₂CrO₄
H₂SO₃
H₂C₂O₄
H₃BO₃
B. Ternary Salts

Rule:

Examples to name:

Sn(NO_2)_2 ________________________________

or ________________________________

Fe(NO_3)_3 ________________________________

or ________________________________

Ternary Salt Practice Set. Pause the video and try these. Then start the video again to see the answers.

ZnC_2O_4

Hg_2(NO_3)_2

Mn(BrO_3)_3

KMnO_4

NiCO_3

Ag_2SO_4
III. **Quaternary Compounds**

A. **Salts from**

1. **Diprotic acids**: What are diprotic acids?

Rule:

Give two names for each of these:

- **NaHCO₃**
  - _____________________________________________________
  - or _____________________________________________________

- **KHSO₄**
  - _____________________________________________________
  - or _____________________________________________________

2. **(Salts from ) Triprotic acids**: What are tri-protic acids??

Rule:
It is important to remember to keep track of the number of hydrogens that are
MISSING! (Actually, the name dibasic, etc., comes from the number of moles of a base
such as NaOH required to neutralize one mole of the acid. Think about it.)

Examples:

\( \text{NaH}_2\text{BO}_3 \)

\( \text{Na}_2\text{HBO}_3 \)

Give two names for each of these salts of triprotic acids:

\( \text{K}_2\text{HAsO}_3 \)

or

\( \text{ZnHBO}_3 \)

or

\( \text{KH}_2\text{PO}_4 \)

or

IV. **Hydrates:** (Think about the term)

What are hydrates?

What does the raised dot mean?

Describe how you indicate the number of water molecules?

Examples:

\( \text{CuSO}_4\cdot\text{5H}_2\text{O} \)

\( \text{H}_2\text{C}_3\text{O}_4\cdot\text{2H}_2\text{O} \)
Although the common names are introduced in the latter part of this lesson, the topic will be addressed fully in the next lesson. Start Lesson 7.

### Lesson 7:

V. Names and Formulas of common inorganic compounds:

<table>
<thead>
<tr>
<th>Common name</th>
<th>Chemical name</th>
</tr>
</thead>
<tbody>
<tr>
<td>NaOH</td>
<td></td>
</tr>
<tr>
<td>NaHCO₃</td>
<td></td>
</tr>
<tr>
<td>HC₂H₃O₂</td>
<td></td>
</tr>
<tr>
<td>H₂SO₄</td>
<td></td>
</tr>
<tr>
<td>HCl</td>
<td></td>
</tr>
<tr>
<td>CaO</td>
<td></td>
</tr>
<tr>
<td>Ca(OH)₂</td>
<td></td>
</tr>
<tr>
<td>CaCO₃</td>
<td></td>
</tr>
<tr>
<td>K₂CO₃</td>
<td></td>
</tr>
<tr>
<td>N₂O</td>
<td></td>
</tr>
</tbody>
</table>
### Nomenclature Review and Practice

<table>
<thead>
<tr>
<th>Compound</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
</tr>
<tr>
<td>ZnS</td>
</tr>
<tr>
<td>Cr₂O₃</td>
</tr>
<tr>
<td>S₂O₃</td>
</tr>
<tr>
<td>H₃PO₄</td>
</tr>
<tr>
<td>K₂SO₄</td>
</tr>
<tr>
<td>Fe₂(SO₃)₃</td>
</tr>
<tr>
<td>Na₂CO₃</td>
</tr>
<tr>
<td>NaHCO₃</td>
</tr>
<tr>
<td>Na₃AsO₄</td>
</tr>
<tr>
<td>Na₂HAsO₄</td>
</tr>
<tr>
<td>NaH₂AsO₄</td>
</tr>
<tr>
<td>CuSO₄·5H₂O</td>
</tr>
</tbody>
</table>
More Practice

Sodium ferric sulfate

Ferric ammonium sulfate

Ferrous ammonium phosphate

Lithium aluminum hydride

Barium nitrate

Ferrous borate

Arsenic acid

Nickel (II) nitrate, trihydrate
(Nickel is misspelled on the video)

Diphosphorus trioxide

Magnesium phosphide
## Table of Common Oxidation Numbers

<table>
<thead>
<tr>
<th>1+ elements</th>
<th>2+ elements</th>
<th>3+ elements</th>
<th>4+ elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>H⁺</td>
<td>Mg⁺</td>
<td>Al³⁺</td>
<td>Pb⁴⁺ (plumbic)</td>
</tr>
<tr>
<td>Li⁺</td>
<td>Ca⁺</td>
<td>Bi⁴⁺</td>
<td>Sn⁴⁺ (stannic)</td>
</tr>
<tr>
<td>Na⁺</td>
<td>S⁺</td>
<td>Cr³⁺ (chromic)</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>Ba⁺</td>
<td>Fe³⁺ (ferric)</td>
<td></td>
</tr>
<tr>
<td>Cs⁺</td>
<td>Zn⁺</td>
<td>Mn³⁺ (manganic)</td>
<td></td>
</tr>
<tr>
<td>Ag⁺</td>
<td>Cd⁺ (cuprous)</td>
<td>Ni³⁺ (nickelic)</td>
<td></td>
</tr>
<tr>
<td>Cu⁺ (cuprous)</td>
<td>Cr³⁺ (chromous)</td>
<td>As³⁺ (arsenic)</td>
<td></td>
</tr>
<tr>
<td>Hg₂⁺ (mercurous)</td>
<td>Mg²⁺ (manganous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Fe²⁺ (ferrous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Co²⁺ (cobaltous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Ni²⁺ (nickelous)</td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>Sn²⁺ (stannous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Pb⁺ (plumbous)</td>
<td></td>
<td></td>
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<tr>
<td></td>
<td>Cu⁺ (cupric)</td>
<td></td>
<td></td>
</tr>
<tr>
<td></td>
<td>Hg⁺ (mercuric)</td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>1+ radicals</strong>*</td>
<td><strong>2- elements</strong></td>
<td><strong>3- elements</strong></td>
<td><strong>5+ elements</strong></td>
</tr>
<tr>
<td>H₃O⁺ (hydronium)</td>
<td>O²⁻ (oxydide)</td>
<td>N²⁻ (nitride)</td>
<td>As⁵⁺ (arsenic)</td>
</tr>
<tr>
<td>NH₄⁺ (ammonium)</td>
<td>S²⁻ (sulfide)</td>
<td></td>
<td></td>
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<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>1- elements</strong></td>
<td><strong>2- elements</strong></td>
<td><strong>3- elements</strong></td>
<td><strong>5+ elements</strong></td>
</tr>
<tr>
<td>H⁻ (hydride)</td>
<td>O²⁻ (oxydide)</td>
<td>N²⁻ (nitride)</td>
<td>As⁵⁺ (arsenic)</td>
</tr>
<tr>
<td>F⁻</td>
<td>S²⁻ (sulfide)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cl⁻</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Br⁻</td>
<td></td>
<td></td>
<td></td>
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<tr>
<td>I⁻</td>
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<td></td>
</tr>
<tr>
<td><strong>1- radicals</strong>*</td>
<td><strong>1- radicals (cont)</strong></td>
<td><strong>2- radicals</strong>*</td>
<td><strong>3- radicals</strong>*</td>
</tr>
<tr>
<td>C₂H₃O₂⁻ (acetate)</td>
<td>HSO₄⁻ (bisulfate)</td>
<td>SO₄²⁻ (sulfate)</td>
<td>BO₃⁻ (borate)</td>
</tr>
<tr>
<td>CN⁻ (cyanide)</td>
<td>HSO₃⁻ (bisulfite)</td>
<td>SO₃²⁻ (sulfite)</td>
<td>PO₄³⁻ (phosphate)</td>
</tr>
<tr>
<td>NO₃⁻ (nitrate)</td>
<td>HS⁻ (bulfide)</td>
<td>S₂O₃⁻ (thiosulfate)</td>
<td>PO₃³⁻ (phosphite)</td>
</tr>
<tr>
<td>NO₂⁻ (nitrite)</td>
<td>HCO₃⁻ (bicarbonate)</td>
<td>CO₃²⁻ (carbonate)</td>
<td></td>
</tr>
<tr>
<td>(arsenate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>OH⁻ (hydroxide)</td>
<td>HC₂O₄⁻ (binoxolate)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>SCN⁻ (thiocyanate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ClO⁻ (hypochlorite)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ClO₂⁻ (chlorite)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ClO₃⁻ (chlorate)</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ClO₄⁻ (perchlorate)</td>
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</tr>
<tr>
<td>BrO₃⁻ (bromate)</td>
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<tr>
<td>BrO₄⁻ (perm bromate)</td>
<td></td>
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<tr>
<td>MnO₄⁻ (permanganate)</td>
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</tr>
</tbody>
</table>

* The term “radicals” is an older term for polyatomic ions.
1. (Review problem) How many atoms of oxygen are present in 36.7 L of the gas if the gas has a specific gravity of 1.11 at that temperature?

2. What is the mass number of a lead isotope having 127 neutrons?

3. If bromine has two naturally occurring isotopes, Br - 79 whose relative weight is 78.918 and occurs 50.69% of the time and Br - 81 whose weight is to be determined, what is the relative weight of the Br – 81?

4. The mass of a single atom of an element expressed in grams is approximately
   a. $10^{22}$  b. $10^{-22}$  c. $10^{24}$  d. $10^{-24}$  e. $10^{26}$  f. $10^{-26}$

5. Give correct names for the following
   - K$_2$Cr$_2$O$_7$
   - MnO$_2$
   - Limestone
   - Na$_2$O$_2$
   - NaH$_2$AsO$_4$
   - CoF$_2$
   - HClO
   - (NH$_4$)$_3$P
6. Give formulas for the following

Lithium aluminum hydride

Zinc bisulfate

Manganic phosphate, dibasic

Oxalic acid, dihydrate

Dinitrogen pentoxide

Arsenic thiocyanate

Calcium hypoiodite

Slaked lime

7. Calculate the total number of atoms of any kind in 17.4 g of cupric sulfate, pentahydrate.

8. What is meant by the term “multivalent” metals?

9. If an average chicken egg weighs 2.0 oz, what is the weight of a mole of average chicken eggs expressed in pounds?

10. Why is the atomic mass number of an isotope only approximately equal to the atomic mass for that isotope?
During this unit, Dr. Etheridge will

1. briefly discuss the "Table of Common Oxidation Numbers" which you are expected to learn.
2. explain the Law of Multiple Proportions.
3. demonstrate calculation of a compound’s percent composition.
4. calculate the empirical or simplest formula for a compound when given a quantitative analysis.
5. show how to calculate the molecular or true formula when given an empirical formula or information for finding the empirical formula as well as the quantitative analysis results.
6. demonstrate balancing chemical equations by inspection. Oxidation-reduction reactions and the methods for balancing them will be addressed in Unit 11.
7. perform calculations involving chemical reactions.
8. discuss predicting simple combination reactions, including reactions with oxygen and reactions of oxides with water.
9. introduce the concept of range of oxidation states.
10. discuss how to predict decomposition reactions, specifically the decomposition of nitrates, carbonates, bicarbonates, halates, and (with limitations) peroxides.
11. determine the limiting reagents in a reaction and show how to apply this factor in determining quantities of other reagents involved in the reaction.
12. calculate actual, theoretical, and percent yields.
13. demonstrate how to perform a combustion train analysis.
14. explain how an activity series works and how to use it in predicting simple replacement reactions.
During this unit, we will discuss:

- Determination of Formulas of Compounds
- Calculations from Chemical Equations (Stoichiometry)
- Types of Reactions
- Predicting Reactions

Lesson 8:

**DETERMINATION OF FORMULAS OF COMPOUNDS**

State the Law of Multiple Proportions:

**Percent Composition**

What is meant by the term "Percent Composition"?
For Na$_2$CO$_3$, the molecular weight is ______________

Express the percent of sodium:

Express the percent of carbon:

Determine the percent of oxygen:

**Chemical Formulas:**

Define empirical formula?

Define a molecular formula?

What is their relationship?

**Determination of Chemical Formulas**

How do we find the simplest formula?
A sample of iron oxide contains 34.97 g Fe and 15.03 g oxygen. Find the formula.

solve:

Therefore, the final formula is ________________________________.

An oxide of carbon contains 27.29% carbon. Find the simplest formula.

solve:
A 3.00 g sample of the salt Na$_2$SO$_4$yH$_2$O is heated to drive off all the water. After heating, the salt residue weighs 1.59 g. Determine the number of water molecules, y, present in the formula.

solve:

**Combustion Train Analysis**

Sketch a basic combustion train analysis apparatus and indicate what occurs in each area.
When a 2.040 g sample of TNT is combusted, 3.470 g carbon dioxide, 0.507 g water, and 1.036 g nitrogen dioxide are recovered. What is the simplest formula for TNT?

solve

HOW IS THE QUANTITY OF OXYGEN FOUND?
Finding the Molecular (or True) Formula

Describe the relationship between the simplest formula and the molecular formula?

A compound whose simplest formula is CH₂O was found to have a molecular weight of 61±3. What is the molecular formula?

The molecular weight of CH₂O is ________

Find the molecular formula in the above problem:

What is the molecular weight of the molecular formula you determined? ____________

Is it in the range set by the data?

Combustion of a 42.60 mg sample of Vitamin C produced 63.96 mg CO₃ and 17.28 gm H₂O. If the molecular weight of Vitamin C is found to be 172±10, what is the molecular formula?
CALCULATIONS FROM BALANCED CHEMICAL EQUATIONS (STOICHIOMETRY)

Rules of balancing equations by inspection:

1.

2.

3.

4.

\[
\text{CaCl}_2 + \text{Na}_3\text{PO}_4 \rightarrow \text{Ca}_3(\text{PO}_4)_2 + \text{NaCl}
\]

\[
\text{C}_8\text{H}_{18} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

\[
\text{C}_5\text{H}_{12} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
\]

**Stoichiometry**

Simple Calculations:

\[
\text{P}_4\text{O}_{10} + 6\text{H}_2\text{O} \rightarrow 4\text{H}_3\text{PO}_4
\]

What information is given from a balanced equation?

Why is the mole:mole ratio so important?

How is this ratio expressed?

\[
\text{quantity} \rightarrow \text{moles} \rightarrow \text{moles} \rightarrow \text{quantity}
\]
What weight of phosphoric acid could be produced when 3.6 g $\text{P}_4\text{O}_{10}$ is allowed to react with excess water?

As happens sometimes, Dr. Etheridge stuck her foot in her mouth and started the calculation using quantity of water rather than phosphoric acid. Don’t take notes until you hear her correct her mistake.

Equation:

solve:

---

Calculate the weight of Cu needed to react with 10.3 g $\text{HNO}_3$.

$$3\ \text{Cu} + 8\ \text{HNO}_3 \rightarrow 3\ \text{Cu(NO}_3)_2 + 2\ \text{NO} + 4\ \text{H}_2\text{O}$$

solve:

---

What weight of Cu is needed to produce 6.1 g NO

$$3\ \text{Cu} + 8\ \text{HNO}_3 \rightarrow 3\ \text{Cu(NO}_3)_2 + 2\ \text{NO} + 4\ \text{H}_2\text{O}$$

solve:
(Please remember to learn the common oxidation numbers.)

**Multiple Equations:**

\[
P_{4}O_{10} + 6H_{2}O \rightarrow 4H_{3}PO_{4}
\]

\[
3Zn + 2H_{3}PO_{4} \rightarrow Zn_{3}(PO_{4})_{3} + 3H_{2}
\]

The relationship that exists between the two equations is strictly limited to the quantity of material produced in the first equation that is the quantity consumed in the second one.

Calculate the weight of zinc needed to react with the phosphoric acid produced when 10.0 g P\textsubscript{4}O\textsubscript{10} reacts by this method!

solve:

**Limiting and Excess Reagents**

Limiting Reagents:

What is meant by the term “limiting reagent”?

What is meant by the term “excess reagent”?  

48
3 Zn + 2 H₃PO₄ → Zn₃(PO₄)₂ + 3 H₂

Calculate the weight of hydrogen produced when 5.0 g zinc reacts with 10.0 g phosphoric acid.

solve:

Zn:

H₃PO₄:

Encircle the quantity of hydrogen produced. This is determined by the limiting reagent which happens to be: ______________

Why is that substance the limiting reagent?

Using the following equation, calculate the weight of NO gas produced when 12.3 g Cu reacts with 100.0 mL of 20.0% HNO₃ having a density of 1.10 g/mL.

3 Cu + 8 HNO₃ → 3 Cu(NO₃)₂ + 2 NO↑ + 4 H₂O

Cu:

HNO₃:

The limiting reagent is ______________

How much NO is produced? ________ WHY?
Determine the percent of the excess reagent used in the above problem:

**Percent Yield**

Write the formula for finding percent yield:

When 8.3 g Cu reacts with nitric acid, 2.1 g NO is recovered. Find the percent yield.

\[ 3 \text{ Cu} + 8 \text{ HNO}_3 \rightarrow 3 \text{ Cu(NO}_3)_2 + 2 \text{ NO} + 4 \text{ H}_2\text{O} \]

Solve for theoretical yield:

Solve for percent yield:

This next problem comes after a brief discussion of reaction types. Dr. Etheridge just cannot resist working one more problem!

\[ 3 \text{ Cu} + 8 \text{ HNO}_3 \rightarrow 3 \text{ Cu(NO}_3)_2 + 2 \text{ NO} + 4 \text{ H}_2\text{O} \]

If 5.0 g Cu and 5.0 g HNO\(_3\) are allowed to react, what is the limiting reagent?

solve:
Lesson 11:

TYPES OF REACTIONS

I. Decomposition Reactions:

Five Types of Decomposition Reactions:

A. Metal nitrates + heat → metal oxides + O₂↑ + NO₂↑

   Examples:

   exception:

   IA metal nitrates + heat → metal nitrites + O₂↑

   Examples:

B. Metal carbonates + heat → metal oxides + CO₂↑

   Examples:

   exceptions:

   IA metal carbonates do not tend to decompose. Remember that!
C. IA metal bicarbonates + heat $\rightarrow$ IA metal carbonate + CO$_2$\textsuperscript{↑}

Examples:

Metal bicarbonate beyond the IA metals are rare. In those rare instances when they do occur, heating them produces:

Other metal bicarbonates + heat $\rightarrow$ metal oxides + CO$_2$\textsuperscript{↑} + H$_2$O\textsuperscript{↑}

Examples:

D. Metal halates + heat $\rightarrow$ metal halides + O$_2$\textsuperscript{↑}

define halates

Examples:

E. Peroxides

You are expected to know these peroxides and these reactions.

Na$_2$O$_2$ + H$_2$O $\rightarrow$

BaO$_2$ + heat $\rightarrow$

H$_2$O$_2$ + heat/light $\rightarrow$

This is why peroxides must be kept in tightly closed, brown bottles.
II. Simple Combination Reactions

A. Metals and non-metals

Watch your oxidation numbers!

Examples:

\[ \text{Na} + \text{Cl}_2 \rightarrow \]
\[ \text{Al} + \text{Cl}_2 \rightarrow \]
\[ \text{Zn} + \text{S} \rightarrow \]
\[ \text{Al} + \text{S} \rightarrow \]

B. Combustion Reactions:

Define combustion:

Is oxygen a requirement?

Examples of combustion using oxygen:

\[ \text{C} + \text{O}_2 \rightarrow \]
\[ \text{H}_2 + \text{O}_2 \rightarrow \]
\[ \text{CH}_4 + \text{O}_2 \rightarrow \]
\[ \text{C}_3\text{H}_6\text{O}_2 + \text{O}_2 \rightarrow \]
\[ \text{ZnS} + \text{O}_2 \rightarrow \]
\[ \text{S} + \text{O}_2 \rightarrow \]
Lesson 12:

Combustion Reactions (Continued)

Range of Oxidation States

Give the range of oxidation states for each of the following families/groups

IA

IIA

IIIA

IVA

VA

VIA

VIIA

\[ \text{C} + \text{O}_2 \rightarrow \text{CO}_2 \quad \text{Is this reasonable?} \quad \text{Why?} \]

\[ \text{CO}_2 + \text{O}_2 \rightarrow \text{CO}_3 \quad \text{Is this reasonable?} \quad \text{Why?} \]

\[ \text{S} + \text{O}_2 \rightarrow \text{SO}_2 \quad \text{Is this reasonable?} \quad \text{Why?} \]

\[ \text{SO}_2 + \text{O}_2 \rightarrow \text{SO}_3 \quad \text{Is this reasonable?} \quad \text{Why?} \]

\[ \text{SO}_3 + \text{O}_2 \rightarrow \text{SO}_4 \quad \text{Is this reasonable?} \quad \text{Why?} \]
# Periodic Table of the Elements

<table>
<thead>
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<td>24.305</td>
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<td>Ce</td>
<td>Pr</td>
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<tr>
<td>Th</td>
<td>Pa</td>
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C. Metal oxides with water:

*Metal oxides react with water to form bases.*

Examples:

\[ \text{MgO} + \text{H}_2\text{O} \rightarrow \]

\[ \text{Fe}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \]

\[ \text{Al}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \]

Give another name for metal oxides: ______________________________

D. Non-metal oxides with water:

*Non-metal oxides react with water to form acids.*

Examples:

\[ \text{CO}_2 + \text{H}_2\text{O} \rightarrow \]

\[ \text{SO}_2 + \text{H}_2\text{O} \rightarrow \]

\[ \text{SO}_3 + \text{H}_2\text{O} \rightarrow \]

How can you determine which acid is formed?

*You have to figure out which NON-metal OXIDE will react with water to produce the acid under consideration.*

\[ \text{N}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \text{HNO}_3 \text{ or HNO}_2 \]

\[ \text{ } + \text{H}_2\text{O} \rightarrow \text{HClO}_4 \]
III. Metathetical Reactions: We will deal with this in Unit 4.

IV. Single Replacement Reactions: Note the activity series on the right: You may be required to memorize this.

Describe the way in which the Activity Series works.

Activity Series

Examples:

K
Ba
Sr
Ca
Na
Mg
Al
Mn
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Sb
As
Bi
Cu
Ag
Pd
Hg
Pt
Au

What is another name for single replacement reactions?

A + BC → AC + B iff* A is above B on the activity series.

*iff means “if and only if”
V. Reactions to produce Hydrogen

A. Single Replacement reactions: Again, you will need your activity series.

Metals above hydrogen replace hydrogen from an acid to produce hydrogen gas.

Examples:

Lesson 13:

Reactions to produce hydrogen (continued)

A. Active metals with water.

What are the “active metals”?

Active metals react with cold water to produce bases and H₂ (gas)

Examples:

B. Aqueous bases with aluminum and zinc (or post-transition elements)

What are “post-transition” elements?

Aqueous bases with Al and Zn produce hydrogen gas:

Examples:
What is meant by “galvanized”? ________________________________

There is another term “anodized” which means “to coat with a protective metal.” Frequently, the metallic coating is of aluminum. Therefore, when you hear the term “anodized” you may expect it to refer to an aluminum coating, unless otherwise indicated.

VI. Reactions to produce oxygen

1. Decomposition of halates

2. Decomposition of nitrates

3. Decomposition of peroxides

4. Addition of sodium peroxide to water

Joseph Priestly, who is credited with the discovery of oxygen, use the following reaction for preparation of the pure gas:

\[ 2 \text{HgO} + \Delta \rightarrow 2 \text{Hg} + \text{O}_2 \uparrow \]

red oxide               quicksilver
of mercury

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Review/Practice of Reactions

1. A piece of zinc is placed in a nickelous nitrate solution. Write a balanced equation if a reaction will occur:

2. Calcium oxide is dissolved in water. Write a balanced equation if a reaction occurs.

3. Your gold pin falls into acetic acid solution. Write a balanced equation if a reaction occurs.

A reactions quiz follows on the next page!
Give a complete, balanced equation for each of the following reactions, but only if a reaction occurs. If no reaction will occur, write “no reaction.”

1. Carbon burned in limited air.

2. The product of #1 reacts with excess oxygen.

3. Calcium oxide is dissolved in water.

4. Aluminum turnings are dropped into hydrochloric acid.

5. Limestone is heated.

6. Magnesium is burned in an atmosphere of bromine.

7. Potassium bromate is heated.

8. Iron and sulfur are ground together and ignited. (more than one equation is correct)
9. Butane, $\text{C}_4\text{H}_{10}$ is burned in a limited amount of oxygen.

10. Butane, $\text{C}_4\text{H}_{10}$, is burned in air.

11. Sulfur dioxide is dissolved in water.

12. Magnesium metal is roasted (burned) in air and the product is dissolved in water (2 equations).

13. Copper metal is dropped into nickel (II) chloride solution.

14. A galvanized bucket is used to hold hydrochloric acid.

15. Baking soda is thrown on a fire.

16. Your silver bracelet is dropped into a beaker of hydrochloric acid.

17. Vinegar is poured into a copper bowl.
18. Potassium nitrate is heated gently.

19. Chromic carbonate is heated

20. Your bottle of hair bleach (hydrogen peroxide) is left open.

21. Washing soda, Na₂CO₃, is heated strongly.

22. Diphosphorus pentoxide is dissolved in water.

23. Br₂O is dissolved in water.

24. Fe₂O₃ is dissolved in water.

25. CuO is dissolved in water.

26. An industrial discharge is removed using CaO as a scrubber. CaSO₄ is produced.

27. Sodium chlorate is heated.
28. A strip of magnesium ribbon is placed in a copper (II) sulfate solution.

29. A calcium shaving is placed in a copper (II) sulfate solution.

30. A silver ribbon is dropped into hydrochloric acid.

31. A calcium shaving is placed in water.

32. Sodium peroxide is sprinkled onto water.

33. When a gas is dissolved in water, chloric acid formed.

34. When two compounds react, calcium hydroxide is produced.

35. When a gas is dissolved in water, bromous acid is produced.
In this unit, Dr. Etheridge will

1. discuss terms such as concentrated, dilute, saturated, unsaturated, solute, solvent, titration, etc.

2. explain the concept of strong and weak electrolytes, specifically discussing strong acids, strong bases, and soluble salts.

3. give you a list of strong acids and bases and the solubility rules, and expect you to commit them to memory.

4. explain how to predict methathetical reactions which are dependent on
   a. solubilities
   b. formation of water
   c. formation of a gas.

5. explain the difference between total and net ionic equations.

6. demonstrate how to complete and balance total and net ionic equations.

7. calculate the molarity of a solution.

8. discuss finding molarity of solutions from assays.

9. explain how to perform calculations using molarity (including titrations).

10. demonstrate how to find concentrations of solutions using percent by mass, ppm, ppb, etc.

11. perform dilution calculations.
In this unit we will cover:

Definitions

Metathetical Reactions

Electrolytes

Types of Solutions

Concentration of Solutions

Stoichiometry

PPM and PPB

Lesson 14:

Definitions:

Solution -

How many kinds of solutions are there?  

Solute -

Solvent -

Insoluble –
THREE TYPES OF METATHETICAL REACTIONS:

What are metathetical reactions?

What form do they take?

For these reactions to occur, at least one of the reactants must be mobile in solution.

The three cases for metathetical reactions are:

   a. A precipitate is formed
   b. Water is formed
   c. A gas is produced

**a. The reaction occurs because an insoluble material (usually, a precipitate) is produced**

**Solubility Rules:** (You must learn these)

1. 

2.  

   exceptions:

3.  

   exceptions:

4.
Consider the reaction:

\[ \text{AgNO}_3 + \text{NaCl} \rightarrow \]

How could you have predicted this?

Stop the video for a moment and think about these two questions:

What do you think would happen when a Na\(^+\) collides with a Ag\(^+\)?

What do you think would happen when a Cl\(^-\) collides with a NO\(_3^-\)?

Consider the reaction:

\[ \text{KI} + \text{Pb(NO}_3)_2 \rightarrow \text{PbI}_2 + \text{KNO}_3 \]

At what point should you balance the equation?

Now try working these:

\[ \text{Na}_2\text{CO}_3 + \text{KI} \rightarrow \]

\[ \text{Na}_2\text{CO}_3 + \text{Ni(NO}_3)_2 \rightarrow \]

\[ \text{Na}_2\text{CO}_3 + \text{CuSO}_4 \rightarrow \]
A few more to try:

\[ \text{Na}_2\text{ClO}_4 \ + \ \text{CaCl}_2 \rightarrow \]

\[ \text{ZnSO}_3 \ + \ \text{Ni(NO}_3)_2 \rightarrow \]

\[ \text{NH}_4\text{Cl} \ + \ \text{Hg}_2\text{(NO}_3)_2 \rightarrow \]

b. The reaction occurs because water is formed - Neutralization

(1) Acid + Base → Salt + Water

examples:

(2) Metal oxide (basic anhydride) + Acid → Salt + Water

another name for a metal oxide is ___________________________

a metal oxide will be have like a _______________________

examples:
(3) Non-metal oxide (acid anhydride) + base → Salt + Water

another name for a non-metal oxide is ________________________________

non-metal oxides behave chemically like ___________________________

examples:

(4) Similar Form (but no water is formed)

Metal oxide + Non-metal oxide → Salt (notice: NO WATER is formed)

Review/Practice Equations

CaO + H$_3$PO$_4$ →

ZnO + HC$_2$H$_3$O$_2$ →

CaO + HCl →

Al$_2$O$_3$ + SO$_2$ →
Lesson 15:

c. The reaction occurs because a gas is formed (and probably bubbles off):

Particularly, we are concerned with the formation of CO$_2$, SO$_2$, H$_2$S, and HCN

1. Formation of CO$_2$. Look for the formation of H$_2$CO$_3$, but write as H$_2$O + CO$_2$

   Note how carbonic acid should be written (when it is a product).

   examples:

   

2. Formation of SO$_2$. Look for the formation of H$_2$SO$_3$, but write as H$_2$O + SO$_2$

   How should sulfurous acid be written (when it is a product)?

   example:

   

3. Formation of H$_2$S (gas) which you will write as H$_2$S

   Describe the physical properties of H$_2$S:

   example of formation

4. Formation of HCN (gas) which you will write as HCN

   Describe HCN

   example of formation:
Reactions for Practice

1. FeS + H$_2$SO$_4$(aq) →

2. CaCO$_3$ + HBr$_{(aq)}$ →

3. K$_2$SO$_3$ + HCl$_{(aq)}$ →

At this point, we have covered these types of chemical reactions and the factors which drive them:

1. Combination (in Unit 3)
2. Decomposition (in Unit 3)
3. Simple Replacement (in Unit 3)
4. Reactions to produce hydrogen (in Unit 3)
5. Reactions to produce oxygen (in Unit 3)

3. Metathesis (in Unit 4)
   a. A precipitate is formed
   b. Water is formed
   c. An insoluble gas is formed

We have one more type reaction to cover: Oxidation-Reduction.

This will be introduced in Unit 11.
ELECTROLYTES:

Define electrolytes:

Characteristics:

What is meant by the term “fused”?

What are strong electrolytes?

What are weak electrolytes?

What are non-electrolytes?

What types of substances are strong electrolytes? (These substances ionize virtually 100%)
Strong Acids: (What is meant by strong acid?)

List the strong acids and learn them:

How should strong acids be written (we are assuming they are in solution)?

Give examples:

What is the hydronium ion and how is it written?

How should weak acids be written? ___________________________

Why?

Strong Bases:

Strong Bases (what does the strength depends on?)

What about NH₄OH? Is it strong or weak? __________

How should NH₄OH be written? _____________________________
Which are the strong bases?

Except for ammonium hydroxide, why are the weak bases weak?

**Why** is ammonium hydroxide a weak base?

Why are strong bases strong? (If this seems redundant, it is because we are trying to make a point!)

How should strong bases, in general, be written? ______________

examples:

**Soluble Salts**

Soluble salts should be written as ____________.

examples:

Remember: Strong acids, strong bases, and soluble salts are written as IONS!
NET AND TOTAL IONIC EQUATIONS

What does a total ionic equation show?

Example: \( \text{NaOH} + \text{HCl} \rightarrow \)

Write as an ionic equation:

That equation is called a ________________________________.

A net ionic equation shows only______________________________________________

When ions remain throughout the process and are not changed, these ions are called spectator ions.

Write the net ionic equation for the \( \text{NaOH} + \text{HCl} \) reaction.

Give both total and net ionic equations when solutions of silver nitrate and sodium chloride are mixed.

Total ionic equation:

Net ionic equation:

We shall use this next reaction in the next lesson.

The reaction between solutions of ammonium hydroxide and aluminum nitrate.
Ammonium hydroxide can be written several ways. Give several:

Solutions of ammonium hydroxide and aluminum nitrate are mixed. Give:

Total ionic equation:

Net ionic equation:

What is/are the spectator ion(s) for this reaction?
TYPES OF SOLUTIONS

What is meant by the term, “saturated solution”?

Explain what is meant by the term “equilibrium” as it is used here:

What happens when the temperature is changed?

Unsaturated solution:

Supersaturated solution:

How are supersaturated solutions produced?
Concentration of Solutions:

Percent Solutions (m = mass, v = volume):

For all practical purposes, there are three (3) types of percent solutions:

1. m/m

What weight of sugar is needed to make 100.0 g of a 15.0% sugar solution?

2. v/v

What volume of alcohol is needed to make 100.0 mL of a 15.0% alcohol solution?

Why can you not use calculations to determine with certainty that it would require 85.0 mL water to complete preparing this solution?

When is this type concentration most likely to be used?
What weight of sugar is needed to make 100.0 mL of a 15.0% sugar solution?

When is this type solution most likely to be used?

**Molarity**

Molarity (M) is *moles of solute per liter of solution*

As a mathematical formula, it is expressed:

![Volumetric Flask](image)

This is calibrated to _________________________

Devices calibrated to contain may not accurately deliver that volume. Devices calibrated to deliver may not accurately contain that volume.

Remember: The volumetric flask is calibrated to contain, but not to deliver.
Calculate the molarity of a solution made by dissolving 5.4 g NaCl in 250 mL solution.

Calculate the molarity of a solution made by dissolving 10.3 g of sodium sulfate in 600.0 mL of solution.
Assays

What information is given in an assay?

The assay for a shipment of hydrochloric acid indicates it is 38.0% HCl and has a specific gravity of 1.19. Determine the molarity of the shipment.

FYI: A solution in which alcohol is the solvent is often called a “tincture” in medical chemistry.

A sulfuric acid shipment was received with an assay of 93.2% sulfuric acid and a specific gravity of 1.84. Determine the molarity.
Dilution of Solutions

When you dilute a solution, what 2 factors **change**?

What DOES NOT change? ________________________________

Therefore,  $\text{Moles}_{\text{before dilution}} = \text{Moles}_{\text{after dilution}}$

   another way:...

   another way…

**REM: THE NUMBER OF MOLES DOES NOT CHANGE**

Approximately what volume of water is needed to dilute 25.0 mL of 18.0 M sulfuric acid to 3.0 M

Solve:

What goes into the container first? WHY
Calculate the molarity of a solution made by diluting 40.0 mL of 6.00 M acid to 250.0 mL.

Solve:

**Stoichiometry**

Recall:

\[
\text{quantity} \rightarrow \text{moles} \rightarrow \text{moles} \rightarrow \text{quantity}
\]

Calculate the volume of 0.20 M sulfuric acid needed to titrate 25.0 mL of 0.30 M NaOH to a good endpoint.

Solve:

buret

this device is calibrated to deliver, not to contain
Calculate the molarity of a NaOH solution if 20.0 mL required 17.2 mL of 0.80 M H₃PO₄ for good endpoint.

solve:

Calculate the volume of 3.0 M Nitric acid needed to react completely with 5.0 g Cu

\[ 3 \text{Cu} + 8 \text{HNO}_3 \rightarrow 3 \text{Cu(NO}_3)_2 + 2 \text{NO} + 4 \text{H}_2\text{O} \]

solve:

Pause the video and attempt to work this problem. The solution appears on the adjoining page, however, try not to look until you think you have it worked.

If 20.0 mL of 0.30 M AgNO₃ and 40.0 mL of 0.10 M CaCl₂ are mixed, find the mass of precipitate formed and the concentration of all ions remaining in solution.

Write the equation

Write the total ionic equation

solve:
Solution to problem on page 85

\[
\text{Ca}^{++} + \text{Cl}^{-} + \text{Ag}^{+} + \text{NO}_3^{-} \rightarrow \text{AgCl} \downarrow + \text{Ca}^{++} + \text{NO}_3^{-}
\]

40.0 mL of 0.10 M  20.0 mL of 0.30 M

4.0 mn  8.0 mn  6.0 mn  6.0 mn \leftarrow \text{quantities before the reaction occurs}

- 6.0 mn  - 6.0 mn  \text{which forms} 6.0 mn

4.0 mn  2.0 mn  0 mn  6.0 mn  6.0 mn \leftarrow \text{after the reaction}

Remember \( M = \frac{n}{L} \) or \( M = \frac{mn}{mL} \)

Note the volumes: 40.0 mL + 20.0 mL = 60.0 mL (With these dilute solutions, this is a reasonable assumption.)

For \( \text{Ca}^{++} \), \( \frac{4.0 \text{ mn}}{60.0 \text{ mL}} = 0.067 \text{ M} \)

For \( \text{Cl}^{-} \), \( \frac{2.0 \text{ mn}}{60.0 \text{ mL}} = 0.033 \text{ M} \)

For \( \text{NO}_3^{-} \), \( \frac{6.0 \text{ mn}}{60.0 \text{ mL}} = 0.10 \text{ M} \)

For \( \text{AgCl(s)} \), 6.0 mn x 143.4 g/n = 860.4 mg = 0.860 g

Did you get the correct answers? If so, congratulations! If not, you should rework the problem to make sure your error wasn’t a simple math error. If the error was conceptual, it is important that you rework this problem until you understand every detail. Otherwise, these concepts may come back to haunt you, later.
PPM and PPB

PPM or ppm (parts per million) and PPB or ppb (parts per billion) are units frequently encountered when dealing with extremely dilute solutions such as those encountered in water pollution, nutrition, and similar situations. The same principles used to manage ppm and ppb problems can be used for parts per thousand and parts per trillion (PPT or ppt), but care must be taken so as to have no confusion about whether ppt is referring to parts per thousand or parts per trillion. In all cases the reference is to the number of grams of a substance per total grams of solution. For example: 3 ppm means 3 grams solute per 1 million grams of solution. As a rule, these solutions are so dilute, the density of water may be considered to be unchanged by the solute and 1.00 g/mL may be used.

Let’s try a sample problem.

If the limit for copper in drinking water is 1.3 ppm, calculate the maximum allowed mass of copper in 1.0 L of drinking water. Express your answer in milligrams.

Solution:

\[
1.0 \text{ L} \times \frac{1000 \text{ mL}}{\text{L}} \times \frac{1.0 \text{ g}}{\text{mL}} \times \frac{1.3 \text{ g}}{100000 \text{ g}} \times \frac{1000 \text{ mg}}{\text{g}} = 1.3 \text{ mg Cu}
\]
1. Indicate whether each of the following is soluble or insoluble:

PbCl₂ ______________________

Hg₂Br₂ _____________________

NaI  ______________________

FeCl₃ ______________________

PbI₂ ______________________

AgBr ______________________

MgSO₄ _____________________

BaSO₄ _____________________

Zn(C₂H₃O₂)₂ ______________________

(NH₄)₂PO₄ _____________________

Na₃AsO₃ ______________________

BaCO₃ ______________________
2. Antimony, a contaminant found in discharge from petroleum refineries, fire retardant industries, etc, has a MCLG (maximum contaminant level goal) set at $6 \times 10^{-2}$ ppm. If a 500 mL sample was analyzed and found to contain $2 \times 10^{-3}$ mg antimony, did that sample exceed the MCLG?

3. Calculate the concentration of each ion remaining in solution and the quantity precipitate formed (if any) when 12.0 mL of 0.400 M HCl is allowed to react with 15.6 mL of 0.500 M Pb(C$_2$H$_3$O$_2$)$_2$. 


In this unit Dr. Etheridge will

1. discuss the properties of gases including motion, volume, attraction of gas molecules, elasticity, kinetic energy and average kinetic energy.

2. review Boyle’s, Charles’, Amonton’s, Avogadro’s, Dalton’s, and Graham’s laws and explain them in part using the Kinetic Molecular Theory of Gases.

3. explain the workings of a barometer and a manometer, how they are used, and their units.

4. expand the mole roadmap to include gases.

5. discuss and apply the concept of STP.

6. demonstrate how to perform calculations using the ideal gas law.

7. explain and use the concept of wet vs. dry gases in the context of Dalton’s law.

8. describe and explain terms such as effusion and diffusion, and demonstrate calculations related to these concepts.

9. explain stoichiometric relationships in chemical equations when gases are involved.

10. discuss kinetic energy relationships in gases.

11. briefly discuss root mean square speed and show how to find the root mean square speed of a particular type atom.

12. describe the progression of change occurring on a Maxwell-Boltzmann distribution curve and give particular attention to the changes which occur as the temperature is increased.

13. discuss the ways in which real gases deviate from ideal gases.

14. demonstrate how to apply the van der Waals equation in calculations involving volume and pressure in non-ideal circumstances.
During this unit, we will discuss the following

Properties of Gases

Gas Laws

Stoichiometry

Kinetic Molecular Theory

Deviations from Ideality

Lesson 18:

Kinetic Theory and the Properties of Gases:

Describe the motion of gas molecules:

Explain what is meant by “perfectly elastic”?

How significant is the actual volume of a gas molecule?

Under what conditions does the volume of a single molecule become significant?
What kind of attraction do gas molecules have for one another?

*Every body in the universe attracts every other body with a force that is directly proportional to their masses and inversely proportional to the square of the distance between them!*

The average kinetic energy of all gases is ______________ at the same temperature.

Average kinetic energy of gases depends **ONLY** upon ________________!

Give the formula describing the kinetic energy of a gas:

**The Gas Laws:**

**Boyle’s Law** states:

Express the law, algebraically:

How do we measure pressure?

Explain how a manometer works:

Explain how a barometer works:
Why is atmospheric pressure measured in units of length?

Why is mercury used in many barometers?

What is a Torr? ______________

Give a variety of pressure units equivalent to an atmosphere.

Calculate the volume occupied by 300 mL of a gas collected at 750 Torr if the pressure is changed to 820 Torr.

solve:

Calculate the pressure required to change 650 L of gas at 700 Torr to a volume of 500 L at constant temperature?

solve:
Charles Law states:

What would the volume of a gas at 0 Kelvins be? _________ WHY?

Why do we use temperature in Kelvins, rather than °C?

(We’ll ask this question and explain “why” in the next lesson)

(The comment Dr. Etheridge was attempting to make when the college logo came on was this… If you use formulas, remember the algebraic expressions for direct and inverse proportions. PV is inverse, as defined by xy = k. However, V/T is direct, as defined by x=ky.)

Lesson 19:

Before we continue with Charles’ Law, let us give attention to:

Motion of gas molecules:

Describe/define each of the following:
Translational motion

Rotational motion

Vibrational motion
State Charles’ Law

Why must temperature be expressed in Kelvins, rather than in °C?

Express Charles’ Law as a formula.

Calculate the temperature required to change 650 L of gas collected at 21°C to a volume of 500 L, if the pressure is held constant.

solve:

**Amontons’ Law** states

What temperature scale must be used? ____________________

What would happen if a closed (sealed) container were heated strongly?

Express Amontons’ Law as a formula.
If a gas had a pressure of 680 Torr at 25 °C, what would the pressure in the closed container be at 100° C?

solve:

Gay-Lussac’s Law proposed:

Avogadro’s Law states:

Knowing the density of a gas at STP, it is simple to accurately calculate the molar volume of that gas at STP. We’ll see density of a gas discussed in just a bit.

One mole of a gas at STP occupies ________________ (on the average)

What is STP?

An expanded version of the mole roadmap appears on the next page:
How many moles of hydrogen are present in 350 L at 105 kPa and 14 °C?

solve:

**Ideal Gas Law** (General Gas Law):

Give the formula for the ideal gas law (general gas law)

Give the value for R, including the units! Show how it can be found, mathematically in the event you should forget the value. Take the time to use other units, such as mL, Pa, etc.
Lesson 20:

STP is the temperature/pressure/volume relationship that allows us to compare gases.

\[ R = 6.236 \times 10^4 \text{ mL} \cdot \text{Torr} \text{ mol}^{-1} \text{ K} \]  

(\text{Dr. Etheridge left out the power of 10 on the video})

Having \( n \) designating moles allows the ideal gas law to be used in many ways.

Calculate the number of molecules of CO in 286 mL gas at 41 °C and 726 Torr.

solve:

275 mL of gas collected in a gas bulb was found to weigh 0.812 g at 29°C and 740 Torr. Calculate the molecular weight (molar mass) of the gas.

You should pause the video at this point and work this problem out. Dr. Etheridge will show you how it is worked. If you have difficulty with this, be sure to see your instructor or a tutor.

solve:
Density of Gases

Derive the formula for density of a gas, using PV=nRT as your starting point.

Calculate the density of oxygen at –24 °C and 790 Torr.

(You should stop the video and solve this problem. The solution will appear when you return.)

solve:

Dalton’s Law or Dalton’s Law of Partial Pressures states:

Give the mathematical expression for Dalton’s Law:

Give the mathematical expression for the expanded view of Dalton’s Law.
**Mole Fraction**

What is mole-fraction?

What is the sum of all mole fractions in the mixture? _________________

How may the partial pressure of a gas in a mixture be expressed?

---

If a gas mixture composed of 3.2 g He, 4.9 g oxygen and 5.5 g Ar exerts a total pressure of 980 Torr, what is the pressure of the He?

(Pause the video and find the number of moles of each gas present. Then, restart the video.)

solve:

---

We will begin the next lesson with “Collecting Gases over Water”

---

GO TO DISC 6
Lesson 21:

Collecting Gases Over Water

Think about this: If H₂ gas is generated and collected over water, what is the composition of the gas present in the collecting bottle?

Vapor pressure is a function of temperature (a gas collected over water is always considered to be “saturated” with water vapor). Why can this be said?

340 mL of gas is collected over water at 20 °C and a pressure of 758 Torr. Find the volume of the dry gas at the same conditions.

Graham’s Law

What is the difference between effusion and diffusion?

(The laws we address here will cover both diffusion and effusion. )
Graham’s law states:

An expression or formula is:

What is rate? What are some appropriate units of rate?

If oxygen effuses from a container at the rate of 3.64 mL/sec, what is the molecular weight of a gas which effuses at 4.48 mL/sec from the same container?

solve:

Calculate the rate at which CO₂ will effuse from a container if helium effused from that container in 3.8 sec.

solve:
KINETIC MOLECULAR THEORY OF GASES

The kinetic energy, $E_k$, for a mole of gas molecules can be expressed as:

$$E_k =$$

Further, we know that for an ideal gas

$$PV = \frac{1}{3} nM \bar{v}^2$$

Substitute the first equation into the second equation and you will have…
(Pause the video and perform the substitution.)

To relate this to temperature…..

$$E_k = \frac{3}{2} RT$$
(Think about the values for R)

Thus, the kinetic energy of a gas is dependent on and only on the temperature! It matters not what the gas is or whether the gas is pure or a mixture of gases.
The information on the next two pages is for your individual study. Dr. Etheridge does not address this in the video, but it is a logical extension of this topic.

This first item is informational and can be quite useful. Some other values for R

<table>
<thead>
<tr>
<th>Value</th>
<th>Description</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.08205 L·atm/mol·K</td>
<td>8.314 J/mol·K</td>
</tr>
<tr>
<td>8.314 kg·m²/s²·mol·K</td>
<td>8.314 kPa·dm³/mol·K</td>
</tr>
<tr>
<td>8.314 kPa·dm³/mol·K</td>
<td>1.9872 cal/K·mol</td>
</tr>
</tbody>
</table>

You will find these values for R handy when you are performing calculations using various units. As we continue the study of chemistry, you will find even greater use for these constants.

**Root Mean Square Speed (rms)** is the speed of a molecule having the average molecular kinetic energy.

Consider: \( PV = \frac{1}{3} nMu^2 \), where “u” is the root mean square speed.

And \( PV = nRT \)

Since: \( PV = PV \)

Then: \( \frac{1}{3} nMu^2 = nRT \)

\[
\frac{1}{3} Mu^2 = RT
\]

\[
u^2 = \frac{3RT}{M}
\]

\[
u = \sqrt{\frac{3RT}{M}}
\]

This is the equation for the root mean square speed of a gas molecule. Note, it is dependent on temperature and molar mass/molecular weight/atomic weight.
Maxwell-Boltzman Distribution Curve of Molecular Speeds

Notice: As the temperature increases, the average speed and range of speeds increase.

This will be relevant to our study of evaporation in a later unit.

Also, the majority of the molecules lie within one standard deviation of the mean, but some molecules still fall within the very high and very low speed areas under the curve. Therefore, some molecules will always have a leaving velocity, an important factor in evaporation and the rate of evaporation.
STOICHIOMETRY USING GASES

Calculate the weight of Zn needed to react with HCl in order to produce enough hydrogen to fill 3 bottles, each having a volume of 250 mL at 28 °C and 738 Torr.

Write your balanced equation, first.

solve:
What volume of ammonia gas at 10 °C and 30.24 in. Hg should be dissolved in water to produce 2.5 L of a 4.8 M NH₄OH solution?

Be very careful, this problem involves two different “kinds” of volume!

solve:

When are gases most likely to behave as “ideal”?

What conditions are necessary for ideality?

In reality:

Gas molecules may have a significant volume. When?

Gas molecules may be attracted to one another. When is this significant?
THE VAN DER WAALS EQUATION:

Write the van der Waals equation and indicate the units of $a$ and $b$.

What do each of the following represent?

$a$

$b$

Calculate the pressure exerted by 3.0 moles of chlorine gas held in a 10.0 L container at -40 °C. Use both methods. ($a = 6.493$ L atm/mol, $b = 0.05622$ L/mol.)
ACID RAIN:

What causes acid rain?

What are some of the problems acid rain produces?

What volume of nitrogen gas will react with 8.9 L hydrogen gas at 21 °C and 700 Torr?

solve:

(Dr. Etheridge has an error in the answer to the first part. It should be 2.96 L, not 2.904 L)
1. How many molecules does a 5.2 liter balloon filled with oxygen gas at STP contain?

2. How many molecules of oxygen will that balloon contain if it is heated to 93°C and a pressure of 1.8 atm?

3. A diver is breathing a mixture of helium and oxygen. What happens to the pressure of oxygen in the tank if a quantity of nitrogen equivalent to the pressure of the helium is introduced?

4. When 3.4 g Zn is added to an excess of HCl_{aq}, the gas produced is collected at 774 torr and 23°C. What volume of gas should be expected?

5. At what rate will SO_{2} effuse from a container allowing 5.0 L of CO_{2} to diffuse in 3.0 minutes?

6. The greatest deviations from the kinetic molecular theory are expected under what conditions of temperature and pressure?
7. Hydrogen, helium, carbon dioxide, and ammonia gases, all having the same pressure, are allowed to mix well before a small opening is made in the container. After 10 sec, the container will have the least of which gas? Why?

8. Hydrogen gas is collected over water at a total pressure of 738 torr. At that temperature, the water vapor pressure is 17.5 mm. What is the mole fraction of hydrogen in the mixture?

9. What will be the pressure if 0.300 mole He and 0.250 mol N₂ are mixed in a 5.0 L container at 24°C?

10. Nitric oxide, NO, and oxygen react to form nitrogen dioxide. If 300 mL of nitric oxide is allowed to react with 200 mL of oxygen, all at STP, what is the final volume of the reaction mixture upon reaction completion?
In this unit Dr. Etheridge will

1. discuss heat capacity, specific heat, joules, calories, endothermic and exothermic reactions, enthalpy, and the work concept.

2. discuss the Law of Conservation of Energy which is also called the First Law of Thermodynamics.

3. explain the difference between heat and temperature and discuss which units are used for each concept.

4. introduce internal energy ($\Delta E$), work ($w$), and heat transferred ($q$), and demonstrate thermodynamic calculations.

5. discuss the concept of Hess's Law of Constant Heat Summation and demonstrate its application to processes.

6. introduce the concept of a state function, specifically using enthalpy. Other state functions will be introduced in a later course.

7. demonstrate how to use enthalpies of formation to predict the heat of a reaction.

8. discuss the concepts of heat capacity, molar heat capacity, and specific heat capacity.

9. introduce the calorimeter as a device employed in the study of heat capacity and show how to determine the heat capacity of a particular calorimeter.
Energy in Reactions

During this unit, we will discuss the following:

The First Law of Thermodynamics
Heats of Reaction
Enthalpy
Hess’s Law of Constant Heat Summation
Heat Capacity, etc.
Calorimetry

Lesson 23:

THE FIRST LAW OF THERMODYNAMICS

State the law:

Describe Kinetic Energy:

Describe Potential Energy
Terms:
  system

  surroundings

  universe

What does “E” represent?

**Internal Energy**

\[ \Delta E_{\text{system}} + \Delta E_{\text{surroundings}} = 0 \]

State this in simple terms:

What is \( \Delta E \)?

How is heat transferred?

a)

b)
\[ \Delta E_{\text{system}} = q + w \]

\[ q \]

\[ w \]

Relative to the SYSTEM, what is meant by

\[ +q \]

\[ +w \]

\[ -q \]

\[ -w \]

Explain what is meant by \( q_p \)

Another expression for \( q_p \) is \( \Delta H \), which is called ______________________

Define this:

A thought: Could we have a \( q_v \)? ______ What would it mean?
HEATS OF REACTION

Remember, $\Delta E = q + w$

$$q = \Delta E - w$$

and

$$w = p\Delta V$$ at constant pressure

- $-P\Delta V$ if volume increases
  
  WHY?

+ $+P\Delta V$ if volume decreases
  
  WHY?

Consider the reaction: $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO(} \text{s) + CO}_2(\text{g})$

$$q = \Delta E - w$$

but, $w = -P\Delta V$  WHY?

therefore

$$q = \Delta E - (-P\Delta V)$$

or

$$q = \Delta E + P\Delta V$$

A reaction may be:

Exothermic:

Endothermic:
ENTHALPY:

Enthalpy is a state function.

What is a state function:

\[ \Delta H = H_{\text{products}} - H_{\text{reactants}} \]

For a constant pressure reaction, \( q_p = \Delta H \)

Define and explain enthalpy of formation

When bonds are formed, energy is __________________________ -

When bonds are broken, energy is __________________________

Describe/define the parts of the following expression:

\[ \Delta H_f^0 \]

The units are:

List those factors considered as standard state conditions?

Consider the following equation: This is an example of a compound, HCl, being formed from the elements. Note the diatomic nature of the hydrogen and chlorine. Don’t forget to include the state of matter of each substance. Give the enthalpy of formation for each substance and we will begin here with the next lesson.

\[ \text{H}_2(\text{g}) + \text{Cl}_2(\text{g}) \rightarrow 2 \text{HCl}(\text{g}) \]

(You should find a table of thermodynamic values in the back of your textbook or a good reference book such as a CRC Handbook of Chemistry and Physics.)
Calculate the heat of this reaction:

\[ \text{H}_2(g) + \text{Cl}_2(g) \rightarrow 2 \text{HCl}(g) \]

How much heat is evolved in the formation of just **one mole** of HCl?

**Exothermic vs. Endothermic**

If \( \Delta H \) is negative, the reaction is _____________________________

If \( \Delta H \) is positive, the reaction is _____________________________

Calculate the enthalpy change for the combustion of acetylene, \( \text{C}_2\text{H}_2(g) \).

Look on page 125 at the end of this unit to see a sample thermodynamics table with the values you need for this problem.

\[ 2 \text{C}_2\text{H}_2(g) + 5 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]
Calculate the heat evolved when 2.0 g acetylene undergoes combustion

\[ 2 \text{C}_2\text{H}_2(g) + 5 \text{O}_2(g) \rightarrow 4 \text{CO}_2(g) + 2 \text{H}_2\text{O}(g) \]

**Hess’s Law of Constant Heat Summation:**

State the law:

Show an example:

Show how the equations cancel.
Using the following equations, calculate the molar heat of formation of MnO₂.

\[
\begin{align*}
\text{MnO}_2(s) & \quad + \quad \text{Mn}(s) \quad \rightarrow \quad 2 \text{ MnO}(s) \quad \Delta H = -240 \text{ kJ} \\
2 \text{ MnO}_2(s) & \quad \rightarrow \quad 2 \text{ MnO}(s) \quad + \quad \text{O}_2(g) \quad \Delta H = +264 \text{ kJ}
\end{align*}
\]

The final equation should be \( \text{Mn(s)} + \text{O}_2(g) \rightarrow \text{MnO}_2(s) \).

**HEAT CAPACITY**

Define:

Units?

**Molar Heat Capacity:**

Define:

**Specific Heat Capacity:**

Define and give units:

The molar heat capacity of most metals heavier than Calcium is very close to 25.0 J/mol°C.

\[
\text{Specific heat} \times \text{atomic weight} = 25.0 \text{ J/mol °C}
\]
Lesson 25:

CALORIMETRY

Calorimetry is a study of heat changes using a device called a “calorimeter.”

Describe the

simple, “coffee-cup” calorimeter

a thermos-type calorimeter

Remember, for a constant pressure reaction or situation:

\[ q_{\text{in}} + q_{\text{out}} = 0 \]
If a 20.0 g piece of hot iron at 440 °C is placed in 300 mL water at 20 °C, what is the final temperature we should expect?

\[ q_{\text{iron}} + q_{\text{water}} = 0 \]

\[ q = mc\Delta T \quad (\text{where } q \text{ is at constant pressure}) \]

where \( q \) is heat

\( m \) is mass

\( \Delta T = T_f - T_i \)

now, solve the problem:

**Finding the heat capacity of the calorimeter:**

When 20 mL water is added to a calorimeter, the temperature of the calorimeter plus water is found to be 20 °C. A 30 mL sample of water at 90 °C is added, and the final temperature is found to be 56 °C. What is the heat capacity of the calorimeter?

solve:

(Pause the video and work the problem. Note that Dr. Etheridge gets 1265.4 J for the first part of the problem, but the answer to the first part should be 1255.2 J.)
When 2.4 g NaOH(s) is added to 100 mL water in a calorimeter at 21.2 °C, the temperature rose to 24.3 °C. Calculate the molar heat of dissolving NaOH. (The specific heat of the solution is 4.21 J/g°C.

(There is a unit error on the screen. The $q_{\text{reaction}}$ is found to be $-1445\text{J}$ for 2.4 g NaOH, not the $-1445\text{kJ}$ shown. Apparently Dr. Etheridge was having a bad calculator day …sigh.)

solve:

Consider this process:

\[
\begin{align*}
\text{SO}_2(g) & \rightarrow \text{S(s)} + \text{O}_2(g) \quad \Delta H = +297 \text{ kJ} \\
\text{S(s)} + \frac{3}{2} \text{O}_2(g) & \rightarrow \text{SO}_3(g) \quad \Delta H = -396 \text{ kJ}
\end{align*}
\]

Calculate $\Delta H$ for

\[
\text{SO}_2(g) + \frac{1}{2} \text{O}_2 \rightarrow \text{SO}_3(g)
\]
## Brief Table of Thermodynamics Values

<table>
<thead>
<tr>
<th>substance</th>
<th>$\Delta H^\circ_f$ (kJ/mol)</th>
<th>$S^\circ$ (J/mol·K)</th>
<th>$\Delta G^\circ_f$ (kJ/mol)</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{C}_2\text{H}_2$ (g)</td>
<td>+226.7</td>
<td>+200.8</td>
<td>+209.2</td>
</tr>
<tr>
<td>$\text{CO}_2$ (g)</td>
<td>-393.5</td>
<td>+213.6</td>
<td>-394.4</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}$ (l) *</td>
<td>-285.8</td>
<td>+69.9</td>
<td>-237.2</td>
</tr>
<tr>
<td>$\text{H}_2\text{O}$ (g) *</td>
<td>-241.8</td>
<td>+188.7</td>
<td>-228.6</td>
</tr>
<tr>
<td>$\text{O}_2$ (g)</td>
<td>0</td>
<td>+205.0</td>
<td>0</td>
</tr>
</tbody>
</table>

*Pay very close attention to these states of matter. Make sure you know which state of matter your problem is using. This occurs especially with water vapor as opposed to the condensed, liquid form of water, or occasionally ice.

A quiz follows on the next page.
1. Which of the following is not true for an endothermic reaction?
   a. The change in enthalpy for the reaction is positive.
   b. Heat flows from the surroundings into the system.
   c. The temperature of the surroundings decreases.
   d. The products have a lower enthalpy than the reactants.

2. If a beaker in which a reaction is occurring feels cold, the reaction is
   a. endothermic
   b. exothermic
   c. spontaneous
   d. a closed system

3. A student makes the mistake of performing an endothermic reaction in a calorimeter, but
   did not consider the heat absorbed or given off by the calorimeter, itself. Under these
   circumstances, the $\Delta H$ calculated would
   a. have a larger negative value.
   b. have a smaller negative value.
   c. have a larger positive value.
   d. have a smaller positive value.
   e. be the same since the calorimeter value would cancel.

4. Consider the reaction for the combustion of ethyl alcohol:

   \[
   \text{CH}_3\text{CH}_2\text{OH}(l) + 3 \text{O}_2(g) \rightarrow 2 \text{CO}_2(g) + 3 \text{H}_2\text{O}(l) \quad \Delta H = -1370 \, \text{kJ}
   \]

   Which of the following is a true statement?
   a. The reaction is exothermic.
   b. The $\Delta H$ would be different if the water produced were a gas.
   c. The product volume is greater than the reactant volume.
   d. both a and b are true

5. Which of the following are state functions?
   a. energy and work   b. work and heat   c. energy and enthalpy   d. work and enthalpy

6. Which is true?
   a. Hess’s Law violates the 1st Law of Thermodynamics
   b. $\Delta H$ is the same as heat at constant pressure
   c. Ice melting is exothermic   d. all of these are true
7. In a given process the following is found:
   \( q = 18 \text{ kJ} \) and \( w = 14 \text{ kJ} \)
   Which of the following statements is/are true?
   a. The system does work on the surroundings.
   b. Heat flows from the system to the surroundings.
   c. \( \Delta E = 32 \text{ kJ} \)
   d. All of the above are correct.

8. When 1.00 g of C\(_8\)H\(_{18}\) (liquid octane) undergoes complete combustion in the presence of oxygen, carbon dioxide gas and liquid water are formed and 47.9 kJ heat is given off. Write a balanced equation for the reaction and include the \( \Delta H \) associated with the balanced equation.

9. From the following data:

   \[
   \begin{align*}
   \text{C}_2\text{H}_4 (g) + 3 \text{O}_2(g) & \rightarrow 2 \text{CO}_2(g) + 2 \text{H}_2\text{O}(l) \quad \Delta H = -1410.9 \text{ kJ} \\
   2 \text{C}_2\text{H}_6(g) + 7 \text{O}_2(g) & \rightarrow 4 \text{CO}_2(g) + 6 \text{H}_2\text{O}(l) \quad \Delta H = -3119.4 \text{ kJ} \\
   2 \text{H}_2(g) + \text{O}_2(g) & \rightarrow 2 \text{H}_2\text{O}(l) \quad \Delta H = -571.6 \text{ kJ}
   \end{align*}
   \]
   calculate the \( \Delta H \) for
   \[
   \text{C}_2\text{H}_4(g) + \text{H}_2(g) \rightarrow \text{C}_2\text{H}_6(g)
   \]
In this unit, Dr. Etheridge will

1. explain the relationships among frequency, wavelength, and energy.

2. discuss the way in which selected wavelengths of energy in the electromagnetic spectrum impact molecules and atoms.

3. place radio waves, microwaves, infrared radiation, visible light, ultraviolet radiation, x-rays and cosmic rays on an energy continuum. Students should know the order of energies for the colors red, orange, yellow, green, blue, indigo, and violet.

4. discuss emission spectra, absorption spectra, line emission spectra, continuous emission spectra, and the electromagnetic spectrum.

5. briefly discuss the work of Bohr, Lyman, Balmer, Paschen, Brackett, Pfund, and Rydberg. The contributions of Bohr and Rydberg are especially important.

6. explain what is meant by the term: the Quantum Theory of Radiation. Students should be conversant with terms such as ground state, excited state, quanta, and quantum numbers.

7. describe the Bohr model and its limitations.

8. go over wave-particle duality (matter waves) and the de Broglie equation.

9. explain the fundamentals of the Heisenberg Uncertainty Principle.

10. briefly discuss some of the ideas of Erwin Schrödinger.

11. discuss wave functions and orbitals.

12. describe, and show how to use four quantum numbers.

13. discuss the concept of electron density

14. demonstrate how to draw the s, p, and d orbitals.
During this unit, we will discuss the following:

A Brief History Lesson

The Electromagnetic Spectrum

The Bohr Atom

The Concept of Wave-Particle Duality

Quantum Numbers

Lesson 26:

A BRIEF HISTORY LESSON

The Conclusions of J. J. Thomson:

Approximately when did he do this work? _______________

Describe his conclusions:
Ernest Rutherford’s work and his conclusions:

Begin by describing the alpha particle:

Describe his “gold foil” experiment:

What were his conclusions and the logic he used in arriving there:
(You may need to pause the video in order to write up a thorough note on this.)
THE ELECTROMAGNETIC SPECTRUM

Light and the Properties of Light:

Describe wavelength

What units are used? ________________________________
What Greek letter represents it? ___________________

What is frequency?

What Greek letter represents frequency? __________
Give the units for frequency: ______________________

What is the speed of light? _______________________

Give the equation for the relationship between wavelength, frequency, and the speed of light:

The college radio station broadcasts at 90.7 Mhz. Calculate the wavelength of this signal in nm.

solve:
The energy factor

\[ E = h\nu \]

The college radio station broadcasts at 90.7 Mhz. Calculate the energy of one wavelength of this signal.

(You should pause the video and try working this. Then restart the video.)

solve:

Your cup holding 200 mL coffee is placed in front of the radio set to 90.7 MHz. How much will the temperature change when the coffee absorbs 1.00 mole of radio waves at that frequency?

(Again, you should pause the video and try working this. Then restart the video.)

solve:
Characteristics of the Electromagnetic Spectrum

Pause the video and sketch the electromagnetic spectrum, noting radio/TV waves, microwaves, the infrared region, the visible region, the ultraviolet region, X-rays, and gamma rays.

In which direction does wavelength increase? (use arrows)

In which direction does frequency increase?

In which direction does energy increase?

Give the order of colors in the visible spectrum (in order of increasing energy) and note approximate wavelengths in nanometers. (The spectrum on page 137 may help or refer to a spectrum in your textbook.)

What happens to molecules absorbing radio-frequency radiation?

What happens to molecules absorbing microwave radiation?

What happens to molecules absorbing infrared radiation?

What happens to molecules absorbing ultraviolet radiation?
Lesson 27:

More terms you need to know:

Emission spectrum -

Continuous spectrum -

Line spectrum -

Absorption spectrum -

Prelude to the Bohr Atom:

Briefly describe the research of Balmer, Lyman, Paschen, Brackett, and Pfund.
How was Rydberg’s work different?

You should know that Rydberg was exciting hydrogen contained in a glass tube. The hydrogen gas would glow, hence emitting light that produced a line emission spectrum. Rydberg derived his equation by trial and error, until he arrived at the constant that would work within the mathematical relationship he found; not by proposing a theory and testing it.

Write the Rydberg Equation and indicate what each letter represents.

There is more than one “Rydberg Equation.” State another variation:

What do the stationary states represent?

Give the expression for the energy of a hydrogen electron in the $n$th stationary state.
Calculate the wavelength, frequency, and energy that corresponds to a change from \( n = 4 \) to \( n = 2 \).

solve:

<table>
<thead>
<tr>
<th>transitions from</th>
<th>produces lines in the</th>
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<tbody>
<tr>
<td>&gt;1 to 1</td>
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<td>&gt;2 to 2</td>
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<td>&gt;3 to 3</td>
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<tr>
<td>&gt;4 to 4</td>
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<td>&gt;5 to 5</td>
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An excellent site to visit for additional information regarding the electromagnetic spectrum is found at:

THE BOHR ATOM

Describe Bohr’s Solar System Model of the atom

How do the “n” stationary states relate to the Bohr atom?

Describe what happens when electrons move from one level (or orbit or stationary state) to another, according to the Bohr model.

Remember: Bohr described electrons as if they were particles

Electrons are in an orbit depending on their _____________.

The mathematical expression showing the number of electrons in an orbit is:

If n = 8, how many electrons could theoretically be present? __________

Give some limitations of the Bohr atom:
Heisenberg Uncertainty Principle.

State the principle:

The sports car illustration Dr. Etheridge used may not be the best one to use here.

WAVE-PARTICLE DUALITY (de Broglie concept)

Describe what is meant by “wave-particle” duality.

Write the de Broglie equation and label each variable:

Describe the properties of a moving particle.

Why is the matter-wave concept important?
A baseball weighing 5.50 oz is thrown at 100 mph. What is the characteristic wavelength of this moving baseball?

solve:

REMEMBER: \[ J = \text{kg} \cdot \text{m}^2/\text{s}^2 \] You should know this! Memorize it.

Calculate the de Broglie wavelength of a neutron traveling at 100 mph. The relative weight of a neutron is 1.00866.

(You should pause the video and try solving this yourself.)

solve:

Which had the greater mass-velocity component, the baseball or the neutron? ______________

Which had the greater wavelength component, the baseball or the neutron? ______________

As the mass-velocity component increases, what happens to the wavelength component?
Describe the concepts introduced by Erwin Schrödinger

Two terms you need to know:

wave function

orbitals (this is NOT referring to orbits)

QUANTUM NUMBERS (The first three were introduced by Schrödinger)

\[ n \]

\[ \ell \]

\[ m_{\ell} \]

\[ s \]
From the information on the previous page, you should be able to fill in the following table:

<table>
<thead>
<tr>
<th>$n$</th>
<th>$\ell$</th>
<th>$m_l$</th>
<th>$s$</th>
<th># electrons</th>
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</table>

Remember, every electron has to have a unique set of quantum numbers. No two electrons may have the same set of four quantum numbers. For example $1\ 0\ 0\ +\frac{1}{2}$ and $1\ 0\ 0\ -\frac{1}{2}$ are two different sets of quantum numbers. At least one number must be different in each set.

GO TO DISC 8
Return to the previous page and fill in the table for \( n = 5 \).

| For the known elements of today, the 5th energy level is like the 4th, except further from the nucleus, the 6th is like the 3rd, but further from the nucleus, and the 7th is like the 1st. |

Theoretically, the table can be expanded indefinitely.

What is the purpose of having a model for the atom?

Information from quantum numbers (and the table):

**size of energy levels:** How many different shells does the model propose for the known atoms of the day?

**shape:**
Note the repeating of the values from one energy level to the next.

What is the model shape when \( l \) is 0

Draw the shape:

Which line in the emission spectrum is the source of the designation \( s \)?

In which energy level or levels does \( s \) appear?
Describe or draw the shape when $l$ is 1

Which line in the emission spectrum is the source of the designation $p$?

In which energy level does $p$ appear for the first time? __________

In which energy levels does $p$ appear after that? ________________

Describe its orientation(s)

What shape is present when $l$ is 2

Which line in the emission spectrum is the source of the designation $d$?

In which energy level does it appear for the first time? ____________

Thereafter, in which energy level(s) does it appear? ________________

Describe its orientations

What is the source of the $f$ designation?

In which energy levels does it appear? ________________

How many orientations does it have? ________________
Questions Asked by Dr. Etheridge
(The answers appear on the videotape)

1. How do you remove an electron from an atom?

2. How many different wavelengths of light (or lines on an emission spectrum) would you expect when an electron excited from \( n = 1 \) to \( n = 5 \) returns to the ground state?

   What is ground state?

   Sketch the diagram used to prove your answer for the number of wavelengths of light that should be expected.

3. Which of the following is/are impossible?

   a. \( 2 \ 1 \ -1 \ +3/2 \)
   b. \( 2 \ 1 \ -1 \ -1/2 \)
   c. \( 3 \ 2 \ 0 \ +1/2 \)
   d. \( 3 \ 3 \ -2 \ -1/2 \)
   e. \( 4 \ 1 \ -2 \ +1/2 \)
   f. \( 23 \ 14 \ -8 \ -1/2 \)
4. Give the four quantum numbers which describe following:

a. The first electron in the 4d
b. The fifth electron in the 6p
c. The eighth electron in the 5f
d. The third electron in the 4f.

A Quick Quiz

1. When NaCl is burned, the flame produced is yellow. The frequency of light given off by this flame is greater than ___ , but less than ___.
   a. radio waves, infrared waves
   b. blue light, orange light
   c. red light, green light
   d. blue light, x-rays

2. What is the wavelength of a photon of light whose frequency is 4.6 X 10^{14} Hz?
   a. 153 nm   b. 460 nm   c. 652 nm   d. 814 nm

3. Alpha particles directed to a thin sheet of gold foil may
   a. be slightly diverted by attraction to electrons.
   b. be reflected back when in contact with nuclei.
   c. pass straight through with no change in path.
   d. all of the above.

4. Of the following, which has the shortest wavelength?
   a. infrared radiation   b. ultraviolet radiation   c. red light
   d. blue light

5. When an electron in the hydrogen atom transitions from n = 4 to n = 1,
   I. energy is absorbed
   II. energy is emitted
   III. the electron gains energy
   IV. the electron loses energy
In this unit, Dr. Etheridge will

1. discuss the four quantum numbers, what each represents, and the relationships among them.

2. show how to convert theoretical orbital notation for an electron to four permissible quantum numbers.

3. show how to give the ground state electron configuration for any element using the aufbau process and will pay special attention to the common exceptions, Cu, Cr, Ag, and Pd.

4. demonstrate the use of the noble gas or simplified notation as well as the full notation.

5. explain the Pauli Exclusion Principle, Hund's Rule, and the Heisenberg Uncertainty Principle. It is important that students verbalize the concepts within these principles.

6. address and use paramagnetism and diamagnetism.

7. describe the contributions of the following to the history of the modern periodic table: Dobereiner, Newlands, Mendeleev, Meyer, and Moseley. (You may find that the work of Dobereiner, Newlands, and Moseley are not included in your textbook.)

8. describe the Periodic Table in detail, discussing types of elements, the several ways of numbering families, family names and characteristics, as well as changes within both families and periods.

9. explain the term, "periodicity," and describe the progression of properties with respect to sizes of atoms and ions, isoelectronic series, metallic-nonmetallic character, first ionization potential, and electron affinity.
During this unit, we will discuss the following:

The Aufbau Process
Hund’s Rule of Maximum Multiplicity
The Pauli Exclusion Principle
Important Exceptions
The Periodic Table

In the portion of this unit involving the Periodic Table,

we will discuss the following:

Historical Aspects
Periodicity and Electron Configurations
Types of Elements
Progression of Properties
THE AUFBERG PROCESS

Recall the chart:

<table>
<thead>
<tr>
<th>$n$</th>
<th>$\ell$</th>
<th>$m$</th>
<th>$s$</th>
<th># e⁻</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0 (s)</td>
<td>0</td>
<td>+1/2</td>
<td>2 2e⁻</td>
</tr>
<tr>
<td></td>
<td>1 (p)</td>
<td>-1, 0, +1</td>
<td>-1/2</td>
<td></td>
</tr>
<tr>
<td>2</td>
<td>0 (s)</td>
<td>0</td>
<td>+1/2</td>
<td>2 8e⁻</td>
</tr>
<tr>
<td></td>
<td>1 (p)</td>
<td>-1, 0, +1</td>
<td>-1/2</td>
<td>6</td>
</tr>
<tr>
<td>3</td>
<td>0 (s)</td>
<td>-1, 0, +1</td>
<td>+1/2</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>1 (p)</td>
<td>-2, -1, 0, +1, +2</td>
<td>-1/2</td>
<td>10</td>
</tr>
<tr>
<td>4</td>
<td>0 (s)</td>
<td>-1, 0, +1</td>
<td>+1/2</td>
<td>2</td>
</tr>
<tr>
<td></td>
<td>1 (p)</td>
<td>-2, -1, 0, +1, +2</td>
<td>-1/2</td>
<td>6 32e⁻</td>
</tr>
<tr>
<td></td>
<td>2 (d)</td>
<td>-3, -2, -1, 0, +1, +2</td>
<td></td>
<td>10</td>
</tr>
<tr>
<td></td>
<td>3 (f)</td>
<td>-3, -2, -1, 0, +1, +2, +3</td>
<td></td>
<td>14</td>
</tr>
</tbody>
</table>
Rules of the Aufbau Process: This relates to ground state electron configurations.

1.

2.

3.

4. (an observation)

When orbitals have several orientations, these orbitals are said to be ______________________.

   These orbitals differ only about their ________________________________.

**HUND’S RULE OF MAXIMUM MULTIPLICITY**
State the rule:

Explain what is meant by “spin”?

What do + and – spins mean?
ELECTRON CONFIGURATIONS

H (1)

He (2)

Li (3)

Be (4)

B (5)

C (6)

N (7)

O (8)

F (9)

What is meant by “paramagnetic”?

What is meant by “diamagnetic”?
Now, let’s look at Potassium:

\[ K \ (19) \]

Diagram the order of filling:

What is the \( n + l \) formula?

Use the \( n + l \) formula for potassium

What is the Noble Gas Core method?

Use the Noble Gas core for Nitrogen:

Use a Noble Gas core for Aluminum (#13)

**PAULI EXCLUSION PRINCIPLE**

State the Pauli Exclusion Principle.

Explain what is meant by “spin paired” electrons.
Dr. Etheridge misspoke herself when she said the 3 d was half filled. She meant the 3 d was filled. You probably caught that quickly.

Ag  (Pause the video and give this configuration)

What are the 4 quantum numbers of the indicated electron? _______________________

Pd

What is unusual about the electron configuration of Pd?
Electron Configurations of Ions:

Na⁺

S²⁻

Now, let’s look at ions of some of the transition elements:

Fe

Fe²⁺

Fe³⁺

Show the electron configurations for these elements and ions and note the electrons lost first for this transition element.

Cu

Cu⁺

Cu²⁺
HISTORICAL ASPECTS

Describe the contributions of:

William Prout:

Johann Döbereiner:

John Newlands:

What was missing from Newlands’ list? ________________________________
Why were they missing?

Dmitri Mendeleev

Lothar Meyer

Henry Moseley
Refer to the periodic table on the adjacent page

Mark two additional ways of numbering the families/groups on the table.

Note the

Alkali metals
Alkaline earth metals
Halogens
Noble gases
Iron group
metals
non-metals
metalloids

PERIODICITY AND ELECTRON CONFIGURATIONS

Make sure you can relate electron configurations to position on the periodic table. You may wish to pause the video and color note the following elements:
Note the s block elements

the p block elements

the d block elements

the f block elements

Give the electron configuration of Sb from its position on the periodic table:
FOUR TYPES OF ELEMENTS

The representative elements

The Noble Gases

The transition elements

The inner transition elements

PROGRESSION OF PROPERTIES

Describe the progression of metallic character over the table.

Describe the progression of atomic radius over the table. Intervening shells of electrons may be expected to mitigate the attraction which exists between the nucleus and the outer electrons, but do not remove the force of attraction.
Describe an isoelectronic series:

In that series, which would have the largest radius? Explain

Sketch the 1st ionization potential diagram for Li through Ar
What is electron affinity?

Removing an electron is an exothermic process, thus the SIGN of the energy will be _________.
If the process is endothermic, the SIGN of the energy will be _______________.

Explain why the alkali metals would have a negative electron affinity value:

Why do the alkaline earth metals have a positive value for electron affinity?

Why does chlorine have the most negative (exothermic) electron affinity?

An important question:
1. How do the sizes of atoms and ions compare?
A Quick Quiz

1. Using the noble gas core format, give the electron configuration of Cu and the four quantum numbers of the last electron in the 3d.

2. Are all elements listed in order of increasing atomic weights? __________
   Where are the exceptions?

3. What is “Periodic Law”? Is this the original meaning or definition? If not, how as the term been redefined?

4. An element, E, has an electron configuration of [Xe]4f⁴5d⁷6s² belongs to which type of elements?
   a. representative   b. transition   c. inner transition   d. noble gases

5. How many valence electrons do all halogens have? ________

6. How many valence electrons do all alkali metals have? ________ Is hydrogen an alkali metal? Why or why not?

7. How many valence electrons do the alkaline earth elements have? ____________ Is helium an alkaline earth element? Why or why not?

8. List the following atoms in order of increasing first ionization energy: Li, Na, C, O, F

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9. Which statement is true?
   a. A sodium atom has a larger radius than a potassium atom.
   b. A fluorine atom has a smaller first ionization potential than an oxygen atom.
   c. A neon atom has a larger radius than an oxygen atom.
   d. A cesium atom has a larger first ionization potential than a lithium atom.

10. The newly discovered element Mt, named for Lise Meitner of Germany, is expected to have properties similar to those of
    a. Lr      b. W      c. Ir      d. Rn

11. Lenses that change color in the presence of sunlight contain silver chloride as a component of the glass. When exposed to light, the AgCl dissociates according to the following equation:

    \[ \text{AgCl(s)} \rightarrow \text{Ag(s)} + \text{Cl} \]

    The \( \Delta H \) for this reaction is 310 \( \text{kJ/mol} \). If all the energy is provided by the light striking the lens,

12. How many electrons in an atom can have the following quantum numbers?

   a. \( n = 2 \)
   b. \( n = 3, \ell = 0 \)
   c. \( n = 2, \ell = 2, m = 0 \)
   d. \( n = 2, \ell = 0, m = 0, m_s = \frac{1}{2} \)
Unit 9

In this unit, Dr. Etheridge will

1. address the two major types of bonds and explain how all bonds are a “mixture” of these two types.
2. describe the ionic lattice that is formed.
3. characterize the ionic bond.
4. describe the relationship between lattice energy and size, lattice energy and magnitude of charge; as well as explain how these factors impact melting points and solubilities.
5. discuss the three types of covalent bonds.
6. explain the term “bond length.”
7. explain the term “bond strength.”
8. show how to develop Lewis dot structures for molecules and demonstrate how these structures may be used to identify resonance hybrids.
9. use the concept of formal charge to determine whether or not a proposed structure is feasible.
10. explain the octet rule and predict exceptions to that rule.
11. show how to recognize and describe a free radical.
12. develop the concept of electronegativity and use it to predict the type of bond formed between elements.
13. discuss the periodic trends in electronegativity.
14. explain what is meant by dipole moments, show how to predict dipole moments and how to use the direction of these moments to predict polar covalent bonding.
15. demonstrate predicting types of bonding using the periodic table and the concept of net dipole moment.
During this unit we will discuss the following:

- Ionic Bonds
- Covalent Bonds
- Lewis Structures
- Resonance Hybrids
- Formal Charge
- Exceptions to the Octet Rule
- Polar Bonds

**Lesson 33:**

**IONIC BONDS**

What is an ionic bond?

What are cations? ___________________________ anions? ___________________________

Explain the term, ionic lattice

List the important characteristics of the ionic bond:
How is it that the ionic bond is “non-directional”?

What is lattice energy?

The force of attraction between the ions is directly proportional to the product of the charges and inversely proportional to the square of the distance between the ions. Write the formula and identify the parts of the formula describing this:

Lattice energy is related to:

- magnitude of charge (explain)
- ionic radius (explain)

The higher the lattice energy, the _________________ the melting point.

Explain how melting point relates to

**size of ions**

Example: For the NaX compounds,

- NaF > NaCl > NaBr > NaI

Explain how melting point relates to
magnitude of charge

\[ \text{TiCl}_4 \ > \ \text{FeCl}_3 \ > \ \text{CaCl}_2 \ > \ \text{KCl} \]

To summarize:

The Bonding Continuum:

ionic \hspace{2cm} \text{covalent}

Are molecules “totally” ionic or covalent? ____________________
COVALENT BONDS

Define the covalent bond:

Give the three types:

Define bond length?

Define bond strength:

Using the bonds between two carbon atoms, describe the relationship between single bonds, double bonds, and triple bonds using
- bond strength

-bond length
LEWIS STRUCTURES (Lewis Dot Diagrams)
What are Lewis Structures?

The rules for producing Lewis structures:
1. Determine the central atom.
2. Determine the total number of electrons to be distributed.
3. Assign 8 electrons to each atom (with some exceptions that we will discuss).

Give a Lewis Structure for \( \text{SO}_3^- \)
(Pause the video and attempt the structure)

Give a Lewis Structure for \( \text{NO}_2^- \)
(Pause the video and attempt this structure.)

RESONANCE HYBRIDS

Sketch the two resonance structures for \(\text{NO}_2^-\) and explain whether or not they are the same.
Describe what is meant by “parents” in resonance hybrid structures, as opposed to the resonance hybrid structure itself.

**Lesson 34:**

What is a resonance hybrid?

Do all parents contribute equally to the hybrid offspring structure? ________ Explain.

What is a resonance structure?

**FORMAL CHARGE**

Rules for determining formal charge:

1. 

2. 

Formal charge =
Put in the electrons and show the bonds for HNO₃

O

N O H

O

Now, determine the formal charge on the atoms in the structure and write them in.

The sum of the charges on a molecule must equal ________________ or for a radical, the sum of the charges must equal ________________________________.

What are the characteristics of a good Lewis diagram (structure)?

Repair the above structure for HNO₃ by showing another resonance structure. Figure the formal charge on each of the atoms. Check the formal charges for this structure:

O

N O H

O
Describe how formal charge can be used to determine which “parents” contribute more to the resonance hybrid.

Formal charge allows us to discriminate among those structures which contribute more or less to the overall structure of a resonance hybrid.

Using formal charge, determine which is better, and WHY? (Pause the video and complete these structures.)

A. \( \text{O} = \text{C} = \text{O} \)

B. \( \text{O} - \text{C} \equiv \text{O} \)

EXCEPTIONS TO THE OCTET RULE

What are the three major types of exceptions?

1. 

2. 

3. 

1. The Incomplete Octet

Sketch the two structures for BF\(_3\) and determine the better structure using the Formal Charge concept.
The fluorine atom will NEVER have a charge that has any positive character. It will be either negative or zero.

When do you expect to have an incomplete octet?

2. The Expanded Octet:

   Sketch the Lewis structure for ICl₃ and use it to explain the expanded octet.

3. Odd Number of Electrons:

   These atoms (radicals) are called ________________________.

   These structures are in a highly reactive state!

   Sketch the Lewis dot diagram for the chlorine atom and for the NO₂ molecule. Do you see that these structures can “double up” to find stability?
POLAR BONDS

Polar bonds are defined as:

Define electronegativity:

Partial Table of Electronegativities

<table>
<thead>
<tr>
<th></th>
<th></th>
<th>B</th>
<th>C</th>
<th>N</th>
<th>O</th>
<th>F</th>
</tr>
</thead>
<tbody>
<tr>
<td>H</td>
<td>2.2</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Li</td>
<td>1.0</td>
<td>Be</td>
<td>1.6</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Na</td>
<td>0.9</td>
<td>Mg</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>K</td>
<td>0.8</td>
<td>Ca</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Rb</td>
<td>0.8</td>
<td>Sr</td>
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<td></td>
</tr>
<tr>
<td>Cs</td>
<td>0.8</td>
<td>Ba</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fr</td>
<td>0.7</td>
<td>Ra</td>
<td>0.9</td>
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<td>In</td>
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<tr>
<td></td>
<td></td>
<td>Sn</td>
<td></td>
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<tr>
<td></td>
<td></td>
<td>Sb</td>
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<tr>
<td></td>
<td></td>
<td>Te</td>
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</tr>
<tr>
<td></td>
<td></td>
<td>I</td>
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<td></td>
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<td></td>
</tr>
<tr>
<td></td>
<td></td>
<td>Xe</td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

Note: These values may or may not agree exactly with the values in your textbook as these values are updated periodically. The principle is the same.

Describe the periodic trends for electronegativity:

The most electronegative element is ________________.

The least electronegative element is ________________.
Describe the way in which electronegativities are used to determine

ionic bonds

polar covalent bonds

non-polar covalent bonds

Dipole Moments

What are dipole moments?

Describe the dipole moment and its notations for gaseous HCl.

Draw the structure of the water molecule and indicate dipoles and the net dipole moment.
A Quick Quiz

1. The most ionic bond will be found in
   a. NaCl  b. LiF  c. KCl  d. LiCl  e. KF

2. Which of the following should you expect to be most ionic?
   a. SO₂  b. CaF₂  c. CF₄  d. N₂  e. H₂O

3. Which has the highest electronegativity?

4. Which of the following molecules has a net dipole moment?
   a. Br₂  b. PCl₃  c. SiCl₄  d. BCl₃  e. CO₂

5. Which ion has the smallest radius?
   a. Li⁺  b. Be²⁺  c. F⁻  d. Ca²⁺  e. O²⁻

6. Which of the following constitutes an isoelectronic series?
   a. Al, Si, P, S
   b. Be, Mg, Ca, Sr
   c. I, Br, Cl, F
   d. Na⁺, Mg²⁺, Al³⁺, Si⁴⁺
   e. C²⁺, C¹⁺, C⁰, C¹⁻, C²⁻
7. As the number of bonds between 2 carbon atoms increases
   a. the number of electrons between the two atoms decreases
   b. the bond length decreases
   c. the bond energy decreases
   d. b and c

8. Which of the following exists as a resonance hybrid?
   a. NH₃  b. PCl₃  c. O₃  d. SO₃²⁻

9. Which of the following will have an expanded octet?
   a. IO₂⁻  b. I₂  c. ICl₂⁻  d. none of these

10. Which of the following has polar bonds but is a non-polar molecule?
    a. CCl₄  b. PCl₃  c. OF₂  d. both a and c

11. For the following structure in which nitrogen is the central atom, give a Lewis diagram,
    all resonance structures, and indicate the best resonance structure. The structure is linear.

    CNO¹⁻

12. In the Lewis structure for N₂, there exists
    a. a single bond between the nitrogens
    b. a double bond between the nitrogens
    c. a triple bond between the nitrogens
    d. three unpaired electrons overall
In this unit Dr. Etheridge will

1. describe covalent bonds and the types of overlap that form these bonds.

2. using VSEPR theory, predict electronic geometry, actual shape, and, when present, the direction of net dipole moments that exist for molecules or radicals.

3. discuss valence bond theory and apply this theory to simple molecules.

4. predict hybridizations for molecular structures.

5. give a molecular orbital potential energy diagram for both homonuclear and heteronuclear diatomic species.

6. produce the molecular orbital electron configuration for given diatomic species.

7. figure the bond order of given homonuclear or heteronuclear species.

8. predict diamagnetism or paramagnetism from molecular orbital diagrams.
During this unit we will discuss the following:

The Covalent Bond
VSEPR Theory
The Role of Carbon
Valence Bond Theory
Molecular Orbital Theory

Lesson 35:

THE COVALENT BOND

How are covalent bonds formed?

There are two types of bonds formed, $\delta$ (sigma) and $\pi$ (pi).

Define the term “sigma bond”:

Sketch and characterize a simple, sigma bond formed from overlap of two s-orbitals, showing how the bond works to hold two atoms together.

Why don’t the nuclei collapse onto the electron cloud of the bond?
Sketch and characterize end-to-end overlap of two p-orbitals.

Sketch and characterize overlap between an s-orbital and a p-orbital.

(We shall cover formation of pi bonds a little later)

**VSEPR THEORY**

Allows us to predict the shape of polyatomic molecules.
Deals with the spatial arrangement of volumes occupied by electrons.

**Rules for determining molecular geometry (spatial arrangement of electron volumes):**

1. Select the central atom. How is this done?

2. Determine the number of electrons about the **central** atom.
   How is this done? Be careful with sulfur and oxygen.

   examples:

   \[ \text{CCl}_4 \]

   \[ \text{NO}_2^- \]

   \[ \text{SO}_3 \]

3. Divide the result of rule 2 by 2 and select the appropriate geometry from the following table.
What does dividing by 2 permit us to determine?

<table>
<thead>
<tr>
<th># of e^- areas</th>
<th>electron density (geometry)</th>
<th>bond angles</th>
<th>hybridization</th>
</tr>
</thead>
<tbody>
<tr>
<td>2</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>3</td>
<td></td>
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<td>4</td>
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<tr>
<td>5</td>
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<td></td>
</tr>
<tr>
<td>6</td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

We will use examples below and on the following pages to fill in this table.

2 areas of high electron density

Consider BeBr₂

The geometry is _____________ . Bond angles are ___________.

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3 areas of high electron density

Consider BBr₃

The geometry is ___________________. Bond angles are ________

(The sketch on the video doesn’t show the lone pair electrons on the top bromine. Please put them in.)

Two examples for your consideration:

CO₂

SO₃
(pause the video and figure this one)
4 areas of high electron density

Consider CCl₄

The geometry is ________________________. Bond angles are _________. (This is chemistry’s most frequently occurring geometry.)

Now, consider H₂O and determine its “actual shape.” Actual shape refers to the geometry of the atomic nuclei, not the geometry which includes the lone-pair electrons.

Lesson 36:

Consider NH₃. It is another example of 4 areas of high electron density. Draw the molecule and note the location of the fourth pair of electrons. Then, note the shape of the molecule when the fourth pair is not considered, i.e. the actual shape of the molecule.
5 areas of high electron density

Consider PCl$_5$

The geometry is ________________________________.

The bond angles are ____________, ____________, and ____________.
(Please give all three angles in the table)

Redraw the structure and note the axial positions and the equatorial positions.

Try this one: ICl$_3$

Give the three possible structures.
List the three types of 90° repulsions as found in the ICl₃ molecule.

Based on these 90° repulsions, which of the three possible structures is best? Encircle it.

Draw the actual shape of the “best” structure.

Repeat the exercise using the ICl₄⁺ cation.
(Pause the video and try this exercise)

Note the number of 90° repulsions found of each type for each structure.

Give the actual shape for the cation.
6 areas of high electron density

Consider $\text{SCl}_6$

The geometry is ____________________________. Bond angles are ________.

Consider $\text{XeCl}_4$
(Pause the video and try this structure)

The geometry is __________________________ and the bond angles are _____________.

Determine the best structure from the above. Draw and describe the actual shape of this molecule.
Let’s return to $\text{ICl}_4^+$

The geometry is ________________________.

Give the possible arrangements (sketch them) using the accepted method of a “flying wedge” to indicate a bond or electron pair coming out of the page towards you and a dotted line or dotted wedge to indicate a bond or electron pair behind the page.

---

**Lesson 37:**

**THE ROLE OF CARBON**

List some of the characteristics of carbon bonding:
Consider this structure:

To determine the geometry about a carbon, consider the number of atoms to which each carbon is directly bonded.

When a carbon is directly bonded to

2 other atoms, the geometry is _______________________
3 other atoms, the geometry is _______________________
4 other atoms, the geometry is _______________________

**VALENCE BOND THEORY:** (refer to the table on page 180, the last column)

Consider BeCl₂

The electron configuration of Be:  1s²  2s²

Describe the hybridization formed:

How many sp hybrid orbitals does Be form? _________

Sketch this sp hybrid orbital.
Now, let’s consider the other element in that molecule:

The electron configuration for Cl: \([\text{Ne}] \ 3s^2 \ 3p_x^2 \ 3p_y^2 \ 3p_z^1\)

Sketch the molecule and note the bond and overlap forming the bond.

The type bond formed is ________________________________

**FILL IN THE CHART**

Consider \(\text{BCl}_3\)

B: \([\text{He}] \ 2s^2 \ 2p_x^1\)

Show the formation of the hybrid.

The hybrid is called a ______ hybrid. How many hybrids are produced? ______?

The angles of the hybrids are ________.

Sketch the molecule, noting the bonds and overlaps producing the bonds:
(Each of the bonds formed is the overlap of an sp\(^2\) (hybrid) to an unhybridized p.)
Consider CH₄

C  1s² 2s² 2px¹  pₓ¹

Show the expected hybridization. What is it called? ___________________

Try to sketch the molecule and show the overlap forming each of the bonds.

FILL IN THE CHART.

Consider NH₃

Electron configuration of N:

Show hybridization:
(Dr. Etheridge misspoke herself. It forms 4 sp³ hybrid orbitals, not 3 sp³ hybrid orbitals.)

Describe the location of the lone pair electrons.
Consider PCl$_5$

The electron configuration of phosphorus is:

Show the hybridization:

The hybrid produced is called the ________________.

How many of the hybrids exist? ______________

Sketch the molecule, indicating location of the lone pairs and the bond pairs. Note the bond formed and the overlap producing each of the bonds.

FILL IN THE CHART.

Consider ICl$_4^+$

\[ I = [Kr] \ 4d^{10} \ 5s^2 \ 5p_x^2 \ p_y^2 \ p_z^1 \]

Show the hybridization. Note that the “d” being referred to in the dsp$^3$ is the “d” from the 5$^{th}$ energy level, not the 4$^{th}$:

Sketch the molecule, give particular attention to the location of the lone pair, and FILL IN THE CHART.
Consider \( \text{SCl}_6 \)

The electron configuration for S is: \([\text{Ne}]\ 3s^2\ 3p_x^2\ p_y^1\ p_z^1\)

Show the hybrid formed:

Be sure you understand where, spatially, these orbitals are located.

**The pi (\(\pi\)) bond.**

How is the pi bond formed?

Show how the pi bond can result in attraction between two atoms.

What comprises a double bond? ___________________________________

What comprises a triple bond? ___________________________________
MOLECULAR ORBITAL THEORY

Review: A sigma bond forms from:

Describe the two types of bonds formed from $p$ atomic orbitals

MO theory considers energy relationships. Recall, when a bond is formed, energy is ________________, but when a bond is broken, energy is _________________.

How does the energy of the molecule compare to the energy of the respective atoms?

Draw the energy diagram for the hydrogen diatomic molecule, $H_2$

Is the molecule reasonable? ________________ Why?
Consider the Helium diatomic molecule, He$_2$

Is the He$_2$ molecule reasonable? Why or why not?

Consider H – He molecule. Is it reasonable? Explain:
**Bond Order:**

Define the term, “bond order”

The higher the bond order, the __________________________ the bond.

How is bond order calculated?

Calculate the bond order for the H-He structure:

Consider the diatomic lithium molecule, $\text{Li}_2$

\[ \begin{array}{c}
\text{energy} \\
\uparrow \\
\text{What is the bond order for this molecule?} \quad \underline{} \\
\text{This is an important question: Why isn’t Li}_2 \text{ listed as one of the diatomic molecules?}
\end{array} \]
Consider the diatomic beryllium molecule, Be₂

(Pause the video and place the orbitals in their proper energy positions.)

\[
\text{energy}
\]

What is the bond order? _________

We need to note the kinds of bonds can be formed when all three of the orientations of the \( p \) orbitals overlap. Diagram this overlap.

When three \( p \) orbitals overlap, ____________ pi and __________ sigma bonds form.

Describe the relative energies of the three overlaps.
Review the last part of the previous lesson to recall the type of overlap formed when all three orientations of the $p$ orbital overlap.

Now, consider the diatomic boron molecule, $B_2$ (The 1s is done for you)

Is $B_2$ reasonable? ________

Calculate its bond order ______________

Is it paramagnetic or diamagnetic? __________________________
Is there a carbon diatomic molecule, $C_2$? Prove it.

Determine the bond order? ____________

Is it paramagnetic or diamagnetic? _______________

Consider the diatomic nitrogen molecule, $N_2$, and calculate its bond order.
Consider diatomic oxygen, O₂. It offers an interesting situation.

What is the bond order? ___________

Is it paramagnetic or diamagnetic? ___________________

Sketch the π* (pi antibond)

Placing electrons in a pi antibonding orbital results in the relative energies of the sigma and pi bonds for that energy level to reverse their orders of energy.
Consider the case for diatomic fluorine, $\text{F}_2$, and be sure you get the proper bond order.

Now for $\text{Ne}_2$ (Can you predict it before you even begin the diagram?)
It’s time to try a heteronuclear diatomic, \( \text{CO} \)
(Pause the video and set up the atomic orbital energy levels.)
Do not be concerned because each atom is contributing a different number of electrons.

Try the superoxide ion, \( \text{O}_2^- \) (Note the differing number of electrons on each oxygen)

\[ \text{A shorthand notation:} \quad \text{Give the shorthand notation for } \text{O}_2 \]
(You may wish to return to the oxygen molecule energy chart to do this)
1. Here is a problem for you to consider. Is it possible for PCl$_5$ to form? Explain your answer using Valence Bond Theory.

2. If PCl$_5$ can form, can NCl$_5$ also be expected to form? (Explain your answer)
3. Give the actual shape (consider only nuclei) for each of the following:

   a. H₂O

   b. CO₂

   c. SF₄

   d. ICl₄⁺

4. For the following molecule, designate the bond as sigma or pi and describe the overlap forming the bond (ex. overlap between an sp² – s)
During this unit, Dr. Etheridge will

1. review the oxidation numbers given on the handout early in the course.

2. explain what is meant by oxidation and reduction, and the terms oxidizing agents, and reducing agents.

3. demonstrate how to assign oxidation numbers to each element in a compound or radical.

4. balance equations using the “half-cell method” which includes completing and balancing each individual half-cell, both oxidation and reduction, regardless of whether it occurs in acidic, basic, or neutral solution.

5. demonstrate a quick method of balancing equations using change of oxidation numbers.

6. demonstrate a good way to manage disproportionation reactions.

7. discuss spectator ions and their role in oxidation-reduction reactions.

8. go through calculations from oxidation-reduction reactions that you have already balanced.

9. demonstrate how to predict whether elements, ions, radicals, and compounds would behave as oxidizing agents or reducing agents or both.

10. show how to convert balanced net ionic equations to complete molecular equations.

11. demonstrate predicting whether diatomic oxygen will produce H₂O or OH⁻ in a basic solution.
During this unit we will discuss the following:

- Definitions
- Oxidation Numbers
- The Half-Cell Method
- The Oxidation Number Method
- Stoichiometry
- Oxidizing-Reducing Agents

Lesson 40:

*****About 4 minutes into the video, the unit stops and restarts. Apparently the college made a technical error and restarted the video. All is well in the new beginning.*****

DEFINITIONS

The old definitions says

Oxidation occurs when a substance _____________ oxygen.

Reduction occurs when a substance _____________ oxygen.

However, consider these equations as we develop a broader concept:

\[
\text{CuO} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{O}
\]

\[
\text{CuS} + \text{H}_2 \rightarrow \text{Cu} + \text{H}_2\text{S}
\]
A substance which gains electrons is ________________.

A substance which loses electrons is ________________.

Oxidation is the _______________ of electrons.

Reduction is the _______________ of electrons.

The oxidizing agent ________________ electrons.

The reducing agent ________________ electrons.

ASSIGNING OXIDATION NUMBERS

For neutral molecules, the sum of oxidation numbers is _______.

For ions/radicals, the sum of oxidation numbers is equal to _________________.

More positive elements are generally written on the _________________.

More negative elements are generally written on the _________________.

Look at KMnO$_4$ and let’s assign each element its oxidation number

\[ \text{KMnO}_4 \]

Because Mn is the “middle element,” its oxidation state will adjust within reason to accommodate the other two elements. Is +7 reasonable for Mn? Why?

Now, consider these:

\[ \text{AsO}_4^{-3} \]

\[ \text{CuCr}_2\text{O}_7 \]

Pause this video and look carefully at Cr’s position on the periodic table. Describe here how you would justify its oxidation number from its position on the periodic table.
Consider the following net ionic equation. Determine what is oxidized and what is reduced:

\[ \text{Fe}^{+2} + \text{MnO}_4^- + \text{H}^+ \rightarrow \text{Fe}^{+3} + \text{Mn}^{+2} + \text{H}_2\text{O} \]

_______________ loses electrons.
_______________ gains electrons.

The reducing agent is ________________

The oxidizing agent is ________________

THE HALF-CELL METHOD

An oxidation-reduction reaction is composed of 2 half-reactions. These two parts may occur independently, provided a method is provided for transport of electrons. Sketch a lab set-up in which the two half-cells occur independently.

Sketch a typical dry cell battery and explain briefly how it works:
Remember:

\textbf{electrons lost = electrons gained}

The three fundamental steps for balancing oxidation-reduction reactions via the half-cell method are:

1. Divide the reaction into oxidation and reduction half-cells.

2. Complete and balance each half-cell (We will expand this considerably, below).

3. Add the completed and balanced half-cells as you would equations in algebra.

Use the half-cell method to balance the following reaction. The expanded instructions appear below:

\[ \text{Cu} + \text{NO}_3^- \rightarrow \text{Cu}^{+2} + \text{NO} \quad \text{(acid solution)} \]

\[ \text{Here are the specific steps needed to balance the equation:} \]

1. Write the substance being oxidized (or reduced) and its product.

2. Note electrons lost on the right and electrons gained on the left.

3. Balance charges using H\(^+\) in acid, OH\(^-\) in base, and whatever ions are unchanged in neutral solutions.

4. Balance oxygen using water. (Repeat with the other half-cell.)

5. Make electrons lost equal to electrons gained.

6. Add the half-cells algebraically.
Let’s go through this equation. Note: It occurs in a basic solution instead of an acidic one.

\[ S^{−} + MnO_4^{−} \rightarrow S + MnO_2 \]  (base solution)

GO TO DISC 11

Lesson 41:

Pause the video and try balancing this equation. Then continue the video and check your answer.

\[ NO_2^{−} + MnO_4^{−} \rightarrow NO_3^{−} + MnO_2 \]  (base solution)
This next reaction simultaneously is both a disproportionation reaction and one which occurs in a neutral solution.

First, what is a “disproportionation” reaction?

It is not unusual to see the halogens involved in these reactions, as well as ions/radicals in which an element takes an intermediate oxidation number. This statement will make more sense in the latter part of this unit.

\[ \text{Br}_2 + \text{CO}_3^{2-} \rightarrow \text{Br}^- + \text{BrO}_3^- + \text{CO}_2 \] (neutral solution)

What is another name for a disproportionation reaction?

**THE OXIDATION NUMBER METHOD**

Unlike the half-cell method, the oxidation number method utilizes the entire equation at once.

\[ \text{Cu} + \text{HNO}_3 \rightarrow \text{Cu(NO}_3)_2 + \text{NO} + \text{H}_2\text{O} \]
Try this one:

\[
\text{Zn} + \text{NO}_3^- + \text{OH}^- + \text{H}_2\text{O} \rightarrow \text{NH}_3 + \text{Zn(OH)}_4^- 
\]

There is a neat trick to use whenever nitrogen and hydrogen are bonded: Assign N a charge of ______ and assign H a charge of ______.

It is important to remember that the net charge on the left must equal the net charge on the right. Otherwise, there are electrons moving that have not been documented.

**STOICHIOMETRY**

It is a two-step process:

1. Complete and balance the equation.

2. Perform the calculations.

Calculate the volume of 0.0199 M sodium oxalate needed to react with 25.00 mL of 0.0233 M acid potassium permanganate solution. Carbon dioxide and the manganous ion are produced.

[Do you recall the definition of spectator ions? They are those ions that are not directly involved in the reaction, but whose presence is necessary for compounds to be complete. They just sit and “watch” the reaction occur. Potassium ions and sodium ions are spectator ions in this reaction.]
Here’s the “quick” method: IT DOESN’T ALWAYS WORK!

Lesson 42:

DETERMINING WHAT CAN BE OXIDIZING AND REDUCING AGENTS

Review of Range of Oxidation States (fill these in)
(If you don’t already have these marked on your periodic table, you should mark them now.)

IA = 0 and +1
IIA = 0 and +2
IIIA = ______________
IVA = ______________
VA = ______________
VIA = ______________
VIIA = ______________

Review the ranges for the elements in transition groups

The scandium group ______________
The titanium group ______________
The vanadium group ______________
The chromium group ______________
The manganese group ______________
Recall:

Oxidizing agents _____________ electrons.
If it cannot gain electrons and stay within its range of oxidation numbers, it cannot be an oxidizing agent!

Reducing agents _____________ electrons.
If it cannot lose electrons and stay within its range of oxidation numbers, it cannot be a reducing agent!

What can be an oxidizing agent?

What can be a reducing agent?

What about O₂ as an oxidizing agent? Explain.

Can the S⁻ act as an oxidizing agent? Explain.

Can MnO₄⁻ act as an oxidizing agent?
(In these radicals, keep oxygen at a -2 charge for our purposes)

What part of the permanganate radical behaves as an oxidizing agent?

Can O₂ act as a reducing agent? Elaborate on this.

Can Na (metal) act as a reducing agent? Explain.

METALS AT THEIR ZERO OXIDATION STATE ARE WONDERFUL REDUCING AGENTS!

Can C₂O₄⁻ act as a reducing agent? ___________________________
Make note of the several reducing agents mentioned here and if you are not absolutely clear on WHY they are good, then make note of what Dr. Etheridge says about them:

Make note of the several oxidizing agents mentioned here and if you are not absolutely clear on WHY they are good, then make note of what Dr. Etheridge says about them:

Pay very careful attention to these substances which can act as both oxidizing and reducing agents and WHY!
Here is an exercise for practice: (You may need to pause the video and try your hand at some of these as you go. Don’t let Dr. Etheridge give you all the information before you have an opportunity to try for yourself.) Some of these blanks may have multiple answers.

<table>
<thead>
<tr>
<th>Substance</th>
<th>Produces as an Oxidizing Agent</th>
<th>Produces as a Reducing Agent</th>
</tr>
</thead>
<tbody>
<tr>
<td>ClO₄⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>H₂O₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>MnO₂</td>
<td></td>
<td></td>
</tr>
<tr>
<td>NO₃⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>IO₃⁻</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Fe³⁺</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Zn</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

**Molecular Equations:**

Give a complete, balanced, molecular equation when KMnO₄ reacts with H₂S in the presence of HCl. Mn²⁺ and S⁰ are formed. Account for all spectator ions.
1. Chromium in $\text{K}_2\text{Cr}_2\text{O}_7$ has an oxidation number of
   a. +2  b. +3  c. +4  d. 7

2. Which of the following cannot act as an oxidizing agent?
   a. $\text{H}_2$  b. $\text{I}^-$  c. $\text{NO}_2^-$  d. $\text{O}_2$

3. How many electrons are gained in the following reaction?
   \[ \text{SO}_3^{2-} + \text{MnO}_4^{1-} \rightarrow \text{SO}_4^{2-} + \text{Mn}^{2+} \]

4. How many electrons are gained in the reaction in question #3?

5. The following reaction occurs in aqueous NaOH
   \[ \text{KMnO}_4 + \text{NaC}_2\text{O}_4 \rightarrow \text{MnO}_2 + \text{CO}_3^{2-} \]
   Show each balanced half-cell and the balanced net ionic equation. Then convert the net ionic equation into a balanced molecular equation.
APPENDIX

Periodic Table 221
Activity Series 222
Common Oxidation Numbers 223
Answers to Quick Quizzes 224
## Periodic Table of the Elements

<table>
<thead>
<tr>
<th>1A</th>
<th>2A</th>
<th>3A</th>
<th>4A</th>
<th>5A</th>
<th>6A</th>
<th>7A</th>
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</thead>
<tbody>
<tr>
<td>H</td>
<td>He</td>
<td>B</td>
<td>C</td>
<td>N</td>
<td>O</td>
<td>F</td>
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<tr>
<td>Li</td>
<td>Be</td>
<td>N</td>
<td>O</td>
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<td>Ne</td>
<td>Ne</td>
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<td>Mg</td>
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<td>Si</td>
<td>P</td>
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<td>Ti</td>
<td>V</td>
<td>Cr</td>
<td>Mn</td>
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<td>Rb</td>
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<td>Y</td>
<td>Zr</td>
<td>Nb</td>
<td>Mo</td>
<td>Tc</td>
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<td>Ba</td>
<td>La</td>
<td>Hf</td>
<td>Ta</td>
<td>W</td>
<td>Re</td>
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<tr>
<td>Fr</td>
<td>Ra</td>
<td>Ac</td>
<td>Rf</td>
<td>Db</td>
<td>Sg</td>
<td>Bh</td>
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### Lanthanides

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<tr>
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<th>Ce</th>
<th>Pr</th>
<th>Nd</th>
<th>Pm</th>
<th>Sm</th>
<th>Eu</th>
<th>Gd</th>
<th>Tb</th>
<th>Dy</th>
<th>Ho</th>
<th>Er</th>
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</table>

### Actinides

<table>
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<tr>
<th>Th</th>
<th>Pa</th>
<th>U</th>
<th>Np</th>
<th>Pu</th>
<th>Am</th>
<th>Cm</th>
<th>Bk</th>
<th>Cf</th>
<th>Es</th>
<th>Fm</th>
<th>Md</th>
</tr>
</thead>
</table>

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**Note:** The image is a periodic table of the elements with classifications for metals, metalloids, and non-metals, along with the lanthanides and actinides.
ACTIVITY SERIES OF THE ELEMENTS

K
Ba
Sr
Ca
Na
Mg
Al
Mn
Zn
Cr
Fe
Cd
Co
Ni
Sn
Pb
H
Sb
As
Bi
Cu
Ag
Pd
Hg
Pt
Au
### Table of Common Oxidation Numbers

<table>
<thead>
<tr>
<th>1+ elements</th>
<th>2+ elements</th>
<th>3+ elements</th>
<th>4+ elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>H⁺</td>
<td>Mg⁺</td>
<td>Al⁺⁺⁺</td>
<td>Pb⁺⁺⁺⁺ (plumbic)</td>
</tr>
<tr>
<td>Li⁺</td>
<td>Ca⁺⁺</td>
<td>Bi⁺⁺</td>
<td>Sn⁺⁺⁺⁺ (stannic)</td>
</tr>
<tr>
<td>Na⁺</td>
<td>S⁺⁺</td>
<td>Cr⁺⁺⁺⁺ (chromic)</td>
<td></td>
</tr>
<tr>
<td>K⁺</td>
<td>Zn⁺⁺</td>
<td>Fe⁺⁺⁺ (ferric)</td>
<td></td>
</tr>
<tr>
<td>Cs⁺</td>
<td>Mg⁺⁺⁺⁺ (manganous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Ag⁺</td>
<td>Cr⁺⁺⁺⁺ (chromous)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Cu⁺⁺⁺ (cuprous)</td>
<td>Mg⁺⁺⁺⁺ (manganous)</td>
<td></td>
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</tr>
<tr>
<td>Hg⁺⁺⁺⁺⁺ (mercurous)</td>
<td>Fe⁺⁺⁺⁺⁺ (ferrous)</td>
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<td></td>
</tr>
<tr>
<td>1+ radicals*</td>
<td>2+ radicals*</td>
<td>3+ radicals*</td>
<td></td>
</tr>
<tr>
<td>H₃O⁺⁺ (hydronium)</td>
<td>Cu⁺⁺⁺⁺⁺ (cupric)</td>
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<td></td>
</tr>
<tr>
<td>NH₄⁺⁺ (ammonium)</td>
<td>Hg⁺⁺⁺⁺⁺ (mercuric)</td>
<td></td>
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</tbody>
</table>

<table>
<thead>
<tr>
<th>1- elements</th>
<th>2- elements</th>
<th>3- elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>H⁻ (hydride)</td>
<td>O⁻⁻</td>
<td>N⁻⁻⁻⁻⁻⁻ (nitrate)</td>
</tr>
<tr>
<td>F⁻</td>
<td>S⁻⁻</td>
<td>H⁻⁻⁻⁻⁻⁻ (thiosulfate)</td>
</tr>
<tr>
<td>Cl⁻</td>
<td>P⁻⁻</td>
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<tr>
<td>Br⁻</td>
<td>1- radicals*</td>
<td>1- radicals (cont)</td>
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<tr>
<td>I⁻</td>
<td>C₂H₃O₂⁻ (acetate)</td>
<td>HSO₄⁻ (bisulfate)</td>
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<td>1- radicals (cont)</td>
<td>2- radicals*</td>
<td>3- radicals*</td>
</tr>
<tr>
<td>CN⁻ (cyanide)</td>
<td>SO₄⁻⁻⁻⁻⁻⁻ (sulfate)</td>
<td>BO₃⁻⁻⁻⁻⁻⁻ (borate)</td>
</tr>
<tr>
<td>NO₃⁻ (nitrate)</td>
<td>SO₃⁻⁻⁻⁻⁻⁻ (sulfite)</td>
<td>PO₄⁻⁻⁻⁻⁻⁻ (phosphate)</td>
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<td>NO₂⁻ (nitrite)</td>
<td>S₂O₃⁻⁻⁻⁻⁻⁻ (thiosulfate)</td>
<td>PO₃⁻⁻⁻⁻⁻⁻ (phosphite)</td>
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<tr>
<td>(arsenate)</td>
<td>CO₃⁻⁻⁻⁻⁻⁻ (carbonate)</td>
<td>AsO₃⁻⁻⁻⁻⁻⁻ (arsenate)</td>
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<td>OH⁻ (hydroxide)</td>
<td>(hydroxide)</td>
<td>2- radicals*</td>
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<td>SCN⁻ (thiocyanate)</td>
<td>(thiocyanate)</td>
<td>C₂O₄⁻⁻⁻⁻⁻⁻ (oxalate)</td>
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<tr>
<td>ClO⁻ (hypochlorite)</td>
<td>O₂⁻⁻⁻⁻⁻⁻ (peroxide)</td>
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</tr>
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<td>ClO₂⁻ (chlorite)</td>
<td>SiO₃⁻⁻⁻⁻⁻⁻ (silicate)</td>
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<tr>
<td>ClO₃⁻ (chlorate)</td>
<td>CrO₄⁻⁻⁻⁻⁻⁻ (chromate)</td>
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<tr>
<td>ClO₄⁻ (perchlorate)</td>
<td>Cr₂O₇⁻⁻⁻⁻⁻⁻ (dichromate)</td>
<td></td>
</tr>
<tr>
<td>BrO⁻ (bromate)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>BrO₂⁻ (perbromate)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>MnO₄⁻ (permanganate)</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

* The term “radicals” is an older term for polyatomic ions.
Answers

Quiz 1
1. b   2. d   3. b   4. a   5. d   6. a   7. b

Quiz 2
1. $1.98 \times 10^{24}$ atoms   2. 209   3. 80.918   4. b

5. potassium dichromate
   manganese dioxide
   calcium carbonate
   sodium peroxide
   sodium arsenate, monobasic
   cobaltous fluoride or cobalt (II) fluoride
   hypochlorous acid
   ammonium phosphide

6. LiAlH$_4$    Zn(HSO$_4$)$_2$    K$_2$HPO$_4$    H$_2$C$_2$O$_2$.2H$_2$O
   N$_2$O$_5$    As(SCN)$_5$    Ca(IO)$_2$    Ca(OH)$_2$

7. $8.8 \times 10^{23}$ atoms

8. Multivalent metals are metals having more than one common oxidation number, i.e. Fe$^{2+}$ and Fe$^{3+}$.

9. $7.5 \times 10^{22}$ lbs (that is certainly a lot of eggs)

10. The atomic mass number only considers the number of nucleons, each of which weighs slightly more than 1.
Quiz 3

1. $2\ C + O_2 \rightarrow 2\ CO$
2. $2\ CO + O_2 \rightarrow 2\ CO_2$
3. $CaO + H_2O \rightarrow Ca(OH)_2$
4. $2\ Al + 6\ HCl \rightarrow 2\ AlCl_3 + 3\ H_2$
5. $CaCO_3 + \Delta \rightarrow CaO + CO_2$
6. $Mg + Br_2 \rightarrow MgBr_2$
7. $2\ KBrO_3 + \Delta \rightarrow 2\ KBr + 2\ O_2$
8. $Fe + S$ (or $S_8$ may be used) $\rightarrow$ FeS (Or $Fe_2S_3$ could be produced)
9. $2\ C_4H_{10} + 9\ O_2 \rightarrow 8\ CO + 10\ H_2O$ (limited oxygen forms CO)
10. $2\ C_4H_{10} + 13\ O_2 \rightarrow 8\ CO_2 + 10\ H_2O$
11. $SO_2 + H_2O \rightarrow H_2SO_3$
12. $2\ Mg + O_2 \rightarrow 2\ MgO;\ MgO + H_2O \rightarrow Mg(OH)_2$
13. no reaction
14. $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$
15. $2\ NaHCO_3 + \Delta \rightarrow Na_2CO_3 + H_2O + CO_2$
16. no reaction
17. no reaction
18. $2\ KNO_3 + \Delta \rightarrow 2\ KNO_2 + O_2$
19. $Cr_2(CO_3)_3 + \Delta \rightarrow Cr_2O_3 + 3\ CO_2$
20. $2\ H_2O_2 \rightarrow 2\ H_2O + O_2$
21. no reaction
22. $P_2O_5 + 3\ H_2O \rightarrow 2\ H_3PO_4$
23. \( \text{Br}_2\text{O} + \text{H}_2\text{O} \rightarrow 2\text{HBrO} \)
24. \( \text{Fe}_2\text{O}_3 + 3\text{H}_2\text{O} \rightarrow 2\text{Fe(OH)}_3 \)
25. \( \text{CuO} + \text{H}_2\text{O} \rightarrow \text{Cu(OH)}_2 \)
26. \( \text{CaO} + \text{SO}_3 \rightarrow \text{CaSO}_4 \)
27. \( 2\text{NaClO}_3 + \Delta \rightarrow 2\text{NaCl} + 3\text{O}_2 \)
28. \( \text{Mg} + \text{CuSO}_4 \rightarrow \text{MgSO}_4 + \text{Cu} \)
29. \( \text{Ca} + \text{CuSO}_4 \rightarrow \text{CaSO}_4 + \text{Cu} \)
30. \( \text{Mg} + \text{HCl} \rightarrow \text{MgCl}_2 + \text{H}_2 \)
31. \( \text{Ca} + \text{H}_2 \rightarrow \text{Ca(OH)}_2 + \text{H}_2 \)
32. \( 2\text{Na}_2\text{O}_2 + 2\text{H}_2\text{O} \rightarrow 4\text{NaOH} + \text{O}_2 \)
33. \( \text{Cl}_2\text{O}_5 + \text{H}_2\text{O} \rightarrow 2\text{HClO}_3 \)
34. \( \text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2 \)
35. \( \text{Br}_2\text{O}_3 + \text{H}_2\text{O} \rightarrow \text{HBrO}_2 \)

**Unit 4**

PbCl\(_2\) insol  BaSO\(_4\) insol  
Hg\(_2\)Br\(_2\) insol  Zn(C\(_2\)H\(_3\)O\(_2\))\(_2\) sol  
NaI sol  (NH\(_4\))\(_3\)PO\(_4\) sol  
FeCl\(_3\) sol  Na\(_3\)AsO\(_3\) sol  
PbI\(_2\) insol  BaCO\(_3\) insol  
AgBr insol  
MgSO\(_4\) sol
Unit 4 (continued)

2. The water sample contained $4 \times 10^{-2}$ ppm antimony, less than the MCLG.

3. After the reaction, the following amounts of each substance remained:

   \[
   \begin{align*}
   H^+ &= 0, \quad Cl^- = 0, \quad Pb^{2+} = 5.4 \text{ mmol}, \quad C_2H_3O_2^- = 10.8 \text{ mmol} \\
   PbCl_2 &= 2.4 \text{ mmol precipitate}, \quad HC_2H_3O_2 = 4.8 \text{ mmol (weak acid)}
   \end{align*}
   \]

The concentrations of all substances remaining in solution are:

   \[
   \begin{align*}
   Pb^{2+} &= 0.195 \text{ M}, \quad C_2H_3O_2^- = 0.39 \text{ M}, \quad HC_2H_3O_2 = 0.17 \text{ M}
   \end{align*}
   \]

The quantity of precipitate remaining is $PbCl_2 = 0.67 \text{ g}$

Remember, the total volume of solution is 27.6 mL

Unit 5

1. $1.4 \times 10^{23}$ oxygen molecules
2. The same number. No molecules were lost or gained.
3. Nothing. The total pressure of the mixture will increase, but the partial pressures of each gas present remains unchanged.
4. 1.24 L
5. 1.38 L/min
6. low temperature and high pressure
7. Helium. It diffuses the fastest.
8. 0.976
9. 2.7 atm
10. 350 mL

Unit 6

1. d
2. a
3. d
4. d
5. c
6. b
7. c
8. $1.09 \times 10^4 \text{ kJ}$ for the two moles of butane involved
9. -137 kJ
Unit 7
1. c  2. c  3. d  4. b  5. d

Unit 8
1. [Ar] 4s\(^1\)3d\(^{10}\); 3 2 2 ½  2. no
3. Periodic Law stated that properties of elements were periodic functions of their atomic weight. However, when the concept of atomic number was developed, the definition changed to “properties of elements are periodic functions of their atomic number.” This eliminated the problems which occurred when accurate determinations of atomic weights showed that progression of atomic weights on the periodic table are inconsistent.
4. b  5. 7
6. 1, no, hydrogen cannot be an alkali metal because it is a non-metal
7. 2, no, helium cannot be an alkaline earth because it is not a metal
8. Na<Li<C<O<F  9. b  10. c
11. 3.86 X 10\(^{-7}\) M or 386 nM  12. a. 8  b. 2  c. impossible set  d. 1

Unit 9
1. e  2. b  3. c  4. b  5. b  6. d
7. b  8. c  9. c  10. a
11. :\(\text{C} = \text{N} = \text{O}\):  :\(\text{C} = \text{N} = \text{O}\):  :\(\text{C} = \text{N} = \text{O}\):  best  12. c

\[\begin{array}{ccc}
-3 & +1 & +1 \\
-2 & +1 & 0 \\
-1 & +1 & -1 \\
\end{array}\]
Unit 10

1. Yes. The hybridization for phosphorus is the dsp3 which phosphorus can manage.

2. No. Nitrogen would have to form a dsp3, also, but nitrogen’s highest energy level is the second level; a level that does not contain the potential for having a “d.”

3. a. bent    b. linear    c. seesaw    d. seesaw

4. 

\[
\begin{array}{c}
\text{H} \\
\text{1} \\
\text{C} \\
\text{2} \\
\text{H} \\
\text{H} \\
\text{1} \\
\end{array}
\quad
\begin{array}{c}
\text{C} \\
\text{3} \\
\text{H} \\
\text{C} \\
\text{H} \\
\text{H} \\
\text{1} \\
\end{array}
\quad
\begin{array}{c}
\text{H} \\
\text{C} \\
\text{4} \\
\text{H} \\
\text{O} \\
\text{H} \\
\text{7} \\
\end{array}
\]

1. sigma, s – sp2    2. double bond has sigma sp2 – sp2 and pi p-p
3. sigma sp2-sp3    4. sigma 1s-sp3    5. sigma sp3-sp3
6. sigma sp3-sp3    7. sp3-s

Unit 11

1. d    2. b    3. 2    4. 5