Appendix 1: Learning objectives for introductory chemistry
Institution 1
Chapters correlate with Tro's *Introductory Chemistry*. The number before the decimal indicates the order of teaching the objectives. For this particular course, chapter 1 is covered in the laboratory, not in lecture, so the learning objectives were added after-the-fact because certain MLC items required them. The chapters are taught in a different order, as was intended to pilot an atoms-first curriculum. Chapter 3 material was covered first, followed by Chapters 2, 4, 6, 9, 5, 10, 12, 11, 13, 7, 16, 8, 15, 14, and 17. As such, the learning objectives reflect this order of instruction for ease of use by the students. However, they are presented here in the chapter order for ease of instructor's adoption.

Chapter 1: The Chemical World
0.0 Recognize that all matter is made of chemicals.
0.1 Identify a process in the scientific method as an observation, a hypothesis, or an experiment.

Chapter 2: Measurement and Problem Solving
2.1 Determine the number of significant figures in a number.
2.2 Apply significant figures to multiplication and division.
2.3 Apply significant figures to addition and subtraction.
2.4 Convert between scientific and standard notation.
2.5 Memorize the meaning of metric prefixes from pico- to tera-.
2.6 Convert between units with the same base unit (e.g., centimeters to kilometers).
2.7 Convert between metric and US units (e.g., kilograms to pounds).
2.8 Combine multiple conversions into one dimensional analysis.
2.9 Modify dimensional analysis to account for units raised to a power (e.g., square centimeters to square inches).
2.10 Convert between Celsius and Kelvin temperature scales.
2.11 Apply density as a conversion factor between mass and volume.

Chapter 3: Matter and Energy
1.1 Classify matter as a pure substance or a mixture.
1.2 Subdivide pure substances as elements or compounds.
1.3 Subdivide mixtures as homogeneous or heterogeneous.
1.4 Group properties as physical or chemical.
1.5 Identify a change as physical or chemical.
1.6 Apply the law of conservation of mass.
1.7 Identify a reaction as exothermic or endothermic based on whether heat is absorbed or released.
1.8 Contrast the physical properties of solids, liquids, and gases.

Chapter 4: Atoms and Elements
3.1 Distinguish the properties of protons, neutrons, and electrons from each other.
3.2 Determine the numbers of all subatomic particles of an isotope or ion.
3.3 Calculate the mass number for a given isotope.
3.4 Discriminate between ions, isotopes, and elements.
3.5 Identify the group (by number and/or name) and period of an element.
3.6 Calculate the atomic mass of an element given relative abundance and mass number of all isotopes.
3.7 Correlate an element's symbol with its name.

Chapter 5: Molecules and Compounds
6.1 Justify the preferred charge of an ion.
6.2 Categorize a compound as ionic, molecular, or acid.
6.3 Generate a name for a compound given its formula.

6.4 Infer the charge of a transition or poor metal from a chemical formula.
6.5 Convert a common name into a systematic name or vice versa for appropriate compounds.
6.6 Produce the formula of a compound given its name.
6.7 Disprove a misnamed compound and modify the compound or name such that both are correct.
6.8 Dissect a chemical formula to determine the identity of elements present and how many there are of each.
6.9 Memorize common polyatomic ions.
6.10 Identify the charge of an ion.
6.11 Identify which elements are diatomic in nature.

Chapter 6: Chemical Composition
4.1 Calculate the molar mass of a compound.
4.2 Distinguish between empirical formula and molecular formula.
4.3 Compare two masses of different substances to determine which has more atoms or molecules.
4.4 Covert a mass of a substance into moles or atoms or vice-versa.
4.5 Determine the mass percent of an element in a compound.
4.6 Construct an empirical compound for a formula given masses of elements or mass percentages of elements.
4.7 Develop a molecular formula for a compound given its molar mass and an empirical formula.

Chapter 7: Chemical Reactions
11.1 Identify components of a reaction as reactants or products.
11.2 Classify a reaction as synthesis (combination), decomposition, double displacement (metathesis), single replacement (substitution), or combustion.
11.3 Predict the products of a metal carbonate undergoing decomposition.
11.4 Predict the products of a double displacement (metathesis) reaction.
11.5 Predict the products of a single replacement (substitution) reaction.
11.6 Determine the coefficients of a chemical reaction so that it is balanced.
11.7 Determine if a given ionic compound will be soluble in water.
11.8 Rearrange a molecular equation into a complete ionic and/or net ionic equation.
11.9 Integrate the concepts of balancing reactions and predicting products of a reaction.

Chapter 8: Quantities in Chemical Reactions
12.1 Compose a dimensional analysis which includes reaction coefficients.
12.2 Calculate the theoretical yield of a product in a reaction given mass or moles of starting materials.
12.3 Determine the limiting reactant in a reaction.
12.4 Calculate the remaining amount of an excess reactant.
12.5 Classify a reaction as exothermic or endothermic based on its reaction enthalpy.
12.6 Calculate the percent yield of a reaction given an actual yield and theoretical yield or vice-versa.
12.7 Integrate reaction enthalpy into dimensional analysis and stoichiometry for use in calculations.
12.8 Integrate concentration concepts into dimensional analysis and stoichiometry for use in calculations.
12.9 Integrate gas laws into dimensional analysis and stoichiometry for use in calculations.

Chapter 9: Electrons in Atoms and the Periodic Table
5.1 Investigate why different elements produce different colors when burned.
5.2 Compare and contrast the quantum mechanical model of electronic structure with the Bohr model.
5.3 Produce an electron configuration for a given atom or ion in standard form.
5.4 Produce an electron configuration for a given atom or ion in noble gas form.
5.5 Arrange atomic orbitals by their relative energy.
5.6 Arrange a group of elements by atomic size, electronegativity, metallic character, or ionization energy.

5.7 Classify electrons as valence electrons or core electrons.

Chapter 10: Chemical Bonding
7.1 For a given compound or polyatomic ion, construct a Lewis structure which obeys the octet rule.
7.2 For an applicable given compound or polyatomic ion's Lewis structure, propose a different resonance form.
7.3 For a given Lewis structure, categorize electron domains as atoms or lone pairs.
7.4 Classify a given Lewis structure according to the number of electron domains it possesses.
7.5 Justify how different electron domains contribute to a Lewis structure's molecular geometry.
7.6 Categorize compounds and polyatomic ions according to their molecular geometries.
7.7 Rank elements in a compound according to their electronegativities.
7.8 Distinguish between ionic bonds, polar covalent bonds, and nonpolar covalent bonds based on electronegativity.
7.9 Evaluate if a molecule is polar or nonpolar based on electronegativity differences and symmetry.
7.10 Predict the solubility or miscibility of compounds in other compounds based on polarity.

Chapter 11: Gases
9.1 Describe the assumptions contained within the kinetic molecular theory.
9.2 Convert the units of pressure into each other, including atmospheres, Pascals, bars, torr, and mmHg.
9.3 Correlate a change in temperature with the pressure of a given gas sample.
9.4 Correlate a change in volume with the pressure of a given gas sample.
9.5 Correlate a change in quantity (moles) with the pressure of a given gas sample.
9.6 Predict a gas's temperature, volume, pressure, or quantity given any three of these.
9.7 Predict how a change in temperature, volume, and/or pressure of a gas sample will alter the temperature, volume, or pressure of a gas.
9.8 Calculate the partial pressure of a gas or the total pressure of a mixture of gases given the composition of a mixture.

Chapter 12: Liquids, Solids, and Intermolecular Forces
8.1 Differentiate between properties of solids, liquids, and gases.
8.2 Diagram how solids, liquids, and gases appear at the molecular level.
8.3 Compare the relative strength of intermolecular forces, including dipole-dipole interactions, ion-dipole interactions, London dispersion forces, and hydrogen bonding.
8.4 Predict which intermolecular force(s) a given molecule might undergo.
8.5 Group compounds by their strongest intermolecular force.
8.6 Arrange similarly-sized molecules by boiling point or melting point.
8.7 Calculate the energy required for a substance to change state.
8.8 Calculate the energy required for a substance to increase or decrease in temperature.
8.9 Synthesize energy calculations that include state changes and temperature changes.
8.10 Categorize solids by type, including metallic solids, ionic solids, and covalent networks.

Chapter 13: Solutions
10.1 Illustrate how ionic and polar covalent substances dissolve in water.
10.2 Arrange a group of compounds by ionic strength in solution (i.e., van't Hoff factor).
10.3 Propose a method to get a substance to dissolve more quickly or more slowly into solution.
10.4 Generalize why certain substances are more soluble in a given solvent than others.
10.5 Calculate the concentration of a solute in % w/w, molar, and molal.
10.6 Dissect how solutes can alter a substance's boiling or melting point.
10.7 Predict the colligative properties of a solution (e.g., boiling point elevation or freezing point depression).
10.8 Classify a substance in a solution as a solute or solvent.

10.9 Classify a mixture as a suspension, solution, or colloid.
10.10 Calculate how a substance’s concentration will change when a solution’s volume is increased or decreased.
10.11 Classify a substance as a strong electrolyte, a weak electrolyte, or nonelectrolyte in water.

Chapter 14: Acids and Bases
14.1 Classify a substance as an acid or base according to the Arrhenius theory.
14.2 Classify a substance as an acid or base according to the Brønsted-Lowry theory.
14.3 Identify the conjugate of a given acid or base.
14.4 Arrange acids by strength given Ka or pKa.
14.5 Arrange bases by strength given Kb or pKb.
14.6 Calculate the pH of a given solution.
14.7 Calculate the pOH of a given solution.
14.8 Calculate the concentration of hydronium in a given solution.
14.9 Calculate the concentration of hydroxide in a given solution.
14.10 Illustrate how buffers achieve an equilibrium that is resistant to pH change.
14.11 Predict how an acid will react with a given base, metal, metal oxide, metal carbonate, or metal bicarbonate.
14.12 Integrate acid-base reactions into dimensional analysis and stoichiometry.

Chapter 15: Chemical Equilibrium
13.1 Illustrate how a chemical reaction takes place at the molecular level.
13.2 Examine how concentration affects the rate of a reaction.
13.3 Examine how temperature affects the rate of a reaction.
13.4 Describe how catalysts affect the rate of a reaction.
13.5 Assemble the components of a reaction into an equilibrium expression.
13.6 Compare equilibrium constants to determine how a reaction will occur.
13.7 Calculate the concentration of a reaction component given an equilibrium constant and other equilibrium concentrations.
13.8 Calculate the equilibrium constant of a reaction based on equilibrium concentrations of aqueous or gaseous reaction components.
13.9 Predict how an equilibrium will shift when the concentration of a component is changed.
13.10 Predict how an equilibrium will shift when the pressure is changed.
13.11 Predict how an equilibrium will shift when the volume is changed.
13.12 Predict how an equilibrium will shift when the temperature is changed.

Chapter 16: Oxidation and Reduction
11.10 Determine the oxidation numbers for each element in a compound or ion.
11.12 Identify which element(s) in a redox reaction is being reduced and which is being oxidized.
11.13 Classify a reaction as redox or non-redox.
11.14 Classify a substance in a reaction as an oxidizing agent or a reducing agent.

Chapter 17: Radioactivity and Nuclear Chemistry
15.1 Classify a nuclear reaction according to the type of nuclear decay.
15.2 Balance a nuclear equation.
15.3 Complete a fusion or fission reaction.
15.4 Calculate the half-life of a given nucleus.
15.5 Calculate the amount of radioactive material remaining after a given time with a given half-life.

Institution 2
Learning Objective (LO): The learning objectives are associated with chapters listed in Tro's chemistry textbook: *Introductory Chemistry*

Chapter 1: The Chemical world
LO 1.1: Understand the definition of chemistry.
LO 1.2: Recognize that chemicals make up everything we come into contact with in our world.
LO 1.3: Identify and understand the key characteristics of the scientific method.

Chapter 2: Measurement and Problem Solving
LO 2.1: Express very large and very small numbers using scientific notation.
LO 2.2: Report measured quantities to the right number of digits.
LO 2.3: Counting significant figures.
LO 2.4: Knowing exact numbers.
LO 2.5: Carry out significant figures in calculations.
LO 2.6: Determine the correct number of significant figures in the results of multiplication and division calculations.
LO 2.7: Determine the correct number of significant figures in the results of addition and subtraction calculations.
LO 2.8: Determine the correct number of significant figures in the results of calculations involving both addition/subtraction and multiplication/division.
LO 2.9: Recognize and work with the SI base units of measurement, prefix multipliers, and derived units.
LO 2.10: Carry out unit conversion: convert between units: one step; multistep; and both the numerator and denominator; units raised to a power.
LO 2.11: Definition of density and calculate the density of a substance.
LO 2.12: Use density as a conversion factor.

Chapter 3: Matter and Energy
LO 3.1: Define matter, atoms, and molecules.
LO 3.2: Classify matter as solid, liquid, or gas and understand the key characteristics of the solid, liquid, or gas.
LO 3.3: Classify matter as pure substance or mixture.
LO 3.4: Distinguish between physical and chemical properties.
LO 3.5: Distinguish between physical and chemical changes.
LO 3.6: Apply the law of conservation of mass.
LO 3.7: Recognize the different forms of energy.
LO 3.8: Apply the law of conservation of energy.
LO 3.9: Identify and convert among energy units.
LO 3.10: Distinguish between exothermic and endothermic processes/ reactions in physical and chemical changes.
LO 3.11: Convert between Fahrenheit, Celsius, and Kelvin temperature scales.
LO 3.12: Know the concept of absolute zero.
LO 3.13: Relate energy, temperature change, and heat capacity.

Chapter 4: Atoms and Elements
LO 4.1: Recognize that all matter is composed of atoms.
LO 4.2: Explain how the experiments of Thomson and Rutherford led to the development of the nuclear theory of the atom.
LO 4.3: Describe the properties and charges of electrons, neutrons, and protons.
LO 4.4: Know the elements is Defined by Their Number of Protons.
LO 4.5: Use the periodic table to classify elements by group.

LO 4.6: Determine ion charge from numbers of protons and electrons.
LO 4.7: Determine the number of protons and electrons in an ion.
LO 4.8: Determine atomic numbers (Z), mass numbers (A), and isotope symbols for an isotope.
LO 4.9: Determine the number of protons and neutrons from isotope symbols.
LO 4.10: Calculate the Average Mass of an Element's Atoms.

Chapter 5: Molecules and Compounds
LO 5.1: Restate and apply the law of constant composition.
LO 5.2: Determine the total number of each type of atom in a chemical formula.
LO 5.3: Classify elements as atomic or molecular. Classify compounds as ionic or molecular.
LO 5.4: Write formulas for ionic compounds.
LO 5.5: Naming ionic compounds.
LO 5.6: Name molecular compounds.
LO 5.7: Naming Acids.
LO 5.8: Calculate formula mass.

Chapter 6: Chemical Composition
LO 6.1: Counting Atoms by the Gram.
LO 6.2: Counting Molecules by the Gram.
LO 6.3: Chemical Formulas as Conversion Factors: A. Molecular formulas give mole ratios of atoms in molecule; B. Converting moles of atoms to moles of molecules; C. Converting grams of atoms to grams of molecules.
LO 6.4: Mass Percent Composition of Compounds; Use mass percent composition as a conversion factor.
LO 6.5: Determine mass percent composition from a chemical formula.
LO 6.6: Determine an empirical formula from experimental data.
LO 6.7: Calculate a molecular formula from an empirical formula and molar mass.

Chapter 7: Chemical Reactions
LO 7.1: Identify evidence of a chemical reaction.
LO 7.2: Identify & write balanced chemical equations.
LO 7.3: Determine whether a compound is soluble.
LO 7.4: Predict and write equations for precipitation reactions.
LO 7.5: Write molecular, complete ionic, and net ionic equations.

Chapter 8: Quantities in Chemical Reactions
LO 8.1: Carry out mole-to-mole conversions between reactants and products based on the numerical relationship (coefficients) in a balanced chemical equation.
LO 8.2: Carry out mole-to-mass or mass-to-mole conversions between reactants and products in a balanced chemical equation and molar masses.
LO 8.3: Carry out mass-to-mass conversions between reactants and products in a balanced chemical equation and molar masses.
LO 8.4: Compare the possible number of product moles (or mass) to determine limiting reactant.
LO 8.5: Use limiting reactant amount to determine theoretical yield.
LO 8.6: Identify actual yield, theoretical yield, and calculate percent yield.

Chapter 13: Solutions
LO 13.1: Define solution, solute, and solvent.
LO 13.2: Relate the solubility of solids in water to temperature.
LO 13.3: Relate the solubility of gases in liquids to temperature and pressure.
LO 13.4: Calculate solution concentration using mass percent.
LO 13.5: Calculate solution concentration using molarity.

LO 13.6: Calculate ion concentration.
LO 13.7: Use the dilution equation in calculations.
LO 13.8: Use volume and concentration to calculate the number of moles of reactants or products and then use stoichiometric coefficients to convert to other quantities in a reaction.

Chapter 9: Electrons in Atoms and the Periodic Table
LO 9.1: Understand and explain the nature of electromagnetic radiation.
LO 9.2: Predict the relative wavelength, energy, and frequency of different types of light.
LO 9.3: Understand and explain the key characteristics of the Bohr model of the atom.
LO 9.4: Understand and explain the key characteristics of the quantum mechanical model of the atom
LO 9.5: Write electron configurations and orbital diagrams for atoms.
LO 9.6: Identify valence electrons and core electrons.
LO 9.7: Explain why the chemical properties of elements are largely determined by the number of valence electrons they contain.
LO 9.8: Identify and understand periodic trends in atomic size, ionization energy, and metallic character.

Appendix 2: Measures of Linked Concepts (MLC, including associated learning objectives)
Institution 1
Chapter 2
A titanium bicycle frame displaces 0.294 L of water and has a mass of 1.32 kg.
1. 0.294 L has 4 significant figures. (2.1)
2. 0.294 L is 2.94×10^{1} L. (2.4)
3. The unit prefix kilo means "x 10^{-3}." (2.5)
4. 1.32kg is 1.32×10^{6} mg. (2.6)
5. The density of the frame is 4.49 kg/L. (2.11)
6. Titanium and water are both chemicals. (0.0)
7. Using a balance to measure the mass of a titanium bicycle frame is 1.32kg is an observation. (0.1)

Chapter 3
A gasoline tank holds 43 kg of gasoline. When the gas burns, 175 kg of oxygen are consumed, and carbon dioxide and water are produced. The complete reaction releases 19.20 GJ of energy.
1. Water is an element. (1.2)
2. The combined mass of water and carbon dioxide is 218 kg. (1.6)
3. It is an endothermic reaction. (1.7)
4. Carbon dioxide is a gas when it leaves the car, and has molecules that are far apart from each other. (1.8)
5. 19.20 has 4 significant figures. (2.1)
6. This reaction theoretically could take place at absolute zero. (2.10)
7. 19.20 GJ means 1.920×10^{7} J of energy. (2.6)
8. Gasoline is a reactant, and oxygen, carbon dioxide, and water are the products. (11.1)

Chapter 4
An empty aluminum soda can have a mass of 14.9 grams. At 20.0 °C, aluminum has a density of 2.70 g/cm^3.
1. Aluminum ion has 10 electrons because it has gained 3 electrons. (3.2)
2. Aluminum-27 has 27 neutrons. (3.2)
3. Aluminum is a transition metal. (3.5)
4. Aluminum is in the same period as chlorine. (3.5)
5. The symbol for aluminum is Am. (3.7)
6. Aluminum is an element. (1.2)
7. Aluminum is a solid, so its atoms do not move around each other and it has fixed volume and rigid shape. (1.8)
8. 14.9 grams is the same as 14900 kg. (2.6)
9. The can has a volume of 40.2 cm^3. (2.11)
10. 20.0 °C = 293.2 K. (2.10)

Chapter 5
Distilled vinegar contains approximately 50.00 grams of $HC_2H_3O_2$ dissolved in a liter of water.
1. $HC_2H_3O_2$ (aq) is an oxyacid. (6.2)
2. The name of $HC_2H_3O_2$ (aq) acetic acid. (6.3)
3. Each of $HC_2H_3O_2$ molecule contains 3 hydrogen atoms. (6.8)
4. $HC_2H_3O_2$ contains the polyatomic ion carbonate. (6.9)
5. $HC_2H_3O_2$ is charge neutral. (1.2)
6. Vinegar is a heterogeneous mixture. (1.3)
7. 50.00 contain 2 significant figures. (2.1)
8. 50.00 grams equal to 5.000×10^{-5} μg. (2.6)
9. Oxide has 10 electrons. (3.2)

10. Oxygen is a noble gas. (3.5)

Chapter 6

0.09545 moles of $C_4H_6O_6$ are the white, powdery substance that coats tart candies such as Sour Patch Kids. Atomic mass: C: 12:01; H: 1.008; O: 15.99

1. The empirical formula of $C_4H_6O_6$ is $C_4H_6O_6$. (4.2)
2. 0.09545 moles of $C_4H_6O_6$ are equivalent to 14.32 grams of $C_4H_6O_6$. (4.4)
3. 0.09545 moles of $C_4H_6O_6$ contain 2.299 x 10^{23} atoms of carbon. (4.4)
4. $C_4H_6O_6$ is a white, powdery substance; this is a hypothesis. (0.1)
5. $C_4H_6O_6$ is a pure substance. (1.1)
6. The molecules in $C_4H_6O_6$ are closely packed and free to move around by each other. (1.8)
7. 0.09545 moles equal to 954.5 millimoles. (2.6)
8. Oxygen-18 has 18 neutrons. (3.2)
9. Carbon is a metalloid. (3.8)
10. $C_4H_6O_6$ is an ionic compound. (6.2)

Chapter 7

Barium ions are highly toxic to humans. Thus, barium chloride is toxic when consumed. One potential antidote of barium poisoning is magnesium sulfate, which reacts according to the reaction below:

barium chloride + magnesium sulfate --> barium sulfate + magnesium chloride

1. This is a double displacement reaction. (11.2)
2. The equation given above will need coefficients unequal to 1. (11.6)
3. Barium sulfate is insoluble, and thus magnesium sulfate in excess will remove all barium ions from solution. (11.7)
4. The oxidation number of sulfur in the sulfate ion is -2. (11.10)
5. This is a redox reaction. (11.13)
6. A physical change has happened in the above equation. (1.5)
7. Mg gains two electrons to form Mg^{2+}. (3.2)
8. Barium has 56 protons. (3.2)
9. 3.12 mol of barium chloride contains 221 grams of Cl. (4.4)
10. Barium chloride's formula is BaCl. (6.6)

Chapter 8

The explosion of fireworks usually contains the following reaction:

$$\mathbf{2KNO_3\ (s) + S\ (s) + 3C\ (s) \rightarrow K_2S\ (s) + N_2\ (g) + 3CO_2\ (g) + energy}$$

If 4.00 mol of KNO_3 reacts with 3.00 mol of S and 3.00 mol of C, 25.0 grams N_2 were produced.

1. The theoretical yield of CO_2 in grams is 132 g. (12.2)
2. The limiting reactant is S. (12.3)
3. The percent yield of N_2 is 10.8%. (12.6)
4. This is an endothermic reaction. (1.7)
5. 25.0 grams N_2 equals to 2.50 x 10^{-13} pg. (2.5)
6. Sulfur has 16 electrons. (3.2)
7. The number of oxygen atoms in 4.00 mol of KNO_3 is 7.23 x 10^{24}. (4.4)
8. The name for K_2S is dipotassium monosulfide. (6.3, 6.7)
9. Nitrogen is a diatomic element. (6.11)
10. If we put a small amount of the solid K_2S produced in the reaction into water, it will dissolve. (11.7)

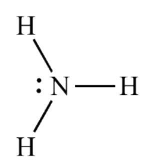
Chapter 9

15.88 nanograms of magnesium. Magnesium only occurs as magnesium-24, magnesium-25, and magnesium-26. The atomic mass of magnesium is 24.31 amu.

1. Magnesium has 6 electrons in s orbitals. (5.3)
2. A magnesium ion has the same electron configuration as a sodium ion. (5.3)
3. The electron configuration of neutral magnesium is [Ar] $3s^2$. (5.4)
4. The atomic size of magnesium is larger than beryllium. (5.6)
5. Magnesium is an alkali metal. (3.5)
6. Magnesium-24 is the most common isotope. (3.6)
7. The chemical symbol for magnesium is Mn. (3.7)
8. Atomic mass units are equivalent to grams/mole. (4.1)
9. There are 3.934×10^{14} atoms of magnesium in the sample. (4.4)
10. Magnesium typically is found in molecular compounds. (6.2)

Chapter 10

NH_3 is commonly used worldwide as a fertilizer. It is also a common cleaning agent when dissolved in water. The Lewis structure of NH_3 is provided below. The electronegativity of N is 3.0 and H is 2.1.



1. The molecular geometry of NH_3 is trigonal planar. (7.5)
2. Nitrogen will have a slightly positive charge. (7.7)
3. N-H bonds are polar. (7.8)
4. This is a nonpolar molecule. (7.9)
5. NH_3 (aq) is a homogeneous mixture. (1.3)
6. Every atom in NH_3 is a nonmetal. (3.8)
7. In one mole of NH_3 there are 6.022×10^{23} atoms of hydrogen. (4.4)
8. Nitrogen has 2 core electrons. (5.7)
9. There are 8 total valence electrons in all the atoms of NH_3. (5.7)
10. The name of NH_3 is ammonium. (6.3)

Chapter 11

SCUBA divers employ a rigid 11.0 L tank at 205 atm filled with heliox, a mixture of 95.0% helium and 5.00% oxygen. The ocean averages a temperature of 17 °C, and the tank will rapidly equilibrate to this temperature.

1. As the diver consumes the gas, the volume of gas in the tank will go down. (9.7)
2. If more heliox were added to the tank, the pressure would go down. (9.7).
3. There are approximately 360. grams of helium in the tank. (9.8, 9.6, 4.4)
4. The partial pressure of oxygen is about 0.0135 torr. (9.8, 2.7)
5. Helium is more electronegative than oxygen. (5.6)
6. Helium is the only noble gas that does not have 8 valence electrons. (5.7)
7. The molar mass of oxygen gas is 16.00 g/mol. (6.11, 4.1)
8. The hypothetical compound HeO would be named helium (II) oxide. (6.3)
9. He has no molecular geometry. (7.5)
10. Helium and oxygen have extremely weak intermolecular forces between each other. (8.3)

Chapter 13

When a patient is given an intravenous therapy (IV) in a hospital, the majority of the fluid is a saline solution containing 0.90 g sodium chloride per 100. mL solution (density=1.00 g/mL).

1. The mass percent of the saline solution is 9.0%. (10.5)
2. The molarity of the saline solution is 0.15 M. (10.5)
3. If we dilute 0.10 L of the saline solution to 1.0 L, the molarity for the diluted solution will be increased by 10 times. (10.10)
4. The saline solution is a strong electrolyte solution. (10.11)
5. Isotopes Cl-35 and Cl-37 have the same number of protons. (3.2)
6. The mass percent of O in water is about 89%. (4.5)
7. The chloride ion mimics the electron configuration of neon. (5.4)
8. Sodium chloride's chemical formula is SCl_2. (6.6)
9. Sodium chloride will dissolve in water because water is nonpolar. (7.10)
10. This solution will have stronger intermolecular forces than pure water. (8.3)

Institution 2
Chapter 2
A titanium bicycle frame displaces 0.294 L of water and has a mass of 1.32 kg.
1. 0.294 L is 2.94×10^{1} L. (LO 2.1)
2. 0.294 L has 4 significant figures. (LO 2.3)
3. If we round 0.294 to 2 significant figures, we will get 0.3. (LO 2.5)
4. The unit prefix kilo means "10^{-3}". (LO 2.9)
5. 0.294 L is the same as 294mL. (LO 2.9; 2.10)
6. 1.32 kg is the same as 1.32×10^{6} mg. (LO 2.9; 2.10)
7. The density of the titanium bicycle frame is 4.49 kg/L. (LO 2.11)
8. Titanium and water are both chemicals. (LO 1.2)
9. The mass of a titanium bicycle frame is 1.32kg is an observation. (LO 1.3)
10. Someone proposed that a titanium bicycle is lighter than a steel bicycle because it is more expensive. This is a theory. (LO 1.3)

Chapter 3
A gasoline tank holds 43 kg of gasoline. When the gas burns, 175 kg of oxygen are consumed, and carbon dioxide and water are produced. The complete reaction releases 19.20 GJ of energy.
1. Carbon dioxide is a gas when it leaves the car, and it has molecules that are far apart from each other. (LO 3.2)
2. Water is an element. (LO 3.3)
3. If we put 43 kg of gasoline and 175 kg of oxygen together in the same tank, this creates a homogeneous mixture. (LO 3.3)
4. Flammability of gasoline is a physical property. (LO 3.4)
5. The combined mass of water and carbon dioxide is 218 kg. (LO 3.6)
6. It is an endothermic reaction. (LO 3.10)
7. This reaction theoretically could take place at absolute zero. (LO 3.12)
8. 19.20 has 4 significant figures. (LO 2.3).
9. 19.20 GJ means 1.920×10^{7} J of energy. (LO 2.9)
10. 43 kg of gasoline is larger than 43 Mg gasoline. (LO 2.9)

Chapter 4
An empty aluminum soda can has a mass of 14.9 grams. At 20.0 °C, aluminum has a density of 2.70 g/cm^3. (K = 273.15 + °C)
1. The symbol for aluminum is Am. (LO 4.4)
2. Aluminum is a transition metal. (LO 4.5)
3. Aluminum is in the same period as Chlorine. (LO 4.5)
4. Aluminum ion has 10 electrons because it has gained 3 electrons. (LO 4.7)
5. Aluminum-27 has 27 neutrons. (LO 4.9)
6. 14.9 grams is the same as 14900 kg. (LO 2.9; LO 2.10)
7. The can has a volume of 40.2 cm^3. (LO 2.12)
8. Aluminum is a solid, so its atoms do not move around each other and it has fixed volume and rigid shape. (LO 3.2)
9. Aluminum is a pure substance. (LO 3.3)
10. 20.0 °C = 293.2 K. (LO 3.11)

Chapter 5
Distilled vinegar contains approximately 50.00 grams of $HC_2H_3O_2$ dissolved in a liter of water. C: 12:01; H: 1.008; O: 15.99
1. Each of $HC_2H_3O_2$ molecule contains 3 hydrogen atoms. (LO 5.2)
2. $HC_2H_3O_2$ contains the polyatomic ion carbonate. (LO 5.5)

3. The name of $HC_2H_3O_2$ (aq) acetic acid. (LO 5.7)
4. $HC_2H_3O_2$ (aq) is an oxyacid. (LO 5.7)
5. The formula mass of $HC_2H_3O_2$ is 60.032 amu. (LO 5.8)
6. 50.00 contain 2 significant figures. (LO 2.3)
7. 50.00 grams equal to 5.000 x 10^{-5}μg. (LO 2.9)
8. Vinegar is a heterogeneous mixture. (LO 3.3)
9. Oxygen is a noble gas. (LO 4.5)
10. Oxide has 10 electrons. (LO 4.7)

Chapter 6

0.09545 moles of $C_4H_6O_6$ are the white, powdery substance that coats tart candies such as Sour Patch Kids. Atomic mass: C: 12:01; H: 1.008; O: 15.99; Avogadro's number = 6.022x10^{23} atoms

1. 0.09545 moles of $C_4H_6O_6$ are equivalent to 14.32 grams of $C_4H_6O_6$. (LO 6.2)
2. 0.09545 moles of $C_4H_6O_6$ contain 2.299 x 10^{23} atoms of carbon. (LO 6.1; LO 6.3)
3. $C_4H_6O_6$ is a white, powdery substance; this is a hypothesis. (LO 1.3)
4. 0.09545 moles equal to 954.5 millimoles. (LO 2.9)
5. The molecules in $C_4H_6O_6$ are closely packed and free to move around by each other. (LO 3.2)
6. $C_4H_6O_6$ is a pure substance. (LO 3.3)
7. Carbon is a metalloid. (LO 4.5)
8. Oxygen-18 has 18 neutrons. (LO 4.9)
9. The empirical formula of $C_4H_6O_6$ is $C_4H_6O_6$. (LO 5.2)
10. $C_4H_6O_6$ is an ionic compound. (LO 5.5)

Chapter 7

$BaCl_2$ is highly toxic, $MgSO_4$ is potential antidotes because $BaCl_2$ reacts with $MgSO_4$ can form $BaSO_4$, which is relatively non-toxic because of its insolubility. 0.00435mol of $MgSO_4$ reacts with 3.12 mol of $BaCl_2$. The complete ionic equation is Mg^{2+}(aq) + SO_4^{2-} (aq) + Ba^{2+}(aq) + 2Cl^-(aq) → Mg^{2+}(aq) + 2Cl^- (aq) + $BaSO_4$ (s). Atomic mass: Ba: 137.2 Cl: 35.45

1. The equation given above is balanced. (LO 7.2)
2. The coefficient for SO_4^{2-} in the equation is 4. (LO 7.2)
3. The type of reaction shown above is a precipitation reaction. (LO 7.4)
4. Ba^{2+} is the spectator ion in the equation. (LO 7.5)
5. 0.00435 mol of Mg is written in scientific notation as 4.35 x 10^{-3}. (LO 2.1)
6. A physical change has happened in the above equation. (LO 3.5)
7. Chlorine has 17 protons. (LO 4.4)
8. Mg gains two electrons to form Mg^{2+}. (LO 4.6)
9. The name of $BaSO_4$ is Barium Sulfate. (LO 5.5)
10. 3.12 mol of $BaCl_2$ contains 221 grams of Cl (LO 6.1; 6.3)

Chapter 8

The explosion of fireworks usually contains the following reaction:

$2KNO_3$ (s) + S (s) + 3C (s) → K_2S (s) + N_2 (g) + $3CO_2$ (g)

If 4.00 mol of KNO_3 reacts with 3.00 mol of S and 3.00 mol of C, 25.0 grams N_2 were produced. Avogadro's number = 6.022 x10^{23} K: 39.10 N:14.01 O: 15.99 S:32.07 C:12.01

1. The limiting reactant is S. (LO 8.4)
2. The theoretical yield of CO_2 in grams is 132 g. (LO 8.5)
3. The percent yield of N_2 is 10.8%. (LO 8.6)
4. 25.0 grams N_2 equals to 2.50 x 10^{-13} pg. (LO 2.9)
5. This is an endothermic reaction. (LO 3.10)
6. Sulfur has 16 electrons. (LO 4.2)
7. Nitrogen is a diatomic element. (LO 5.3)

8. The name for K_2S is dipotassium monosulfide. (LO 5.5)
9. The number of oxygen atoms in 4.00 mol of KNO_3 is 7.23 x 10^{24}. (LO 6.1; 6.3)
10. If we put a small amount of the solid K_2S produced in the reaction into water, it will dissolve. (LO 7.3)

Chapter 13

When a patient is given an intravenous therapy (IV) in a hospital, the majority of the fluid is a saline solution containing 0.90 g sodium chloride per 10<u>0</u> mL solution (density=1.00 g/mL). Sodium chloride (aq) reacts with lead(II) nitrate (aq) will form lead(II) chloride and sodium nitrate. Na: 22.99 Cl:35.45 N: 14.01 O: 15.99

1. The saline solution is a strong electrolyte solution. (LO 13.2)
2. The mass percent of the saline solution is 9.0%. (LO 13.4)
3. The molarity of the saline solution is 0.15 M. (LO 13.5)
4. If we dilute 0.10 L of the saline solution to 1.0 L, the molarity for the diluted solution will be increased by 10 times. (LO 13.7)
5. Isotopes Cl-35 and Cl-37 have the same number of protons. (LO 4.7)
6. The Chemical formula for Lead (II) nitrate is Pb_2NO_3. (LO 5.4)
7. Mass percent of O in sodium nitrate (keep 2 sig.figs) is 19%. (LO 6.5)
8. Sodium nitrate will dissociate into Na^+, N^{3-}, and O^{2-} in water. (LO 7.4)
9. Sodium chloride reacts with Lead (II) nitrate to form a precipitate. (LO 7.4)
10. The spectator ions for the chemical equation when sodium chloride (aq) reacts with lead (II) nitrate (aq) are: sodium ion and lead ion. (LO 7.5)

Chapter 9

15.88 nanograms of magnesium. Magnesium only occurs as magnesium-24, magnesium-25, and magnesium-26. The atomic mass of magnesium is 24.31 amu.

1. The electron configuration of neutral magnesium is [Ar] $3s^2$. (LO 9.5)
2. Magnesium has 6 electrons in s orbitals. (LO 9.5)
3. A magnesium ion has the same electron configuration as a sodium ion. (LO 9.5)
4. The atomic size of magnesium is larger than beryllium. (LO 9.8)
5. The chemical symbol for magnesium is Mn. (LO 4.4)
6. Magnesium is an alkali metal. (LO 4.5)
7. Magnesium-24 is the most common isotope. (LO 4.10)
8. Magnesium typically is found in molecular compounds. (LO 5.5; 5.6)
9. Atomic mass units are numerically equivalent to grams/mole. (LO 5.8)
10. There are 3.934 x 10^{14} atoms of magnesium in the sample. (LO 2.9; 6.1)

Appendix 3 Survey questions and interview protocol

Survey Questions

1. Do you think the true/false/unsure type of questions on quizzes helped you make connections among topics in this class?
2. Please explain why you chose the answer for question 1.
3. How do the true/false/unsure type of questions help you understand chemistry conceptually?
4. Do you think the instruction of the true/false/unsure type of questions is clear enough for you to answer the questions?
5. If you answer was "no" for question 4, what suggestions do you have to improve the instruction?
6. Did you find the "unsure" option to be helpful?
7. Please explain why you chose the answer for question 6.
8. How easy did you find it to answer the true/false/unsure type of questions?
   Do you think the reflection questions in the end of the quizzes are helpful? Here are the reflection questions: “Which learning objectives did you miss on the last [MLC] quiz? What is your plan to master these learning objectives in the future?”
9. Have you ever looked at the learning objectives that you didn’t perform well on the quizzes?
10. Have you ever developed a plan for the learning objectives that you didn’t perform well on the quizzes?
11. If you answer "Yes" for question 10, what plans do you usually have?
12. How did the reflection questions help you evaluate your learning in this course?
13. How did you prepare for the true/false/unsure type of questions on quizzes?
14. Did the true/false/unsure type of questions change the way you think chemistry as a subject?
15. Please explain why for question 14.
16. Any additional comments about the true/false/unsure type of questions?

Interview protocol

First, before the interview starts, the interviewer and participants briefly introduce themselves to each other. Then, the interviewers ask participants to complete the example assessment with a statement below: *“Please don’t worry about your answers as correct or incorrect, we just want to know how you solved the problem.”*

Example assessment: MLC Chapter 13 Solutions

When a patient is given an intravenous therapy (IV) in a hospital, the majority of the fluid is a saline solution containing 0.90 g sodium chloride per 100 mL solution (density=1.00 g/mL). Sodium chloride (aq) reacts with lead (II) nitrate (aq) will form lead (II) chloride and sodium nitrate. Na: 22.99 Cl:35.45 N: 14.01 O: 15.99

1. The saline solution is a strong electrolyte solution. (LO 13.2)

2. The mass percent of the saline solution is 9.0%. (LO 13.4)

3. The molarity of the saline solution is 0.15 M. (LO 13.5)

4. If we dilute 0.10 L of the saline solution to 1.0 L, the molarity for the diluted solution will be increased by 10 times. (LO 13.7)

5. Isotopes Cl-35 and Cl-37 have the same number of protons. (LO 4.7)

6. The Chemical formula for Lead (II) nitrate is Pb_2NO_3*. (LO 5.4)*

7. Mass percent of O in sodium nitrate (keep 2 sig. figs) is 19%. (LO 6.5)

8. sodium nitrate will dissociate into Na^+*,* N^{3-}*, and* O^{2-} *in water. (LO 7.4)*

9. Sodium chloride reacts with Lead (II) nitrate to form a precipitate. (LO 7.4)

10. The spectator ions for the chemical equation when sodium chloride(aq) reacts with lead(II) nitrate (aq) are: sodium ion and lead ion. (LO 7.5)

Answer these reflection questions on your own after viewing the correct answers: (1) What learning objectives did you perform incorrectly for this assessment? please write them down. (2) What is your plan to master these learning objectives in the future?

After the participants complete the assessment, the following interview questions are asked:

1. Can you talk about each statement and explain why you put them as true/false/unsure?
2. Have you ever taken chemistry courses before? What are they? What types of assessments were used in your previous chemistry courses?
3. How confident are you when solving chemistry problems in general?
4. Do you consider chemistry topics as a unified theme or as separate facts?
5. Do you think the above assessment help you make connections among chemistry topics in the courses? Why?
6. To what extent the above question help you learn chemistry conceptually?
7. What do you think about the reflection questions? Have you ever looked at those learning objectives that you didn't perform well on the assessment? How did they help you evaluate/reflect your learning in the courses?
8. How easy did you find it to answer this type of assessment? Do you believe you do better on this assessment or other types of assessments?
9. Can you please talk about how you prepared for this type of assessment on quizzes? Are there any differences when you prepared for this type of assessment versus other types of questions such as multiple-choice questions? How do the above types of questions impact your learning approaches in chemistry courses?
10. Would you like to see this type of assessment in your future chemistry classes? Why?
11. If you have choice, do you prefer this type of assessment or other types of assessment for homework assignments or quizzes or exams? Why?
12. Any additional comments about this type of assessment or suggestions for improving this type of assessment?