

## COURSE OUTLINE

<u>CHE 101</u>	<u>Chemistry</u>	<u>General Chemistry I</u>
Course Number	Science Division	Course Title
<u>4</u>	<u>3</u>	<u>3</u>
Credits	Class Hours/Week	Laboratory Hours/Week

### Required Materials:

#### General Chemistry:

Whitten, Kenneth, Davis, Raymond, Peck, and Stanley, Thompson-Brooks/Cole,  
9th Ed., 2010 with On-line Access

#### General Chemistry I: Course Manual

Alfare, Carlo MCCC, 8th Edition

#### General Chemistry I: Laboratory Manual

Alfare, Carlo MCCC, 10th Edition

Any basic calculator (scientific notation, log., and trig. functions recommended) may be used (Programmable calculators or computers will not be allowed for use on tests or in laboratories, even if memories are cleared.)

Goggles must be worn in the laboratory. They will be supplied or can be purchased in the Bookstore.

15 Weeks  
Length of Semester

Week 15  
Final Examination

### Catalog Description:

Basic concepts introduced. Theoretical principles of modern chemistry integrated with descriptive and practical aspects. Topics discussed include: stoichiometry, atomic theory and the structure of matter, Periodic Table, chemical bonding, kinetic-molecular theory and the states of matter; gas laws, solutions, oxidation-reduction, acid-base systems, and thermochemistry. Laboratory work illustrates selected topics covered in lecture.

Prerequisite: MAT 135 or higher

Corequisite: ENG 101 or higher

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**Grading Procedure:**

Grading will be based on the point system as indicated below.

<u>Activity</u>	<u>% of Total</u>	<u>Basis for Points</u>	<u>Max. Points</u>
Examination I	15.7	Percent	100
Examination II	15.7	Percent	100
Examination III	15.7	Percent	100
Quizzes	15.7	Total (best of 10)	100
Laboratory	21.6	1/2 of Total	138
Final Examination	15.7	Percent	<u>100</u>
TOTAL			638

**Minimum Course Grade Assignment:**\*

A	593 Points	(93%)	B -	510 Points	(80%)
A -	574 Points	(90%)	C +	490 Points	(77%)
B +	555 Points	(87%)	C	435 Points	(68%)
B	530 Points	(83%)	D	370 Points	(58%)

\*Acceptable laboratory and recitation participation and performance along with a passing grade on the final examination are required to pass the course. See the Course Objectives for more details.

**Quizzes and Examinations:**

Weekly quizzes will be given in recitation. No make-up will be given. There will be three hourly examinations given during the semester at approximately the intervals shown in the Topical Outline. Specific dates and locations for these examinations will be announced at least one week in advance. It is your responsibility to be present at all the examinations and the final exam. An unexcused absence will constitute a zero score on any exam or quiz. An absence will be considered unexcused if notification of the course coordinator is not made before hand, where possible, or within 48 hours of the absence.

See Appendix E for Grade Record Keeping Chart.

# TOPICAL OUTLINE

Week	Unit	Lecture & Recitation	Whitten	Lab. Experiments & "Films"	Self Study Materials
1	I	Introduction	1, App. A & C	1 An Introduction to Lab. "The Metric Film" "Powers of Ten"	Text CD – Chapter 1
2	II	Stoichiometry	2	2 Determination of Density	Text CD – Chapter 2 (18 & 19) Course Manual Appendix D: 1-4
3	II	Stoichiometry	3	3 Qualitative Analysis of a Hydrate Mixture	Text CD – Chapter 3 (16-19) Text CD – Chapter 5 Course Manual Appendix D: 5
4	III	Gases	12	4 Formula of a Compound "States of Matter"	Text CD – Chapter 12
5	III	Gases	12	5 Ideal Gas Laws	Text CD – Chapter 12 Course Manual Appendix D: 6
<b>EXAMINATION I                      WEEKS 1 – 5                      OBJECTIVES I - III</b>					
6	IV	Crystals	13	6 Molar Volume of Oxygen	Text CD – Chapter 13
7	V	Liquids and Changes of State	13	7 Crystal Structure "Anitmatter"	Text CD – Chapter 13
8	VI	Atomic Structure	4	8 Spectroscopy and Atomic Structure "The Periodic Table" "Nature of Matter"	Text CD – Chapter 2 & 7

Week	Unit	Lecture & Recitation	Whitten	Lab. Experiments & "Films"	Self Study Materials
9	VI	Periodic Properties	5	8 Spectroscopy and Atomic Structure	Text CD – Chapter 8 Course Manual: App. A "Naming Inorganic Compds."
<b>EXAMINATION II: Weeks 6-9 Objectives IV-VI</b>					
10	VII	Chemical Bonding	7	9 Concentration Units & Preparation of Solutions	Text CD – Chapter 3 & 9
11	VII	Chemical Bonding	8	10 Indicators and Acid-Base Titrations	Text CD – Chapter 10
12	VIII	Chemical Reactions in Aqueous Solutions	6, 10	10 Acid – Base Titrations	Text CD – Chapter 4
13	VIII	Chemical Rxns. in Aqueous Solutions	11	11 Oxidation-Reduction Workshop & Titration	Text CD – Chapter 4 & 21
<b>EXAMINATION III: Weeks 10-13 Objectives VII-VIII</b>					
14	IX	Properties of Solutions	14	12 Determination of Molecular Weight From Freezing Point Depression	Text CD – Chapter 14
15	IX	Properties of Solutions	14	12 Conclusion & Checkout	Text CD – Chapter 14
<b>FINAL EXAMINATION Weeks 1-15 Objectives I-IX</b>					

**HOMWORK SCHEDULE**

Homework assignments are on a weekly basis to help you learn the course material according to the performance objectives and to help you to test your mastery of the material considered. They are not to be considered "exclusive" but representative of the material. If you feel the need to do additional reading or problems, you can ask your instructor for guidance. All but the last column will be found in Whitten, Davis, Peck, and Stanley

**THE ONLY WAY TO LEARN THE MATERIAL IS TO DO IT YOURSELF. CHEMISTRY IS A PROBLEM SOLVING COURSE. YOU CAN ONLY MASTER IT WITH PRACTICE.**

<u>Week</u>	<u>Read Chap.</u>	<u>Questions and Problems</u>	<u>Course Manual</u>
1a	1	3-5, 8, 9, 14-18, 21-23, 29-32, 35, 36, 45(ans. $1.6 \times 10^4$ g), 48, 53	Unit 1: 1 - 14
1b	2	1, 2, 11, 18, 26, 40	Unit 2: 1
2a	2	28, 36, 44 (DO MANUAL FIRST)	Unit 2: 2 - 10
2b	2	54, 58, 62, 76, 78, 112(optional)	Unit 2: 11 - 13
3a	3	<b>DO THE MANUAL FIRST</b> 1, 2, 5, 8, 12, 16, 24	Unit 2: 14 - 16
3b	3	57(ans. 0.255 mole, 75.0 g), 62, 69(ans. 40.3 g), 70, 72	Unit 2: 15 (Optional)

**NOTE:** The programs described in Appendix D can be very helpful. If you notice ANY difficulties in the early part of the course, I recommend immediate recourse to these programs.

4a	12	1-4, 6, 12, 20, 22, 31, 32, 36, 60	Unit 3: 1 - 5
4b	12	14, 26, 34, 66, 67(ans. 32.2 mL), 111	
5a	12	37, 38, 44, 85, 87, 88, 91, 93, 96	Unit 3: 6 - 8
5b	12	40, 45(ans. 4.68 g/L), 46, 56, 98, 108	Unit 3: 9, 10
6a	13	75, 76, 84-88, 89(ans. $3.47 \times 10^{-22}$ g & $3.79 \times 10^{-23}$ cm <sup>3</sup> ), 90- 92, 96, 98, 99(ans. 3.53 Å), 100, 101(ans. 2.350 Å)	Unit 4: 1 - 21
6b	13	2, 4, 6, 8, 12, 15, 16	Unit 5: 1 - 5
7	13	21-24, 27, 28, 30, 31, 34, 39, 42, 45, 46, 65-70, 72, 120, 121	Unit 5: 6 - 8

<u>Week</u>	<u>Read Chap.</u>	<u>Questions and Problems</u>	<u>Course Manual</u>
8	4	1, 2, 4, 7, 9, 15, 16, 18, 20, 22, 24, 39, 40, 42, 49-52, 53a, 54a, 65-67, 69, 70, 79-92, 94, 105, 107-109, 114-117, 126-128	Unit 6: 1 - 13
9	5-1 to 5-6	1-11, 13-21, 24, 26-28, 30, 32, 36, 38, 41, 42	Unit 6: 14 - 16
10a	6-3, 6-4	27-32, 34, 36, 38, 40, 45, 46	Unit 7: 1 - 2
10b	7	1-16, 18-21, 23-25, 28, 30, 32, 34, 36, 52, 57, 59, 62a, 67-69, 71, 74	Unit 7: 3
10c	8	1-4, 7-10, 13-16, 20-22	Unit 7: 4 - 5
10d	8	24, 26, 30, 32, 34, 38, 46, 48-50, (include hybridization, orbital & molecular geometry in all questions)	Unit 7: 6 - 7
11a	6-1, 6-2, 6-9	2-12, 14, 18, 20	Unit 8: 1
11b	5-9	67 - 71	
11c	10	1-4, 6-13, 17-20, 23-27	Unit 8: 2
11d	10	31, 32, 34-36, 40, 41, 43-47, 50, 57, 61, 63	Unit 8: 3 - 6
12a	5-7	46, 48, 49	Unit 8: 7 - 9
12b	6-5	49, 50, 52	
12c	11	1-4, 6, 40, 41(also calc.N, ans. 4.56 N) 44, 47(ans. 0.07365 M)	Unit 8: 10 - 13
12d	11-6	63(ans. 10.7 mL), 64 (Eq. Wt. Method for both)	Unit 8: 14 - 17
13	11-4, 11-5	52a,b, 53b,c, 55a, 61d	Unit 8: 18 - 24
14a	14	1, 2, 4, 5, 7, 8, 10, 14, 22, 23, 32, 36, 38, 40, 43. 44, 91	Unit 9: 1 - 7
14b	14	46, 48, 50, 60, 68, 81	Unit 9: 8 - 12

**SUPPLEMENTARY MATERIALS**

CD is available on the computers in the Library and MS 211

Cengage on-line access

Tutoring sites on my home page: [www.mccc.edu/~calfare](http://www.mccc.edu/~calfare)

**OPTIONAL SUPPLEMENTARY MATERIALS AVAILABLE IN THE BOOKSTORE:**

General Chemistry as a Second Language, David R. Klein.

Chemistry Survival Skills by Margaret Brault and Margaret MacDevitt. This book will help the chemistry student be more successful in the course.

Student Study Guide by Raymond Davis. Chapter summaries, study goals, 80 drill and concept questions per chapter with answers.

Student Solutions Manual by Yi-Noo Tang and Wendy Keeney-Kennicutt. Answers and solutions to all even-numbered end-of-chapter exercises.

Schaum's Outline of College Chemistry by Jerome Rosenberg. Theory and problems with complete solutions.

Goggles and gloves for the laboratory.

**General Chemistry I** is intended to provide you with an initial exposure to a broad realm of fundamental concepts in chemistry. It will assist you in attaining a basic understanding of these concepts, and it will help you to develop essential skills in these areas. The lectures, recitation discussions, laboratory sessions, homework assignments, quizzes, and examinations provide an integrated selection of activities which can lead you to success, provided that you are conscientious. In order to receive credit for the course, you must at least meet the minimum requirements described below. Additional effort and achievement will be especially rewarding however.

**Mercer's Academic Integrity Policy:**

Academic integrity is violated when a student:

- A. Uses or obtains unauthorized assistance in any academic work.
- B. Gives fraudulent assistance to another student.
- C. Knowingly represents the work of others as his/her own, or represents previously completed academic work as current.
- D. Fabricates data in support of an academic assignment
- E. Inappropriately or unethically uses technological means to gain academic advantage

Violators will be penalized in accordance with college policy.

**Performance Objectives:**

- 1. You must satisfactorily complete on an individual basis the assigned laboratory experiments. (Missing 3 or more will constitute an F or W for the course)
- 2. You must participate in weekly recitations (missing 3 or more may constitute an F or W for the course)
- 3. You must complete the weekly quizzes and hour tests, as assigned.
- 4. You must achieve a passing grade on a comprehensive final examination.
- 5. You must complete a minimum of 6 hours of work (not counting lab write-ups) on chemistry at home each week.
- 6. You must demonstrate your level of performance (see page 3 for "grading") by mastering a large part of the material covered by lectures, videos, homework, laboratory work and the textbooks as detailed in the specific course objectives that follow.

**Specific Course Objectives:**

You should be able to:

**Unit I: Introduction:**

1. Convert any number to its equivalent in scientific notation, or any number in scientific notation to its equivalent in decimal notation.
2. Perform mathematical operations (addition, subtraction, multiplication, division, square root) using numbers in scientific notation.
3. Solve simple algebraic equations for one unknown.
4. Perform dimensional analyses to verify the proper units in a mathematical operation.
5. Identify the number of significant figures for any number and for the result of any mathematical operation.
6. Distinguish between precision and accuracy.
7. Distinguish between fundamental quantities and derived quantities.
8. Learn the fundamental units of mass, distance, time, temperature, and charge; and the derived quantities of area, volume, and density.
9. Learn the metric prefixes from micro to mega.
10. Learn, use and convert between the metric prefixes: milli, centi, and kilo.
11. Convert between the English and Metric systems for mass, length, volume, etc., given the conversion factors.
12. Determine the density, mass, or volume of any substance given two of the three quantities.
13. Given the density of two immiscible (do not mix) liquids ( or a solid and a liquid), explain which would be on top (or bottom) and why.
14. State the relationship between solid and liquid volume in the metric system.
15. Estimate (in metric units) the approximate mass and approximate size of common household items.
16. Distinguish between chemical and physical properties and changes.
17. Differentiate between extensive and intensive properties, including examples.

18. Define and give examples for the following items:

- |             |                   |                  |
|-------------|-------------------|------------------|
| a. Atom     | f. Homogeneous    | k. Isotope       |
| b. Element  | g. Heterogeneous  | l. Atomic number |
| c. Molecule | h. Symbol         | m. Mass number   |
| d. Compound | i. Ion            | n. Atomic weight |
| e. Mixture  | j. Polyatomic Ion |                  |

19. Write or interpret isotopes in nuclear notation and the notation: C-12

20. Explain the Law of Definite Proportions and the Law of Multiple Proportions.

21. Explain formulas and equations, using examples, and be able to balance simple chemical equations.

22. Distinguish between kinetic and potential energy; and between exothermic and endothermic processes, using examples.

23. Differentiate between heat and temperature.

24. Define the calorie and specific heat capacity and relate them.

25. Demonstrate a knowledge of the temperature scales (Fahrenheit, Celsius, Kelvin) and be able to convert from one scale to the others.

26. Graph experimental data; determine the slope, intercept and equation of a straight line graph.

27. Take a measurement to the accuracy of the instrument with balances, meter sticks, graduate cylinders, and quantitative glassware.

28. Name from the formula (or give the formula from the name) the simple monatomic ions and the following common polyatomic ions; as well as their partially hydrogenated forms:

- |                     |              |                                       |              |
|---------------------|--------------|---------------------------------------|--------------|
| a. $\text{NH}_4^+$  | ammonium     | i. $\text{C}_2\text{H}_3\text{O}_2^-$ | acetate      |
| b. $\text{OH}^-$    | hydroxide    | j. $\text{MnO}_4^-$                   | permanganate |
| c. $\text{NO}_3^-$  | nitrate      | k. $\text{CO}_3^{2-}$                 | carbonate    |
| d. $\text{NO}_2^-$  | nitrite      | l. $\text{HCO}_3^-$                   | bicarbonate  |
| e. $\text{ClO}_4^-$ | perchlorate  | m. $\text{C}_2\text{O}_4^{2-}$        | oxalate      |
| f. $\text{ClO}_3^-$ | chlorate     | n. $\text{SO}_4^{2-}$                 | sulfate      |
| g. $\text{ClO}_2^-$ | chlorite     | o. $\text{SO}_3^{2-}$                 | sulfite      |
| h. $\text{ClO}^-$   | hypochlorite | p. $\text{PO}_4^{3-}$                 | phosphate    |

**Unit II: Stoichiometry:**

1. Determine the atomic number, atomic weight, or formula weight (using the Periodic Table) of any substance, given its symbol or formula.
2. Explain what the atomic weight of an element represents and why it is not an integer.
3. Distinguish between molecular weight and formula weight.
4. Describe the relationship between the mass of a substance and the number of moles that mass represents.
5. Find the mass, atomic or formula weight, or number of moles of any substance, given two of the three quantities.
6. Describe the relationship between the number of moles of a substance and Avogadro's Number.
7. Calculate the mass, moles, and number of atoms or molecules in a substance given any one of them, a Periodic Table, and Avogadro's Number.
8. Find the percent composition by weight of all elements of any pure compound, given its formula and the Periodic Table.
9. Find the mass (or moles) of an element in a compound given the mass (or moles) of the compound, and vice versa. The formula and Periodic Table are also given.
10. Find the empirical formula of a compound, given the percentage composition of all but one of the elements in the compound and the Periodic Table.
11. Find the empirical formula of a compound given the relative number of grams or moles of each element (or a way to get them) and a Periodic Table.
12. Given the grams or moles of a compound, find the grams or moles of each element in it, and vice versa.
13. Derive the molecular formula of a compound, given the molecular weight and empirical formula.
14. Balance simple chemical equations by inserting the proper coefficient before each symbol or formula.
15. Explain the meaning of a balanced chemical equation, especially with regard to the coefficients of the substances involved.

16. Define and determine the limiting reagent in an equation, given the equation and quantities of reactants.
17. Determine the number of moles, the number of particles, and the mass or volume of all substances involved in a reaction, given the equation and quantities of one or more reactants or products.
18. Define solute, solvent, and solution, giving examples.
19. Define, calculate, and use in calculations:  
Weight(mass) percent, and molarity.
20. Define, explain, and calculate with the dilution formula.
21. Given two of the following: molarity, weight percent, density, calculate the 3rd.

### **Unit III: Gases:**

1. Discuss the three states of matter in terms of the motion and closeness of their molecules and what determines the shape of the state.
2. Outline the basic tenets of the Kinetic-Molecular Theory and relate the concepts in this theory to observable properties of a gas such as temperature or pressure.
3. Relate (and sketch) the kinetic energy (and velocity) of gases vs. temperature and explain how they are affected by temperature. (Maxwell-Boltzman Distribution).
4. Give the formula for kinetic energy and use it to relate the speeds of two gases at the same and different temperatures.
5. Define pressure and explain its units and how it is measured.
6. Measure the atmospheric pressure as well as the pressure of a given gas in an enclosed container in units of millimeter mercury height (or torr), given the appropriate Barometer and a demonstration.
7. Define one atmosphere in terms of mm Hg or Torr.
8. Describe qualitatively the effect on a sample of a gas if any one of the three variables (pressure, volume, temperature) is increased or decreased, using the Kinetic-Molecular Theory.

9. State the ideal gas laws (Boyle's Law, Charles' Law, Gay-Lussac's Law, General law) and use these relationships to find the resulting pressure, volume, or temperature of a sample of a gas if the changes it undergoes in two of these variables are adequately specified.
10. Be able to plot and interpret Boyle's and Charles' Laws.
11. State the generalized gas law relationship,  $PV = nRT$ , and use this equation to find any one of the variables if the remaining quantities are specified or to determine any of the laws in the previous objective.
12. Define and explain S.T.P. conditions.
13. Derive an absolute scale of temperature from the behavior of ideal gases at constant pressure.
14. State Dalton's Law of partial pressures, and use it in calculations, including when water vapor is one of the gases.
15. State Avogadro's Principle and describe what is meant by the "molar volume" of an ideal gas.
16. Calculate the value of R in liter-atm/mole  $^{\circ}\text{K}$ , given the molar volume of an ideal gas at S.T.P.
17. Derive the relationships between the molecular weight and the density of a gas and measurable quantities, and use this equation in calculations.
18. State Graham's Law and show how it can be obtained from the fact that absolute temperature is proportional to the average kinetic energy of the molecules, given by  $1/2mv^2$ .
19. Apply Graham's Law to the separation of gases, the prediction of relative rates of diffusion, or effusion, or finding a molecular (or atomic) weight.
20. Define an ideal gas in terms of its lack of molecular volume and its lack of attractions among its molecules.
21. Explain why real gases do not obey the ideal gas laws perfectly, and explain the meaning of the terms in the Van der Waals equation of state of a real gas.
22. Name the temperature and pressure conditions (i.e., high, low) under which a real gas behaves more ideally or non-ideally, and why this is so.

**Unit IV: Crystal Structure:**

1. Define and relate the wave properties (and their symbols); wavelength ( $\lambda$ ), velocity ( $v$ ), frequency ( $\nu$ ), and amplitude ( $A$ ), and illustrate them on a picture of a wave.
2. Be able to use the relationships between wavelength, velocity and frequency of a wave in calculations.
3. Define and illustrate diffraction and interference, and state Bragg's Law for constructive interference for both crystals and a diffraction grating.
4. Describe the electromagnetic spectrum in terms of wavelength, frequency, energy, and types of waves (especially the color of the visible portion.)
5. Describe the external features of a crystal.
6. Explain the use of x-rays in determining crystal structure.
7. Distinguish between an amorphous solid and a crystalline solid.
8. State and explain Bragg's Law as it applies to crystal lattices and use it in calculations.
9. Define and give examples of unit cell, lattice, and space lattice.
10. Describe the geometries of the seven basic shapes of unit cells: cubic, tetragonal, orthorhombic, monoclinic, triclinic, hexagonal, and rhombohedral.
11. Describe the four types of unit cells: simple, body-centered, face-centered, and end-centered, and state which ones exist for cubic unit cells.
12. Explain why there are only 14 crystal systems.
13. Determine what fraction of a sphere will be in a unit cell for spheres at the following lattice points: corner, edge, face-center, body-center.
14. Determine the number of metal atoms contained in each of the types of unit cells.
15. Explain that face centered cubic (cubic closes packing) is the most efficient and most common type of packing; especially for metals such as Cu, Al, Ag, Au, etc.
16. Be able to apply the Pythagorean Theorem to unit cells.
17. Calculate the radius of a metal atom, given the unit cell length (or vice versa), for a simple cubic, face-centered cubic, and body-centered cubic unit cell.
18. Calculate the face-diagonal of a cubic unit cell given the edge length or sphere size, or vice versa.

19. Calculate the % void (unoccupied) space in or density of a simple, face-centered, and body-centered cubic unit cell with spheres of diameter or radius of one angstrom.
20. Describe the location, shape, number, and relative size of tetrahedral and octahedral sites within the face-centered cubic unit cell, and the simple cubic site, and describe how this leads to ionic crystals.
21. Describe the packing, given the number of formula units per unit cell, and give the general formula and an example for the four types of cubic closest packing of ionic salts: rock salt, zinc blende, fluorite, and antiferite.
22. Find the ionic radii (and the density) for a salt in the rock salt structure, given the length of the unit cell, and vice versa.
23. Describe the packing of the cesium chloride structure for ionic salts, given the number of formula units for unit cell, and be able to do calculations relating the unit cell length to the ionic radii.

### **Unit V: Liquids and Changes of State:**

1. Define and give examples for the following terms:

a. Compressibility	j. Freezing
b. Diffusion	k. Melting
c. Surface tension	l. Fusion
d. Evaporation	m. Crystallization
e. Condensation	n. Sublimation
f. Vapor pressure	o. Boiling point
g. Critical temperature	p. Normal boiling point
h. Critical pressure	q. Melting point
i. Boiling	r. Triple point
2. Define the four states of matter.
3. Explain on a molecular level why diffusion in a liquid is slower than in a gas.
4. Explain on a molecular level the cause of surface tension and which states of matter exhibit it.
5. State the differences among the three states of matter for the physical properties of density, compressibility, and ability to flow.
6. Explain on a molecular level why evaporation is a cooling process.
7. Define the heat of vaporization and the heat of fusion, and use them in calculations.

8. Describe the attractive forces in a liquid and relate them to such properties as heat of fusion, heat of vaporization, vapor pressure, boiling point, melting point, and surface tension.
9. Describe vapor pressure in terms of the dynamic equilibrium that exists between the liquid and gaseous states.
10. Define the principle of Le Chatelier and apply it to a liquid-gas, liquid-solid, and solid-gas equilibria in terms of volume or temperature stresses.
11. Describe the effect of size on the polarizability of a series of like-compounds such as the hydrides of Groups VA, VIA, and VIIIA.
12. Describe the effects of the polarizability of a molecule on the physical properties (e.g. boiling point, vapor pressure, heat of fusion, etc.) of that molecule.
13. Describe the course and nature of hydrogen bonds and its effect on a series of boiling points (and other physical properties) of like compounds such as the hydrides of Groups VA, VIA, and VIIIA.
14. Sketch and interpret all portions of heating and cooling curves.
15. Sketch or interpret a "phase diagram" in terms of triple points, critical points, normal boiling and melting points, and the lines that reflect the equilibrium between the states or phases of a substance. All phase changes should also be included, along with the Kinetic-Molecular Theory, in your description.
16. Explain what would happen to a substance as its temperature or pressure were changed, given its phase diagram.

### **Unit VI: Atomic Structure and Periodic Properties:**

1. Define and give examples of the following terms:

a. Electron	f. Coulomb	k. Atomic number
b. Neutron	g. Radioactivity	l. Mass number
c. Proton	h. Alpha particle	m. Isotope
d. Cathode rays	i. Beta particle	n. X-rays
e. Nucleus	j. Gamma ray	
2. Determine how many protons, neutrons, and electrons a given atom (or ion) has, given its symbol, atomic number, and mass number (and vice versa).
3. Describe the relative mass, charge, and location of the three basic particles in the atom ( $\alpha$ ,  $\beta$ ,  $\gamma$ )
4. State what isotope serves as the current standard for the atomic mass scale.

5. Explain why atomic masses of some elements (i.e., Cu, Cl) are so far from whole numbers.
6. Summarize the contributions made to our early understanding of atomic structure by:
  - a. Joseph J. Thompson
  - b. Robert A. Millikan
  - c. Eugene Goldstein
  - d. James Chadwick
  - e. Antoine Becquerel
  - f. Ernest Rutherford
  - g. Henry Mosley
  - h. Niels Bohr
7. Describe the three important kinds of radiation emitted by radioactive substances.
8. Summarize Rutherford's gold-foil experiment and the conclusions that can be made from its outcome.
9. Describe the evidence for the existence of electronic energy levels or orbits in atoms.
10. Compare and contrast line and continuous spectra, and list sources of each.
11. Explain how atomic spectra are obtained, what they look like, and what they mean.
12. Outline the Bohr concept of atomic structure.
13. Explain the meaning of the Rydberg equation and perform a Rydberg calculation to determine the quantum level change and corresponding energy change for a line in the emission spectrum of an atom such as hydrogen from a measured or given wavelength emitted. From this one should be able to construct an energy level diagram for the atom.
14. State the relationships between energy, wavelength, and frequency and use them in calculations.
15. Explain and state the differences among the Balmer, Lyman, and Paschen series in the hydrogen spectrum and relate them to the Rydberg Equation.
16. State Planck's quantum theory relating energy with wavelength (or frequency) of radiation of a given frequency or wavelength.
17. State the Heisenberg Uncertainty Principle and interpret how it describes the limitation on our simultaneous knowledge of the momentum and position of a moving electron.
18. State Louis De Broglie's contribution relating the wavelength and momentum of a particle.
19. State Erwin Schrodinger's contribution to the quantum picture of the atom.

20. Explain the wave function ( $\psi$ ) in the solution to the Schrodinger Wave Equation describes the electron in an atom, and how its square ( $\psi^2$ ) relates to the probability of finding the electron at any point around the nucleus.
21. Explain the relationship between  $\psi^2$ , orbitals, and electron clouds, and how they reflect our current picture of the atom.
22. Give the symbols for, and name the four quantum numbers which describe an electron; state what values they may assume, and their relationship to each other. This should include the relationship between  $n$  and  $K, L, M, \dots$ ; and  $s, p, d, f, \dots$ .
23. Arrange the orbitals described by these quantum numbers in order of increasing energy.
24. Explain what the four quantum numbers tell us in terms of energy, location, and shape of the electron in the atom, and how many electrons can fit in each orbital.
25. Explain how many electrons fit into each orbital and "construct" successively larger atoms or elements by "filling" these orbitals with the appropriate number of electrons.
26. Define and give examples of the terms shell and subshell.
27. State the similarities and differences between the Bohr picture and the modern quantum picture of the atom.
28. State the Pauli Exclusion Principle and Hund's Rule and apply them to the structure of the atom.
29. Draw an energy level diagram for the lowest energy state of an element. This diagram should be labeled as to the  $n$  and  $l$  values of each level, and with arrows up or down for the spin of the electron in each occupied orbital.
30. Sketch the spatial arrangements of  $s, p$ , and  $d$  orbitals.
31. Explain why there is only one  $s$  orbital in each shell, why there are three and only three  $p$  orbitals in each case, etc.
32. Write the electronic configuration in spectroscopic notation (e.g.  $1s^2, 2s^2 2p^6$  etc.) for an atom of any element, or its ion; or from the atomic number of the element.
33. Explain the few exceptions to the predicted order of filling of subshells in terms of the stability of a half-filled subshell.
34. Explain how the format of the Periodic Table results from the energy levels of the orbitals.

35. Define valence electrons, and write the valence electronic configuration for an element from its position in the Periodic Table or locate its position from its valence configuration.
36. Write the symbols from the names of the chemical elements and vice versa, for the first 20 elements.
37. State the Periodic Law and describe the Periodic Table as an arrangement of the elements in the order of their atomic numbers so that elements of similar electronic structure and similar chemical and physical properties are in the same column.
38. Define period, group, family, A and B family, and give examples from the Periodic Table.
39. Identify the following groups and series in the Periodic Table and correlate their identities with their valence electronic configurations:
- Alkali metals ( $ns^1$ )
  - Alkaline earth metals ( $ns^2$ )
  - Halogens ( $ns^2np^5$ )
  - Noble gases ( $ns^2np^6$ )
  - Representative elements (s and p groups)
  - Transition metals (nd)
  - Lantanides (rare earth elements) (4f)
  - Actinides (5f)
  - Inner transition elements (nf)
40. Define the following terms, giving examples, and describe the trends in any row and in any column of the Periodic Table for each one:
- |                      |                               |
|----------------------|-------------------------------|
| a. Atomic radius     | f. First ionization potential |
| b. Ionic radius      | g. Electron affinity          |
| c. Electronegativity | h. Metallic nature            |
| d. Density           | i. Acidic/basic strength      |
| e. Melting point     | j. Oxidizing/reducing power   |
41. Distinguish among the terms metal, nonmetal, and semimetal (or metalloid) and determine which elements in the Periodic Table fall into which category.
42. Define the term isoelectronic and pick out from a series of atoms and ions those which are isoelectronic, and be able to list them according to decreasing radius (or increasing ionization potential).
43. Describe a metallic lattice and metallic conduction in terms of a "sea" of electrons, and relate this to electrical and thermal conductivity.

44. Define the terms malleability and ductility.
45. Describe the similarities as well as the range of chemical and physical properties of the metals (eg. ease of oxidation, reactivity, melting points, etc.)
46. Describe the range of some of the properties of the nonmetals.

### **Unit VII: Chemical Bonding:**

1. Define, describe and distinguish among ionic, covalent, polar covalent, and metallic bonding.
2. Determine the ion an atom will form from its position in the Periodic Table or its valence electronic configuration, and determine the electron configuration of an ion.
3. Name from the formula (or give the formula from the name) the simple monatomic ions and the following common polyatomic ions; as well as their partially hydrogenated forms:
  - a.  $\text{NH}_4^+$  ammonium
  - b.  $\text{OH}^-$  hydroxide
  - c.  $\text{NO}_3^-$  nitrate
  - d.  $\text{NO}_2^-$  nitrite
  - e.  $\text{ClO}_4^-$  perchlorate
  - f.  $\text{ClO}_3^-$  chlorate
  - g.  $\text{ClO}_2^-$  chlorite
  - h.  $\text{ClO}^-$  hypochlorite
  - i.  $\text{C}_2\text{H}_3\text{O}_2^-$  acetate
  - j.  $\text{MnO}_4^-$  permanganate
  - k.  $\text{CO}_3^{2-}$  carbonate
  - l.  $\text{HCO}_3^-$  bicarbonate
  - m.  $\text{C}_2\text{O}_4^{2-}$  oxalate
  - n.  $\text{SO}_4^{2-}$  sulfate
  - o.  $\text{SO}_3^{2-}$  sulfite
  - p.  $\text{PO}_4^{3-}$  phosphate
4. Determine the formula for an ionic compound from the position of its elements in the Periodic Table or the charges on the ions that form the compound. You should also be able to name these compounds.
5. Describe the Born-Haber Cycle including the terms lattice energy and heats of formation of an ionic substance, and explain how it is used to describe this information.
6. Explain the octet rule for ionic and covalent substances, and list which elements usually obey it as well as those that violate it.
7. Define Lewis Structure, and be able to write Lewis Structure for elements, monatomic ions, polyatomic ions, ionic compounds, and covalent compounds.

8. Apply the concept of electronegativity to predict which compounds are predominantly ionic or covalent, or more ionic or covalent.
9. Define and give examples of the following terms:
  - a. Bond length
  - b. Bond energy
  - c. Bond order
  - d. Dipole
  - e. Dipole moment
  - f. Polar
  - g. Hydrogen bonding
  - h. Coordinate covalent bond
10. Define ionic potential and apply it to determine the relative ionic-covalent character of compounds.
11. Apply the concept of electronegativity to predict if a bond will have a dipole or be polar, then determine if the whole molecule will have a dipole moment, based on its structure.
12. Use the relative size or charge of a cation or anion to predict the relative ionic-covalent character of a compound.
13. Relate the relative ionic-covalent character of a compound to its properties: solubility, acid-base character, color, melting point, and cation hydrolysis.
14. Summarize the two most commonly used elementary descriptions of covalent bonding:
  - a. Valence bond (VB) theory
  - b. Molecular orbital (MO) theory
15. Define sigma ( $\sigma$ ) and pi ( $\pi$ ) bonds and sketch their formation from atomic orbitals of appropriate symmetry.
16. Describe and use the devices of promotion and hybridization to explain covalent bonds and the geometries of molecules.
17. Describe the hybridization, orbital geometry, and molecular geometry from the formula or structure of a compound or ion.
18. Explain how the valence bonds are formed in a molecule or ion, and list the atomic orbitals from which they are formed.
19. Explain the nature of multiple bonds and how they are formed in molecules, using examples (or given an example).
20. Define resonance and give valence bond structures for molecules or ions which exhibit it. (e.g.,  $\text{NO}_2^-$ ,  $\text{NO}_3^-$ ,  $\text{SO}_2$ ,  $\text{SO}_3$ ,  $\text{CO}_3^{2-}$ ).
21. Explain the electron pair repulsion theory and use it to determine molecular geometry.
22. Describe the elemental form and structure of the nonmetals, and relate this to some of their chemical and physical properties.

23. Explain why the nonmetal and semimetal elements of the Second Period of the Periodic Table form stable p-p pi bonds while those of the Third Period cannot; and use this to explain the elemental structure of these elements.
24. Define the term "allotropic", and give examples.
25. Describe and name the allotropic forms of oxygen and carbon, describing some of their properties.

### **Unit VIII: Chemical Reaction in Aqueous Solutions:**

#### **A. Solution Concentrations:**

1. Define and give examples of the following terms:

a. Solution	f. Saturated
b. Solvent	g. Unsaturated
c. Solute	h. Supersaturated
d. Concentrated	I. Solubility
e. Dilute	j. Equilibrium
2. Define, give symbols for, calculate the values of, and describe the preparation of solutions in the following concentration units: (given appropriate data)

a. Weight percent (wt%)	d. Molarity (M)
b. Parts per million (ppm)	e. Molality (m)
c. Mole fraction (X)	f. Normality (N)
3. Given any two of the following three items: Molarity, percent by weight, density; calculate the third one.
4. Convert from one unit of concentration to another, given the information required.
5. Prepare a solution of given concentration and accuracy, given the appropriate equipment and chemicals.
6. Calculate and be able to dilute one solution to obtain another, given all but one of the concentrations and volumes, and the proper equipment.

**B. Acids, Bases, and Salts:**

1. Define and give examples of the following terms:
 

a. Acid	k. Precipitation
b. Base	l. Neutralization
c. Salt	m. Dissociation
d. Anion	n. pH
e. Cation	o. Acid anhydride
f. Ionization	p. Basic anhydride
g. Indicators	q. Conjugate acid
h. Electrolyte	r. Conjugate base
i. Nonelectrolyte	s. Oxoacid
j. Hydration	t. Binary acid
  
2. Explain, compare, and give examples for "strong" and "weak" for the following: acids, bases, electrolytes.
  
3. Explain the "dynamic chemical equilibrium" that reflects the ionization of weak electrolytes, and compare this to a strong electrolyte.
  
4. Write the balanced molecular and ionic equation for a neutralization reaction, given the acid and base involved, or the salt produced.
  
5. Define and give examples of monoprotic and polyprotic acids and write the equations for the stepwise ionization of polyprotic acids.
  
6. Define and give examples of acid salts.
  
7. Name the formulas (and vice versa) of the common acids:
 

HCl	Hydrochloric acid	HMnO <sub>4</sub>	Permanganic Acid
HNO <sub>3</sub>	Nitric acid	HOAc	Acetic Acid
HNO <sub>2</sub>	Nitrous acid	H <sub>2</sub> SO <sub>4</sub>	Sulfuric acid
HClO	Hypochlorous acid	H <sub>2</sub> SO <sub>3</sub>	Sulfurous acid
HClO <sub>2</sub>	Clorous acid	H <sub>2</sub> CO <sub>3</sub>	Carbonic acid
HClO <sub>3</sub>	Chloric acid	H <sub>3</sub> PO <sub>4</sub>	Phosphoric acid
HClO <sub>4</sub>	Perchloric acid	H <sub>2</sub> C <sub>2</sub> O <sub>4</sub>	Oxalic acid
  
8. Define an acid solution as one which contains an excess amount of hydrogen ions, H<sup>+</sup> (sometimes called hydronium ions) over hydroxyl ions, OH<sup>-</sup>, and that it will turn a blue litmus to pink.

9. Define a basic solution as one which has an excess amount of hydroxyl ions,  $\text{OH}^-$ , over hydrogen ions,  $\text{H}^+$ , and that it will turn a pink litmus paper blue.
10. Describe the pH scale in terms of the relative acidic or basic strength of a solution.
11. Describe the use of indicators in determining the pH of a solution.
12. Explain the pH of a solution of a salt made from:
  - a. A strong acid and a strong base
  - b. A strong acid and a weak base
  - c. A weak acid and a strong base
13. Write a precipitation equation given the salts involved and their solubilities.
14. Give the Bronsted-Lowry definition of acids and bases and use it to identify the acid, base, conjugate acid, and conjugate base in a reaction.
15. Write a chemical equation to produce the conjugate acid or conjugate base of a molecule.
16. Write the chemical equation for the autoionization of water.
17. Give the Lewis definition of acids and bases.
18. Describe and explain the relative strength of the oxoacids in terms of their structure:  
eg. O's without H's, central atoms and other electron withdrawing groups.
19. Describe the relative strength of the binary acids of a family or period in terms of the size and electronegativity of the atoms.
20. List the formulas for the strong and weak acids and bases in water.

**C. Oxidation - Reduction:**

1. Define and give examples of each of the following:
  - a. Oxidation
  - b. Reduction
  - c. Oxidation state
  - d. Half-reaction
  - e. Oxidation number
  - f. Oxidizing agent
  - g. Reducing agent

2. Assign the following oxidation numbers:
  - a. 0 Pure elements
  - b. +1 Alkali metals
  - c. +2 Alkaline earth metals
  - d. -1 Halogens as halides (with metals)
  - e. +1 Hydrogen with nonmetals
  - f. -1 Hydrogen as hydride (with metals)
  - g. -2 Oxygen as oxide
  - h. -1 Oxygen as peroxide
3. Using the above assignments and charges on polyatomic ions, assign oxidation numbers to all other elements in a compound or ion.
4. Determine the change in oxidation number of an element in a reaction and use this to identify if it is oxidized or reduced and to isolate the oxidation half-reaction and the reduction half-reaction in terms of the covalent molecules or ions involved.
5. Balance oxidation half-reactions and reduction half-reactions with respect to mass and charge by the ion-electron method.
6. Balance oxidation-reduction reactions by finding balanced ionic equations and balanced molecular equations for both acidic and basic systems.
7. Describe the reaction of metals with acids and give examples with chemical equations.
8. Describe the relative ease of oxidation of the metals.
9. Describe the activity series for metals and use it to explain single displacement reactions.
10. Describe the reaction of metals with oxygen and write chemical equations as examples.
11. Predict the oxidation states of a metal based on its valence electronic configuration, and whether it is a Representative or Transition Metal.
12. State the trend in oxidation states for the Representative Metals that have two oxidation states; and use this to predict which of two metals will be the stronger oxidizing or reducing agent.

**D. Quantitative Aspects of Reactions in Solutions:**

- Define and give examples of the following terms:
  - Titration
  - End point
  - Equivalence point
  - Equivalents
  - Equivalent weight
  - Normality
- Relate grams, equivalents, and equivalent weight; and calculate any one of them, given the other two.
- Determine the number of equivalents that equals one mole of an acid, a base, an oxidized species, and a reduced species; from its reaction (or other appropriate information); and use it to connect between each of the following pairs:

moles	equivalents
N $\left(\frac{\text{equivalents}}{\text{liter}}\right)$	M $\left(\frac{\text{moles}}{\text{liter}}\right)$
eq. wt. $\left(\frac{\text{g}}{\text{eq}}\right)$	M.W. or F.W. $\left(\frac{\text{g}}{\text{mole}}\right)$

- Identify the end point of a titration as when the indicator changes color and the equivalence point as when:
 
$$\# \text{ Eq. ACID} = \# \text{ Eq. BASE} \quad \text{or}$$

$$\# \text{ Eq. OXIDIZED} = \# \text{ Eq. REDUCED}$$
- Use the above equations to determine the number of equivalents, equivalent weight, mass, or percentage purity of a solid titrated in an acid-base or oxidation-reduction reaction, given the necessary information.
- Use the above equations to determine the number of equivalents, equivalent weight, mass, percentage purity, volume, or normality of a solution titrated in an acid-base or oxidation-reduction reaction, given the necessary information.
- Properly use burettes and perform a redox or acid-base titration in the laboratory with a reasonable degree of accuracy.
- Explain the use of permanganate as an oxidizing agent, and list its advantages in a redox titration.

**Unit IX: Properties of Solutions:**

1. List the three kinds of mixtures: suspensions, colloids, and solutions; and describe them in terms of particle size, filtration, and settling; and give examples of each.
2. Define and give examples of each of the following:
  - a. Polar
  - b. Nonpolar
  - c. Solvated
  - d. Hydrated
  - e. Tyndall effect
  - f. Emulsifying effect
  - g. Alloys
3. Name the three types of solutions due to physical state differences and to give at least one example for each of the three.
4. Describe how Polar/Non-polar solute and solvent interactions can account for observed solubilities and explain the term "like dissolves like".
5. Define and give examples of miscible and immiscible liquid systems.
6. Define the heat of solution ( $\Delta H_{\text{SOLN}}$ ) and differentiate between an endothermic solution process and an exothermic solution process.
7. Describe in detail, including molecular and energy considerations as well as molecular attractions, the process of forming a solution with:
  - a. Two liquids
  - b. A solid and a liquid
  - c. A gas in a liquid
8. Apply Le Chatelier's Principle to solubility equilibria for temperature and pressure effects on the solubility of solution, given the state of the components and the heat of solution.
9. Define an ideal solution as one in which the heat of solution is zero.
10. State Henry's Law ( $C_g = k_g P_g$ ) and use it in calculations.
11. Define what is meant by colligative properties of solutions and list four.
12. State Raoult's Law ( $P_A = X_A P_A^0$ ), use it in calculations, and explain why it is so on a molecular level.

13. Apply Raoult's Law to a two component system in which one or both components are volatile, plot the vapor pressure of the components against their mole fractions, find the total pressure on the graph, and determine pressures of all components for any given amounts of the two components. Or, given the pressures, find the mole fractions.
14. Explain what is meant by positive and negative deviations from Raoult's law in terms of:
  - a. The above graph.
  - b. The heat of solution.
  - c. The relative attractions of the components before and after solution.
  - d. What happens to the temperature as the solution is prepared.
15. Describe the freezing point depression or boiling point elevation of a solution in terms of the effect a solute has on the phase diagram of the solvent.
16. State the laws governing boiling point elevation and freezing point depression of ideal solutions:
$$\Delta t_b = K_b m \quad \text{and} \quad \Delta t_f = K_f m \quad \text{and to}$$
  - a. Define the terms in these equations
  - b. Calculate any one term from the other two
  - c. Calculate the molecular weight of a solute
  - d. Calculate the freezing point or boiling point of a solution
17. Explain the effects of electrolytes and solutes on colligative properties, including calculations.
18. Define and give examples for:

a. Osmosis	e. Crenation
b. Osmotic pressure	f. Isotonic
c. Semi-permeable membrane	g. Hypotonic
d. Hemolysis	h. Hypertonic
19. State and explain the van't Hoff equation for osmotic pressure:
$$PV = nRT,$$
 and use it in calculations, including the obtaining of molecular weights.
20. Sketch an experimental apparatus whereby osmotic pressure may be measured and discuss three possible modes of action of a semi-permeable membrane on a molecular level.

**Course Laboratory Objectives:**

1. Expand your understanding of the Course Objectives.
2. Learn to manipulate chemicals and glassware by working alone.
3. Learn to collect and analyze data from an experiment by working alone.
4. Learn how to use laboratory balances.
5. Learn how to do quantitative analysis such as titrations, pipetting and preparation of solutions by working alone.
6. learn how to collect and treat data on the computer.
7. Utilize critical thinking and quantitative reasoning skills in observing, organizing and analyzing data, synthesizing information, interpreting results, and communicating the results of the analyses and laboratory investigations orally and in writing.
8. Perform chemical experimentation in a safe and scientific manner, using proper scientific and laboratory safety procedures.