CHEM 121: General Chemistry I

Spring 2018 – Section 501 – CRN 36461

Instructor: Dr. Jerry Godbout

Office: VAAS 134
Email: jgodbout@unm.edu
Phone: 505-925-8611

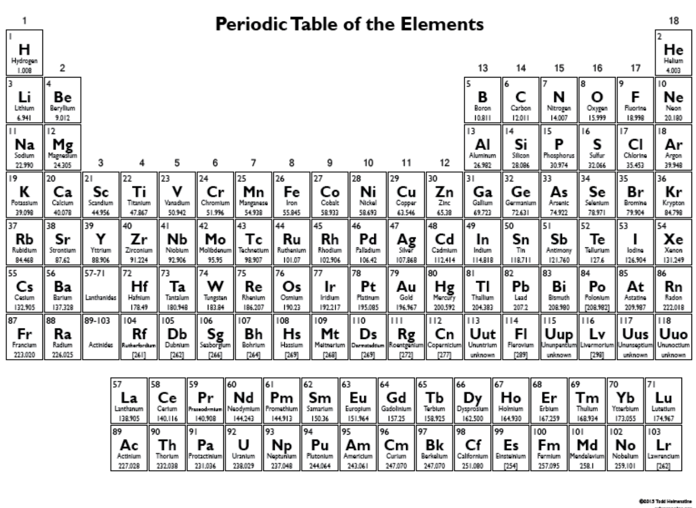
Office Hours: Monday 10:30 am – 12:00 pm,
Tuesday 2:00 pm – 4:00 pm (STEM Center),
Thursday 9:30 am – 11:30, and anytime by appointment

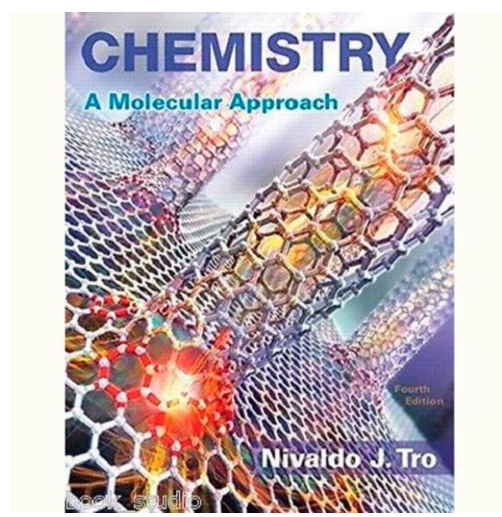
Meeting Times: Lecture: Monday & Wednesday 9:00 – 10:15 am, VAAS 127
Laboratory: Monday 10:30 am – 1:15 pm, VAAS 128

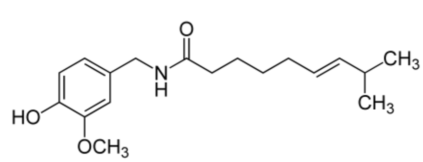
Course Description: The Study of stuff, and what it does (1st of a 2-course sequence)

Course Description: Introduction to the chemical and physical behavior of matter. Meets New Mexico Lower Division General Education Common Core Curriculum Area III: Science (NMCCN 1214). Prerequisite: MATH 121 or MATH 123 or MATH 150 or MATH 153 or MATH 162 or MATH 163 or MATH 180 or MATH 181 or MATH 264 or ACT Math =>25 or SAT Math Section =>590. Pre- or corequisite: 123L. {Summer, Fall, Spring}

Guess which one is the instructor's, and guess which one is has gone through various committees and perhaps a lawyer or two?







What You'll Need
(Required Resources)

- Chemistry: A Molecular Approach
- Mastering Chemistry Access Code
- Calculator (non-graphing) with log/antilog and exponential functions
- Internet Access: *Blackboard Learn* and *UNM email address* ***must be checked regularly (daily)***

What If You Need Help?
(UNM-Valencia Resources)

- **Instructor**: Office hours, STEM Center Hours, email
- **STEM Center**: Tutors*, molecular modelling kits, Laptops, textbooks

* Reminder: when using tutors, it is the **students'** responsibility to make sure they understand well enough to complete the problems on **their own**.

What Will Each Class Be Like?

- **Quiz**: covering material recently covered and any assigned preparation (reading, video, etc)
- **Course Business**
- **Group Activity:** collaborative exercise to help you master that day's topic
- **Reflection:** an opportunity to put the day's lesson into larger perspective, and formulate/ask questions

What Will My Routine Be Like?

- **Before Class**: complete any prepatory assignment (reading, video, etc)
- **During Class:** work with your group to master concepts. The more you put in, the more you'll get out
- **After Class:** work on homework assignment relevant to that day's topic (review notes, **WORK ON PROBLEMS**, think of questions for office hour visits, etc.
- **Repeat 28 times!:**

What You'll Find Useful
(Recommended Resources)

- 3-ring binder for lecture notes, handouts, group activities
- Periodic table (on paper)
- Calculator (non-graphing) with log/antilog and exponential functions
- Mastering Chemistry notebook: keep track of problem solving, identify patterns, record areas of difficulty

How Is Your Grade Determined?
(Exams, Quizzes, Homework, and the Like)

	How Many	Weight
Class Points	1	10 %
Quizzes	15*	10 %
Homework	10*	15 %
Exams	5**	50 %
Final Exam	1	15 %
Total		100 %

* Approximate values

** Each equally weighted, 10 % each

What Do I Need For an A?
(What's the grading scale?)

Earn This %	Get This Grade
98	A+
92	A
90	A-
88	B+
83	B
80	B-
78	C+
73	C
69	C-
67	D+
62	D
60	D-
55	F+
0	F

When We Learn This Stuff?

(Schedule is approximate and subject to change by the instructor)

Meeting	Date	Topics/Events
0	Mon 15 Jan	MLK Day – No meeting
1	Wed 17 Jan	Nuclear Atom GA (2.5 – 2.6)
2	Mon 22 Jan	Dimensional Analysis GA (1.6 – 1.8)
3	Wed 24 Jan	Dalton's Atomic Theory (2.1 – 2.5)
4	Mon 29 Jan	Periodic Table, Average Atomic Mass GA (2.7 – 2.8)
5	Wed 31 Jan	Molar Mass (Counting by Weighing GA) (2.8)
6	Mon 5 Feb	Chemical Bonding, Formulas and Naming (3.1 – 3.7)Exam 1
7	**Wed 7 Feb**	**(Chapters 1 & 2)**
8	Mon 12 Feb	Molar Mass, Balanced Chemical Equations (3.8 – 3.12)
9	Wed 14 Feb	Stoichiometry, L.R., % Yield, Solution Stoichiometry (4.1 – 4.3)
10	Mon 19 Feb	Aqueous Solutions, Molarity (4.4 – 4.6)
11	Wed 21 Feb	Aqueous Reactions, Net Ionic Equations (4.7 – 4.9)
12	**Mon 26 Feb**	**Exam 2 (Chapters 3 – 4)**
13	Wed 28 Feb	Ideal Gas Equation (5.1 – 5.4)
14	Mon 5 Mar	Gas Mixtures, Gas Stoichiometry (5.6 – 5.7)
15	Wed 7 Mar	Kinetic Molecular Theory, Real Gases, Thermodynamics Intro (5.8, 5.10, 6.1 – 6.3)
	Mon 12 Mar	Spring Break (no meeting) Online Thermodynamics Activity
	Wed 14 Mar	Spring Break (no meeting)
16	Mon 19 Mar	Thermochemistry and Calorimetry (6.3 – 6.7)
17	Wed 21 Mar	Hess' Law and Reaction Enthalpies (6.8 – 6.9)
18	**Mon 26 Mar**	**Exam 3 (Chapters 5 – 6)**
19	Wed 28 Mar	Atomic Orbitals and Shapes (7.5 – 7.6)
20	Mon 2 Apr	Electronic Configurations and Periodic Table (8.1 – 8.5)
21	Wed 4 Apr	Periodic Trends (8.6 – 8.9)
22	Mon 9 Apr	Chapter 7 – 8 Wrap-up Activity
23	**Wed 11 Apr**	**Exam 4 (Chapters 7 – 8)**
24	Mon 16 Apr	Lewis Dot Structures (9.1 – 9.5)
25	Wed 18 Apr	Bond Polarity, Dipoles, Bond Characteristics (9.6 – 9.11)
26	Mon 23 Apr	VSEPR Theory (10.1 – 10.5)
27	Wed 25 Apr	Hybridization (10.6 – 10.7)
28	Mon 30 Apr	Molecular Orbital Theory (10.8)
29	**Wed 2 May**	**Exam 5 (Chapters 9 – 10)**
	Mon 7 May	**Final Exam (9:00 – 11:00 a.m.)**

Unit Level Learning Objectives: Exam 1 (Chapters 1-2)

At the end of most learning objectives, there is a reference to a sample problem. These references are the same for both the 3rd and 4th editions of the textbook. The following symbols are used for these references:

CC = Conceptual connection problem within the chapter (answers at the end of the chapter)
Ex = Example within the chapter
EoC = End of chapter problems (answers in Appendix III)

By the end of the chapter, students will be able to…

Ch 1: Matter, Measurement, and Problem Solving

1. Define matter and classify a given substance by physical state.
2. Classify changes in matter as physical or chemical. (Ex 1.1 p 10)
3. Use the appropriate SI units and metric prefixes to express numbers in scientific notation. (Ex 1.4 p 21, Ex 1.5 p 23)
4. Use the concept of density in quantitative and qualitative problems involving masses and volumes. (Ex 1.8 p 29, Ex 1.10 p 31)
5. Report the result of any measurement to the appropriate number of significant figures. (Ex 1.6 p 24)
6. Express the result of any set of simple mathematical operations on measurements to the appropriate number of significant figures. (Ex 1.4 p 21, Ex 1.5 p 23, Ex 1.6 p 24)
7. Analyze a set of measurements for precision and or accuracy. (Ex 1.7-1.8 p 29)
8. Convert between units and prefixed units using dimensional analysis and develop a systematic approach to solving problems involving unit conversion and equations, including the conversion between the three commonly used temperature scales. (Ex 9 p 30, Ex 1.10 p 31)

Ch 2: Atoms and Elements

1. Use the laws of conservation of mass, definite proportions, and multiple proportions to justify Dalton's atomic theory. (Ex 2.1 p 49, Ex 2.2 p 50, CC 2.2 p 50)
2. Justify the nuclear model of the atom with reference to Rutherford, Thompson's, Millikan's experiments, and the scientific method. (CC 2.3 p 53).
3. Identify a set of isotopes from information on the composition of the nucleus. Use atomic notation to write the symbol of any isotope. (Ex 2.3 p 59, CC 2.4 p 59)
4. Identify an element or ion based on the composition of the nucleus and number of electrons. (CC 2.5 p 61, Ex 2.4 p 65)
5. Use the periodic table to classify an element as being a metal (forms cations), nonmetal (forms anions).
6. Identify main group elements and transition elements. Also identify the following groups: alkali metals, alkaline earth metals and halogens and recall the ions commonly formed by elements in these groups.
7. Define the mole and calculate and use average atomic masses to convert between mass, moles and numbers of atoms. (Ex 2.6 p 71, Ex 2.7 p 72, Ex 2.8 p 73, Ex 2.9 p 74)

Unit Level Learning Outcomes: Exam 2 (Chapters 3-4)

At the end of most learning objectives, there is a reference to a sample problem. These references are the same for both the 3rd and 4th editions of the textbook. The following symbols are used for these references:

CC = Conceptual connection problem within the chapter (answers at the end of the chapter)
Ex = Example within the chapter
EoC = End of chapter problems (answers in Appendix III)

By the end of the chapter, students will be able to…

Ch 3: Molecules, Compounds, Chemical Equations

1. Describe the two different forms of bonding that connect atoms - IONIC or COVALENT. Use the periodic table to determine whether a species is molecular or ionic based on chemical formula. (EoC 29 p 130)
2. Determine formulas of ionic compounds, including the use of polyatomic ions, and molecules from their systematic names. (EoC 33 & 35 p 131, Ex 3.2 p 95)
3. Name molecular and ionic compounds using their systematic names. (EoC 37, 41, 47, 49 p131)
4. Determine and use molar mass to convert between mass, moles, and numbers of molecules and atoms in molecules. (Ex 3.13 p 108)
5. Write and balance chemical equations to describe reactions. (Ex 3.22, 3.23, 3.24 p 120-122)

Ch 4: Chemical Quantities and Aqueous Reactions

1. Define molarity and perform calculations involving the composition of solutions, including dilution calculations. (Ex 4.1 p 143, Ex 4.2 p 144, Ex 4.5 p 153, Ex 4.7 p 156)
2. Define and give examples of strong electrolytes, weak electrolytes, and non-electrolytes. Draw molecular level pictures of each type of electrolyte to illustrate the relative degree of ionization in each.
3. Determine the products of a given precipitation reaction by considering the species present in solution and
using a solubility table. (Ex 4.10 & 4.11 p 165)
4. Represent precipitation, acid-base, and gas evolution reactions in solution by molecular, complete ionic, and net ionic equations. (Ex 4.12 p 168, Ex 4.13 p 171)
5. Perform stoichiometric calculations involving precipitation reactions or acid-base neutralization reactions,
including those involving limiting reagent. (Ex 4.14 p 173)
6. Define oxidation and reduction in terms of electron loss and gain. (Ex 4.17 p 179)
7. Assign oxidation states to simple ionic compounds and use oxidation state changes to identify redox reactions, oxidizing and reducing agents. (Ex 4.16 p 178, CC 4.8 p 179)
8. Write balanced equations for combustion reactions, precipitation, and acid-base reactions. (Ex 4.18 p 180,
Ex 4.19 p 182)

Unit Level Learning Outcomes: Exam 3 (Chapters 5-6)

At the end of most learning objectives, there is a reference to a sample problem. These references are the same for both the 3rd and 4th editions of the textbook. The following symbols are used for these references:

CC = Conceptual connection problem within the chapter (answers at the end of the chapter)
Ex = Example within the chapter
EoC = End of chapter problems (answers in Appendix III)

By the end of the chapter, students will be able to…

Ch 5: Gases

1. Recall and use the gas laws (Boyle, Charles and Avogadro) to calculate properties of an ideal gas under changing conditions. (Ex 5.2 p 202, Ex 5.3 p 205, CC 5.1 p 205)
2. Recall and use the ideal gas law, $PV = nRT$ to calculate P, V, n or T given three of the four parameters. (Ex 5.5 p 208, Ex 5.6 p 209)
3. Recall and use the molar volume for an ideal gas 22.42 L at STP (recall that STP is 0 °C (273K) and 1atm). (CC 5.2 p 210, CC 5.3 p 211, Ex 5.7 p 213)
4. Recall and apply Dalton's Law of Partial Pressures to calculate properties relating to mixtures of gases. Use and calculate mole fractions. (CC 5.4 p 216, Ex 5.9 p 216, Ex 5.10 p 218)
5. Apply the ideal gas law to find number of moles from P, V and T conditions, and use this information in stoichiometric calculations. (Ex 5.12 p 221, Ex 5.13 p 223)
6. Recall the three assumptions of Kinetic Molecular Theory and identify situations in which these assumptions fail.

Ch 6: Thermochemistry

1. Define potential energy, kinetic energy and work.
2. State the first law of thermodynamics.
3. Distinguish between heat and temperature. (CC 6.2 p 257)
4. Identify chemical bonds as the source of chemical potential energy.
5. Define energy flow INTO a system as a positive quantity, and energy flow OUT of a system as a negative quantity for the system. Apply the terms 'endothermic' and 'exothermic' to describe the flow of heat between a reaction and its surroundings. Relate these terms to the relative chemical potential energy of reactant and products. (Table 6.3 p 256)
6. Define and use specific and molar heat capacities to calculate temperature changes when heat is applied or removed. (CC 6.3 p 260, Ex 6.3 p 261)
7. Apply stoichimetry to determine enthalpy changes associated with reactions of particular masses of reactants or to form particular amounts of products. (Ex 6.7 p 270, CC 6.5 p 267)
8. Use specific or molar heat capacities to calculate the enthalpy of a reaction in a calorimeter (constant pressure or constant volume). (Ex 6.5 p 266, Ex 6.8 p 271)
9. Use the properties of enthalpy to calculate ΔH for a chemical reaction using Hess's Law. (Ex 6.9 p 274)
10. Look up standard enthalpies of formation for any substance and apply these to calculate ΔH° for a reaction. (Ex 6.10 p 276)

Unit Level Learning Outcomes: Exam 4 (Ch 7-8)

By the end of the chapter, students will be able to…

Ch 7: Electronic Structure of Atoms

1. Use the emission spectrum of hydrogen in the visible region to explain how this line spectrum supports a quantized model of energy levels in hydrogen. (Ex 7.7 p 322)
2. Describe the Bohr model of the hydrogen atom in terms of quantized circular orbits.
3. Use quantum numbers n, l, and ml to describe orbitals. Recall and use the relationships between n, l and ml to determine if any orbital is an allowed one, what type of orbital it is (s, p, d or f orbital), and how many orbitals there are in each l level. (CC 7.4-7.5 p 318, Ex 7.5-7.6 p 320)
4. Sketch the shapes of orbitals designated by s, p, and d. (Figure 7.28 p 327)

Ch 8: Periodic Properties

1. Write electron configurations and orbital diagrams for ground state atoms by applying the Pauli exclusion principle, Hund's rule, the Aufbau principle, and the position of the atom in the Periodic Table. Identify atoms based on electron configurations and orbital diagrams. (Fig 8.5 p 343, Ex 8.1-8.2 p 346)
2. Identify the principle quantum number and the number of valence electrons for an atom or ion and use this information to predict the relative reactivity, size, magnetism, and ionization energy of the atom or ion. (Ex 8.3 p 347, Ex 8.4 p 350)
3. Understand the concept of effective nuclear charge and how it affects atomic size. (Ex 8.5 p 356, Fig 8.12 p 359)

Unit Level Learning Outcomes: Exam 5 (Ch 9, 10)
By the end of the chapter, students will be able to…
Ch 9: Lewis Model of Bonding

1. Describe covalent and ionic bonding with respect to orbitals. (CC 9.1 p 386, Ex 9.3 p 400)
2. Use Lewis structures to represent the valence electrons of molecules and determine bond order and placement of non-bonding electrons. (Ex 9.1 p 388, Ex 9.4-9.5 p 401, Ex 9.6 p 402)
3. Use formal charge considerations to determine the lowest energy resonance structure for a molecule. (Ex 9.7 p 404, Ex 9.8 p 406)
4. Use trends in electronegativity to determine bond polarity. Predict the relative polarity of covalent bonds. (CC 9.4 p 398)
5. Predict relative bond energies and bond lengths in related molecules. (CC 9.8 p 414, Ex 9.11 p 414)

Ch 10: VSEPR and Molecular Orbital Theory

1. Predict the shape of any given molecule by writing the Lewis structure and applying VSEPR to assign the positions of the bonding and non-bonding electrons pairs. (CC 10.1 p 429, CC 10.2 p 431, Ex 10.1 p 432)
2. Compare bond angles in the series methane, ammonia and water to demonstrate that lone pairs repel more than bonded pairs of electrons. (Ex 10.2-10.3 p 438)
3. Draw dipole moments for bonds in molecules, and use these to predict whether a molecule will have a net dipole moment. (Ex 10.5 p 443)
4. Explain what hybridization is and why we invoke it in Valence Bond theory to describe bonding in covalent compounds.
5. Determine the appropriate hybridization of any atom in a molecule using the Lewis structure and the number of electron groups in it (2 to 6 groups). (CC 10.7 p 450)
6. Show how orbitals overlap to form new orbitals with sigma or pi symmetry. Explain why sigma overlap is greater than pi overlap and describe the implications for bond strength. (CC 10.8 p 454)
7. Analyze a given organic 'skeleton' structure to determine geometry of any given atom and the number of sigma bonds and pi-bonds in the structure. (Ex 10.6-10.7 p 459, Ex 10.8 p 460)
8. Draw molecular orbital diagrams for homonuclear diatomics from hydrogen to fluorine and their anion and cation forms. Use MO diagrams to predict bond order, relative bond lengths and strengths, and paramagnetism. (Ex 10.9 p 464, Fig 10.5 p 468)