Name Date

**Answers to Introduction to Chemistry Study Guide**

1. Chemistry is the study of the composition, structure, and properties of matter, the processes that matter undergoes, and the energy changes that accompany these processes.
2. Classify the following as qualitative or quantitative.
	1. Quantitative
	2. Qualitative
	3. Qualitative
	4. Qualitative
	5. Qualitative
3. Give an example of the following terms
	1. An apple
	2. Gold, Au
	3. Sand and water
	4. Air
	5. Soda
4. Intensive properties do not depend on the amount of matter. Extensive properties do depend on the amount of matter
	1. Extensive
	2. Extensive
	3. Intensive
	4. Extensive
5. Elements are made up of only one type of atom such as helium, He. Compounds are two or more elements chemically bonded such as carbon dioxide, CO2.
6. There are approximately 100 elements
7. Physical or Chemical Properties?
	1. Chemical
	2. Chemical
	3. Physical
	4. Physical
	5. Chemical
8. Physical or Chemical Changes
	1. Physical
	2. Chemical
	3. Physical
	4. Chemical
	5. Physical
9. Element, Compound, Homogeneous, Heterogeneous
	1. Homogeneous
	2. Element
	3. Compound
	4. Heterogeneous
	5. Heterogeneous
	6. Homogeneous
	7. Heterogeneous
	8. Homogeneous
10. SI units
	1. Meters
	2. Seconds
	3. Kelvin
	4. Kilograms
11. Common units
	1. Grams
	2. Liters or milliliters
	3. g/mL or g/cm3
	4. celcius
12. Density = mass D = 586 g = 3.12 g/mL

 Volume 188 mL

1. mass = density x volume m = (0.798 g/mL)(17.4 g) = 13.8 g
2. a. 2.7 x 10-7

 b. 3.56 x 102

 c. 4.7764 x 104

 d. 9.4 x 10-2

1. a. 0.0152

 b. 778000

 c. 6750

 d. 0.00082

1. number of sig figs
	1. 2
	2. 5
	3. 3
	4. 2
	5. 3
	6. 4
	7. 1
	8. 2
2. round to 2 sig figs
	1. 0.0045 m
	2. 560000 m
	3. 0.81 m
	4. 0.0046 m
	5. 3.0 x 102 m
3. Calculations
	1. 1700 cm2
	2. 0.94 m
	3. 1.8 x 10-2
	4. 10.6 m
	5. 0.79 g
	6. 16.5 cm g
	7. 1 x 106 g/cm3
4. Student A – not precise, not accurate

Student B – precise and accurate

Student C – precise but not accurate

% error = accepted value – experimental value / accepted value \* 100

 87.0 mL – 87.6 / 87.0 x 100 = 0.698 %

1. a. 23.4 g = 0.00234 kg

 b. 5.23 km = 5230 m

 c. 7.34 mg = 0.00734 g

 d. 7.3 L = 7300 mL

HISTORY OF THE ATOM

1. A philosopher who first proposed the idea that matter is made of indivisible particles called “atomos”.
2. The basic particle of matter is called the atom because the greek word for indivisible is “atomos”.
3. The idea of the atom didn’t occur at this time because there was no evidence from experimentation. The philosopher whose ideas prevailed was Aristotle. The basic elements in nature were Fire, Water, Air, and Earth.
4. Law of Conservation of Mass states that matter cannot be created nor destroyed.
5. Five things Dalton stated:
	1. All matter is composed of tiny particles called atoms
	2. Atoms of a give element are identical in size, mass, and other properties
	3. Atoms cannot be subdivided, create, or destroyed
	4. Atoms of different elements combine in simple whole number ratios
	5. In chemical reactions, atoms are combined, separated or rearranged.
6. The two principals that have been revised are b and c above.
7. Thomson
8. Cathode ray tube
9. Plum pudding model
10. Oil Drop Experiment
11. Robert Millikan
12. Rutherford

This experiment altered Thomson’s model by including all of the protons in the small, dense, center called the nucleus

1. Neutrons
2. The atom is neutral because the number of positively charged particles, protons, equals the negatively charged electrons.
3. The main regions of the atom are the nucleus and the electron cloud.
4. Protons + nucleus, Electrons - electron cloud, Neutrons 0 charge nucleus
5. The nucleus is a small, dense, positively charged center the contains the protons and neutrons.
6. Isotopes are atom with the same numbers of protons but different masses due to the different numbers of neutrons.
7. The atomic numbers for an atom is found on the periodic table. The atomic number tells us the number of protons in the nucleus of an atom.
8. A mass number is the sum of the protons and neutrons in the nucleus. These numbers are not on the periodic table
9. Average atomic mass is the weighted average of all the naturally occurring isotopes of an element. The formula for average atomic mass is:

(% abundance) ( mass of isotope 1) + (% abundance) (mass of isotope 2) +…

 100

1. An amu is 1/12th the mass of the carbon-12 atom. The carbon-12 atom is the reference atom.
2. Iron - 56 $\begin{matrix}56\\26 \end{matrix}$ Fe
3. Zinc – 66 30 protons, 30 electrons and 36 neutrons
4. (75.55%) (34.969 amu) + (24.23%) (36.966 amu) = 35.376 amu

100