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. 1.3695902
	2. 1.37 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)?

To form a hypothesis we need at least one observation, experiment, or idea (from examining nature).

1. 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)?

The snow sublimed to form water vapor.

1. Can an extensive property be used to identify a substance? Explain why or why not (2 points).

No, extensive properties change with the size of the sample and so cannot be used to identify a substance.

1. Convert 46 μs to ps (3 points).

$$46 μs×\frac{10^{12} ps}{10^{6} μs}=4.6×10^{7} ps$$

1. 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.

The two experiments demonstrate the law of definite composition. The unknown compound decomposes the same way both times.

Experiment 1: $\frac{0.64 g solid}{0.36 g gas}=1.8$

Experiment 2: $\frac{2.08 g solid}{1.17 g gas}=1.8$

The experiments also demonstrate the law of conservation of mass since the total mass before reaction equals the total mass after reaction.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Experiment | Blue compound $→$ | White compound + | Colorless gas |  |
| 1 | 1.00 g =  | 0.64 g +  | 0.36 g =  | 1.00 g |
| 2 | 3.25 g =  | 2.08 g +  | 1.17 g =  | 3.25 g |

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).

The atomic mass is an average weighted mass of all of the naturally occurring isotopes of an element. The modern periodic table is ordered by increasing atomic number (number of protons).

Argon (Ar, 39.948 amu) to potassium (K, 39.098 amu)

Cobalt (Co, 58.933 amu) to nickel (Ni, 58.693 amu)

Thorium (Th, 232.04 amu) to Protactinium (Pa, 231.04 amu)

Uranium (U, 238.029) to Neptunium (Np, (237))

Plutonium (Pu, (244)) to Americium (Am, (243))

Darmstadtium (Ds, (281)) to Roentgenium (Rg, (272))

Flerovium (Fl, (289)) to Uup (Ununpentium, (288))

1. 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).

Rutherford concluded that the positive charge in the atom could not be spread out (to form the pudding) in the atom, but must result from a concentration of charge in the center of the atom (the nucleus). Most of the alpha particles were deflected only slightly or passed directly through the gold foil, so Rutherford reasoned that the nucleus must be small compared to the size of the entire atom. The negatively charged electrons do not deflect the alpha particles, and Rutherford reasoned that the electrons took up the remainder of the space of the atom outside of the nucleus.

1. 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?

$$100\%-57.4\%=42.6\%$$

* 1. What is the atomic mass of the element?

$$atomic mass=\sum\_{}^{}\left(m\_{isotope}\right)\left(\frac{\%abundance\_{isotope}}{100}\right)$$

$$atomic mass=\left(120.9038 u\right)\left(\frac{57.4}{100}\right)+(122.9041 u)\left(\frac{42.6}{100}\right)$$

$$atomic mass=69.3987812 u+ 52.3571466 u$$

$$atomic mass=121.7559278 u≈121.8 u$$

* 1. What is the name and symbol of the element? \_\_\_\_\_antimony, Sb
1. How is it that all soluble ionic compounds are electrolytes but soluble molecular compounds may or may not be electrolytes (3 points)?

Solutions of weak electrolytes conduct electricity poorly because only small quantities of ions form when the electrolyte dissolves. Solutions of nonelectrolytes do not conduct electricity because essentially no ions exist in the solution.

1. 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).
2. Magnesium nitrate and zinc chloride

Mg(NO3)2 (aq) + ZnCl2 (aq) → MgCl2 (aq) + Zn(NO3)2 (aq)

Mg2+ (aq) + 2 NO3- (aq) + Zn2+ (aq) + 2 Cl- (aq) → Mg2+ (aq) + 2 Cl- (aq) + Zn2+ (aq) + 2 NO3- (aq)

No reaction

1. Calcium carbonate and acetic acid

 CaCO3 (s) + 2 HC2H3O2 (aq) → Ca(C2H3O2)2 (aq) + H2O (l) + CO2 (g)

 CaCO3 (s) + 2 HC2H3O2 (aq) → Ca2+ (aq) + 2 C2H3O2- (aq) + H2O (l) + CO2 (g)

 CaCO3 (s) + 2 HC2H3O2 (aq) → Ca2+ (aq) + 2 C2H3O2- (aq) + H2O (l) + CO2 (g)

1. Ethylene gas, C2H4, reacts with water at high temperature to yield ethyl alcohol, C2H6O.
	1. Write the balanced chemical equation (10 points).

C2H4 (g) + H2O (l) → C2H6O (l)

* 1. How many grams of carbon are in 15.43 g of ethylene?

 $15.43 g C\_{2}H\_{4}×\frac{1 mol C\_{2}H\_{4} }{28.054 g C\_{2}H\_{4}}×\frac{2 mol C}{1 mol C\_{2}H\_{4}}×\frac{12.011 g C}{1 mol C}=13.21 g C$

* 1. How many grams of ethyl alcohol will result from the reaction of 1.33 × 1024 molecules of water with excess ethylene gas?

$1.33×10^{24} molecules H\_{2}O×\frac{1 mol H\_{2}O }{6.022 ×10^{23} molecles H\_{2}O}×\frac{1 mol C\_{2}H\_{6}O}{1 mol H\_{2}O}×\frac{46.069 g C\_{2}H\_{6}O}{1 mol C\_{2}H\_{6}O}=101.746649 g C\_{2}H\_{6}O≈102 g C\_{2}H\_{6}O$

1. 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?

$$\left⌈H^{+}\right⌉=10^{-pH}=10^{-0.9745}=0.1060 M $$

* 1. If 5.00 mL of this solution is dilute to 25.00 mL, what is the concentration of the new solution?

M1 = 0.1060 M

V1 = 5.00 mL

V2 = 25.00 mL

M2 = ?

$$M\_{1}V\_{1}=M\_{2}V\_{2}⇒M\_{2} =\frac{M\_{1}V\_{1}}{V\_{2}}=\frac{(0.1060 M)(5.00 mL)}{(25.00 mL)}=0.0212 M $$

1. 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.

3 H2 (g) + N2 (g) → 2 NH3 (g)

* + - * 1. Identify the type of reaction: \_\_\_\_\_\_\_combination reaction\_\_\_\_\_\_\_\_
				2. Complete an ICE table when 30.0 g of nitrogen react with 5.02 g of hydrogen.

|  |  |  |  |
| --- | --- | --- | --- |
|  | N2 (g) + | 3 H2 (g) → | 2 NH3 (g) |
| I | $$30.0 g N\_{2}×\frac{1 mol N\_{2}}{28.014 g N\_{2}}=1.07 mol$$ | $$5.02 g H\_{2}×\frac{1 mol H\_{2}}{2.016 g H\_{2}}=2.49 mol$$ | 0 mol  |
| C | -x | -3x | +2x |
| E | 1.07 mol – x =1.07 mol – 0.830 mol =0.24 mol | 2.49 mol – x =0 mol | 2x = 2(0.830 mol) =1.66 mol  |

Compare ratios:$actual \frac{2.49 mol H\_{2} }{1.07 mol N\_{2} }=\frac{2.33 mol H\_{2} }{1mol N\_{2} }<theoretical \frac{3 mol H\_{2} }{1mol N\_{2} } $

Therefore, hydrogen is the limiting reagent. 2.49 mol – 3 x = 0 mol

 x = 0.830 mol

* + - * 1. Calculate the grams of ammonia produced.

$$1.66 mol NH\_{3}×\frac{17.031 g NH\_{3}}{1 mol NH\_{3}}=28.3 g NH\_{3}$$

* + - * 1. How many grams of the excess reactant remain at the end of the reaction?

$$0.24 mol N\_{2}×\frac{28.014 g N\_{2}}{1 mol N\_{2}}=6.7 g N\_{2}$$