Exam 1

# Part 1: Multiple Choice (2 points each)

## Directions: Please circle the *best* answer for each of the following questions.

1. The scientific method
	1. is just a theory.
	2. is a strict set of rules and procedures that lead to inarguable fact.
	3. isn't used much in modern chemistry.
	4. is based on continued observation and experiment.
	5. is a framework for proving an argument you know to be true.
2. Which of the following represents a physical property?
	1. Sodium metal is extremely reactive with chlorine gas.
	2. Mercury is a silvery liquid at room temperature.
	3. Aluminum has a tendency to “rust.”
	4. Argon has an unreactive nature.
	5. none of the above
3. What is the empirical formula for Hg2(NO3)2?
	1. Hg2(NO3)2
	2. HgNO3
	3. Hg(NO3)2
	4. Hg2NO3
	5. Hg4(NO3)4
4. What is the name for Hg2(NO3)2
	1. Mercurous nitrate
	2. Mercury(II) nitrite
	3. Mercury(I) nitrite
	4. Mercurous nirate
	5. a and d
5. The boiling point of neon is 27 K or
	1. -401 °F
	2. -105 °F
	3. -246 °C
	4. 300 °C
	5. none of the above
6. Choose the statement below that is true.
	1. A weak acid solution consists of mostly nonionized acid molecules.
	2. The term “strong electrolyte” means that the substance is extremely reactive.
	3. A strong acid solution consists of only partially ionized acid molecules
	4. The term “weak electrolyte” means that the substance is inert.
	5. A molecular compound that does not ionize in solution is considered a strong electrolyte.

1. Which isotope has 43 protons, 43 electrons, and 56 neutrons?
	1. Barium-99
	2. Neodymium-99
	3. Technicium-99
	4. Scandium-43
	5. none of the above
2. The best way to avoid inhalation of volatile chemicals while working with them it is to
	1. work in a chemical hood when using them.
	2. avoid breathing the vapors by holding beakers and flasks at arm’s length.
	3. wear a common dust mask.
	4. always keep beakers, flasks, and bottle of the chemical capped or covered.
	5. none of the above
3. Carry out the calculation below, paying special attention to significant figures, round, and units.

$$\frac{4.32×10^{7} g}{\frac{4}{3}(3.1416)(1.95×10^{2} cm)^{3} }=$$

* 1. 4.3695902
	2. 4.37 g/cm3
	3. 1.66 × 105 g/cm
	4. 1.66 × 1013 g cm3
	5. 2.01 × 102 g/cm3
1. Consider the following reaction: 2 CH3OH (g) + 3 O2 (g) → 2 CO2 (g) + 4 H2O (g)

Each of the following molecular diagrams represents an initial mixture of the reactants:

How many carbon dioxide molecules would be formed from the reaction mixture that produces the greatest amount of products?

1. 1 molecule
2. 2 molecules
3. 3 molecules
4. 4 molecules
5. not enough information

# Part 2: Short Answer

## Directions: Answer each of the following questions. Be sure to use complete sentences where appropriate. For full credit be sure to show all of your work.

1. What kind of information is needed to formulate a hypothesis (3 points)?
2. A brief winter storm leaves a dusting of snow on the ground. During the sunny but very cold day after the storm, the snow disappears even though the air temperature never gets above freezing. If the snow didn’t melt, where did it go (2 points)?
3. Can an extensive property be used to identify a substance? Explain why or why not (2 points).
4. Convert 46 μs to ps (3 points).
5. State the mass law(s) demonstrated by the following experimental results, and explain your reasoning (8 points).

Experiment 1: A student heats 1.00 g of a blue compound and obtains 0.64 g of a white compound and 0.36 g of a colorless gas.

Experiment 2: A second student heats 3.25 g of the same blue compound and obtains 2.08 g of a white compound and 1.17 g of a colorless gas.

1. Tellurium has an average atomic mass of 127.60 amu, which is heavier than iodine, whose mass is 126.90447 amu, yet iodine comes after tellurium on the modern periodic table. Why does this occur? Find one other instance where heavier atoms precede lighter ones on the periodic table (4 points).
2. Explain how the results of the gold foil experiment led Rutherford to dismiss the plum pudding model of the atom and create his own model based on a nucleus surrounded by electrons (3 points).
3. An element has two naturally occurring isotopes. Isotope 1 has a mass of 120.9038 u and a relative abundance of 57.4%, and isotope 2 has a mass of 122.9042 u (8 points).
	1. What is the percent abundance of the second isotope?
	2. What is the atomic mass of the element?
	3. What is the name and symbol of the element? \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
4. How is it that all soluble ionic compounds are electrolytes but soluble molecular compounds may or may not be electrolytes (3 points)?
5. Interpretation of Reactions by Ionic Type Equations. Aqueous solutions of the following substances or their mixtures with water if they are only slightly soluble, are mixed. Write first the conventional equation, second the total ionic equation, and lastly the net ionic equation. If you predict no appreciable reaction, indicate this, and state why (6 points).
6. Magnesium nitrate and zinc chloride
7. Calcium carbonate and acetic acid
8. Ethylene gas, C2H4, reacts with water at high temperature to yield ethyl alcohol, C2H6O.
	1. Write the balanced chemical equation (10 points).
	2. How many grams of carbon are in 15.43 g of ethylene?
	3. How many grams of ethyl alcohol will result from the reaction of 1.33 × 1024 molecules of water with excess ethylene gas?
9. A solution of hydrochloric acid is prepared by bubbling hydrogen chloride gas into water. If the resulting solution has a pH of 0.9745 (10 points)
	1. what is the hydrogen ion concentration of the solution?
	2. If 5.00 mL of this solution is dilute to 25.00 mL, what is the concentration of the new solution?
10. Ammonia gas is produced via the Haber process when hydrogen gas is reacted with nitrogen gas (12 points).
	* + - 1. Write a balanced chemical equation for the reaction.
				2. Identify the type of reaction: \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_
				3. Complete an ICE table when 30.0 g of nitrogen react with 5.02 g of hydrogen.
				4. Calculate the grams of ammonia produced.
				5. How many grams of the excess reactant remain at the end of the reaction?