Chemistry 1st semester final exam review Part 1

Unit 1 & 2 (Chapters 1, 2, 3)

Unit 1, Chapter 1 & 3 Introduction to Chemistry

1. Define the following words:
    a) Accuracy : how close a measurement is to the TRUE value (must have true value to measure)
    b) Precision: how close measurements are to each other (must have more than one measurement)
    c) Scientific Theory: well-tested explanation of some aspect of the natural world, attempts to explain why something occurs/happens
    d) Scientific Law: a concise statement that summarizes the results of many observations and experiments, does NOT attempt to explain why.
2. What is the difference between a scientific law and a scientific theory? A theory attemps to explain why, a law does not.
3. Describe the following measurement in terms of accuracy and precision: A laboratory analyzed a standard known to contain 140 ppb lead. The following results were obtained.

Trial	ppb Pb
1	169
2	114
3	142
4	115

The measurements were neither accurate nor precise. Only one measurement (trail 3) was close to the actual value. And only two measurements (Trail 2 & 4) were close to each other.

4. Describe the following measurement in terms of accuracy and precision:
    a. Measured value: 5.2 mg known value = 5.0 mg Accurate, NOT precise
    b. Measured value: 5.2 mg, 4.9 mg, 5.1 mg known value = 5.0 mg Both
    c. Measured values: 6.61 mL, 6.99 mL, 7.25 mL NOT Accurate, yet precise
    d. Measured value: 2.134 g/cm^3 Neither accurate, nor precise
5. How many significant figures are in the following numbers?
    Significant figure rules summarized:
    numbers are always significant,
    zeros in between significant numbers are significant
    zeros at the END of the number AND after the decimal DO count
    zeros at the end of a number WITH OUT a decimal don't count
    zeros at the beginning of the number with a decimal don't count

1) 7100	2	2) 260.0	4	3) 0.00010	2	4) 218	3
5) 320	2	6) 0.00530	3	7) 22,568	5	8) 4,755.50	6

6. Express the Following in Scientific Notation:
    a) 0.000 033 3.3×10^{-5} c) 55 000 000 5.5×10^{7} e) 0.000 00733 7.33×10^{-6}
    b) 8 200 000 8.2×10^{6} d) 0.00288 2.88×10^{-3} f) 65 000 6.5×10^{4}
7. Convert the following metric measurements. Show your work:
    a. 1.6 cs to Ks $1.6\ cs \times \frac{10^{-2}s}{1\ cs} \times \frac{1Ks}{10^{3}s} = 1.6 \times 10^{-5}Ks$
    b. 8.8 x 10^{-2} mm to Mm $8.8 \times 10^{-2}\ mm \times \frac{10^{-3}m}{1\ mm} \times \frac{1Mm}{10^{6}m} = 8.8 \times 10^{-11}Mm$
    c. 6.1 x 10^{21} daJ to hJ $6.1 \times 10^{21} daJ \times \frac{1hJ}{10\ daJ} \times = 6.1 \times 10^{20} TJ$

d. 9.5 x 10^{-23} ML to mL $9.5 \times 10^{-23}\, ML \times \frac{10^6 L}{1\, ML} \times \frac{1 mL}{10^{-3} L} = 9.5 \times 10^{-14} mL$

e. 2.35 x 10^8 cg to dag $2.35 \times 10^8\, cg \times \frac{10^{-3} g}{1\, cg} \times \frac{1 dag}{10 g} = 2.35 \times 10^4 dag$

f. 1.78 x 10^8 mg to g $1.78 \times 10^8\, mg \times \frac{10^{-3} g}{1\, mg} = 1.78 \times 10^5 g$

g. 274300 hJ to J $274300\, TJ \frac{10^{12} J}{1\, TJ} = 2.743 \times 10^{17} J$

h. 0.00432 daJ to KJ $0.00432\, daJ \times \frac{10^1 J}{1\, daJ} \times \frac{1 KJ}{10^6 J} = 4.32 \times 10^{-8} GJ$

8. Perform the following calculations and show your answers with the correct number of significant figures.
When multiplying or dividing you round to the SMALLEST number of significant figures.
   a) (4.0 x 10^3 mm) x (1.5 x 10^2 mm) = 6.0 x 10^5 (must have 2 sigfigs)
   b) (5.5 x 10^5 Km^3) / (3.3 x 10^3Km) = 1.7 x 10^2 (must have 2 sigfigs)
   c) 596,000 mg^2 ÷ 0.0023 mg = 2.6 x 10^8 (must have 2 sigfigs)
   d) 6.77 kg x 0.9 kg = 6 (must have 1 sigfig)

9. Find the density of the following items: SHOW WORK **round according to significant figure rules**
Density = mass/volume **volume = length x width x height**
Volume by displacement = final volume – initial volume
   a. What is the density of an object having a mass of 4.0 g and a volume of 39.0 cubic centimeters?
      D=4.0g/39.0cm^3 = 0.10 g/cm^3
   b. What is the volume of an object with a density of 7.73 g/cm^3 and a mass of 5.4010 g?
      Volume = mass/density V=5.4010g/(7.73g/cm^3) = 0.699 g
   c. A cube of a gold-colored metal with a volume of 59 cm^3 has a mass of 980 g. The density of pure gold is 19.3 g/cm . Is the metal pure gold? Show calculations to justify your answer.
      D= 980g/59cm^3= 17 g/cm^3 object is NOT gold b/c density is not close enough
   d. The density of osmium, which is the densest metal, is 22.57 g/cm . What is the mass of a block of osmium that measures 1.00 cm by 4.00 cm by 2.50 cm?
      V= 1.00 cm x 4.00 cm x 2.50 cm = 10.0 cm3
      Mass =density x volume M = 22.57 g/cm^3 x 10.0 cm^3 = 226 g
   e. A cup of gold colored metal beads was measured to have a mass 425 grams. The beads were placed in a graduated cylinder with an initial volume of 20.0 mL and the final volume of the water and beads was read to be 48.3 mL. What is the density of the beads.
      V=48.3 mL -20.0 mL =28.3 mL
      D = 425 g/ 28.3mL = 15.0 g/mL
   f. What is the mass of a metal object with a volume of 2.23 mL and a density of 9.43 g/mL?
      M= 9.43 g/mL x 2.23 mL = 21.0 g
   g. The density of an irregular metal is 29.3g/mL. If the metal nugget that weighs 75.3g and the initial volume was 20 ml, what would be the final volume of the graduated cylinder?
      V = 75.3g / (29.3 g/mL) = 2.57 mL Vf= 20.0 mL + 2.57mL =22.6 mL

10. Complete the following temperature conversions **K = °C + 273**
   a. 250 Kelvin to Celsius 250 - 273 = -23 °C
   b. 339 Kelvin to Celsius 339 - 273 = 66 °C
   c. 17 Celsius to Kelvin 17 + 273 = 290 K
   d. -20 Celsius to Kelvin -20 + 273 = 253 K

Unit 2 review, Chapter 2 (Matter and Change)

11. Complete the following table on the characteristics of the states of matter

	Solid	Liquid	Gas
Shape	Definite (set)	Definite (set)	Indefinite (not set)
Volume	Definite (set)	Indefinite (not set)	Indefinite (not set)

12. Define:

a. Matter: anything that has mass and takes ups space
b. Pure substance : matter that has a uniform and definite composition
c. Element: simplest form of matter, cannot be broken down
d. Compound: the combination of two or more elements, can only be separated chemically.
e. Mixture: a physical blend of two or more substance that are NOT chemical combined
f. Homogeneous mixture: a mixture that is uniform in composition (same throughout)
g. Heterogeneous mixture: a mixture that is NOT uniform in composition (different throughout)

13. **Both elements and compounds are examples of a pure substance, how are they different from each other?** Compound can be broken down chemically into smaller substances (elements)

14. **Classification of Matter**

Classify each as an element, compound, homogeneous mixture (homo) or a heterogeneous (hetero) mixture.

1. table salt compound	10. Salad dressing hetero	19. Iron element
2. gold element	11. Water compound (assume pure)	20. Helium element
3. the air in DHS homo	12. hydrogen chloride (HCl) compound	21. Wood hetero
4. carbon element	13. carbon element	22. blood homo
5. copper element	14. bucket of salt, sand, & water hetero	23. milk (store bought) homo
6. Kool-aid homo	15. water from water fountain homo	24. oily water hetero
7. fruit salad hetero	16. A root-beer float hetero	25. soil (dirt) hetero
8. city air hetero	17. Lucky charms cereal hetero	26. oxygen element
9. glucose compound	18. Flat soda homo	27. pure water compound

Classify each as a chemical or physical change	**Classify each as a chemical or physical property**
1. boiling water physical	1. Copper is a good conductor of heat and electricity physical
2. burning gasoline chemical	2. ice melts at 0°C. physical
3. cooking an egg chemical	3. a piece of sulfur is burned. chemical
4. ironing a shirt physical	4. 0_2 is a gas. physical
5. evaporating alcohol physical	5. Iron can rust chemical
6. rusting iron chemical	6. titanium is an inert metal. chemical
7. water evaporates. physical	7. He is very nonreactive. chemical
8. Ripping paper physical	8. Na is a soft, shiny metal. physical
9. Steel turns red when heated physical	9. ice melts at 0°C physical
10. fermenting orange juice chemical	10. water has a high specific heat. physical
11. rocks are ground to sand. physical	11. Alcohol burns in presence of a flame chemical
12. silverware tarnishes. chemical	12. gold is a yellow metal physical
13. digesting a pizza chemical	13. silver is a soft metal. physical
	14. gold is a very dense metal. physical

14. an ice melting in a drink physical 15. decomposing meat chemical 16. sulfur is burned. chemical 17. Carrots rot. chemical 18. Bread it cut into slices physical 19. Iron rust chemical	15. Hydrogen peroxide will break down into water and oxygen chemical 16. Sodium is highly reactive with water chemical 17. Water condenses at 100°C. physical 18. With electricity water with break down into oxygen and hydrogen chemical

15. What happens to the temperature of a substance during a phase change? The temperate remains constant during a phase change, all the heat added is going to the energy needed to change phases.

16. Round off the measurement 0.0030955 m to three significant figures. 0.00310 m (zero at end counts)

17. What is the product of the number 1000 and the measurement 0.00357 m expressed in the correct number of significant digits? 4 (only has 1 sig fig b/c of 1000 only having 1)

18. The mass of the electron is 9.1093910 kg. Express the mass of the electron to 1, 2, 3, and 4 significant figures. 4 sigfin = 9.109 kg 3 sigfig = 9.10 kg 2 sigfig= 9.1 kg 1 sigfig= 9 kg

19. . Define /describe the following separation techniques

a. Manual separation; physically picking out pieces/parts
b. Filtration: using a filter (or strainer) to allow large objects to remain behind and small object pass through.
c. Evaporation: used to separate a dissolved solid from a solution, heat up solution and the liquid evaporates leaving the solid behind.
d. Distillation : used to separate two or more liquids that have different boiling points, the steam MUST be collected.