Name ______________________________ Section __________ Date __________

Chapter 5: Basic Review Worksheet

1. Explain the scientific meaning of *uncertainty*?

2. Explain how a *unit* is related to a measurement?

3. Explain the terms *conversion factor* and *equivalence statement*.

4. For each of the following, make the indicated conversion.
   a. 122.4×10^5 to standard scientific notation
   b. 5.993×10^{-4} to ordinary decimal notation

5. For each of the following, make the indicated conversion.
   a. 6.0 pt to liters
   b. 6.0 pt to gallons
   c. 5.91 yd to meters
   d. 62.5 mi to kilometers
   e. 88.5 cm to millimeters

6. Evaluate each of the following mathematical expressions, being sure to express the answer to the correct number of significant figures.
   a. $10.20 + 4.1 + 26.0001 + 2.4$
   b. $(1.091 - 0.991) + 1.2$
   c. $(4.06 + 5.1)(2.032 - 1.02)$
   d. $(67.21)(1.003)(2.4)$

7. Make the indicated temperature conversions.
   a. 541 K to Celsius degrees
   b. 221 °C to kelvins

8. Given the following mass, volume, and density information, calculate the missing quantity.
   a. mass = 121.4 g; volume = 42.4 cm^3; density = ? g/cm^3
   b. mass = 0.721 lb; volume = 241 cm^3; density = ? g/cm^3

Name ____________________ Section ________ Date _________

Chapter 5: Review Worksheet

1. How does uncertainty enter into measurements? How is uncertainty indicated in scientific measurements?

2. Why must a unit be included with a measurement?

3. Give an everyday example of how you might use dimensional analysis to solve a simple problem.

4. For each of the following, make the indicated conversion.
   a. 0.0004321×10^4 to standard scientific notation
   b. 5.241×10^2 to ordinary decimal notation

5. For each of the following, make the indicated conversion.
   a. 16.0 L to fluid ounces
   b. 5.25 L to gallons
   c. 8.25 m to inches
   d. 4.25 kg to pounds
   e. 4.21 in. to centimeters

6. Evaluate each of the following mathematical expressions, being sure to express the answer to the correct number of significant figures.
   a. $[(7.815 + 2.01)(4.5)]/(1.9001)$
   b. $(1.67 \times 10^{-9})(1.1 \times 10^{-4})$
   c. $(4.02 \times 10^{-4})(2.91 \times 10^{3})/(9.102 \times 10^{-1})$
   d. $(1.04 \times 10^{2} + 2.1 \times 10^{1})/(4.51 \times 10^{3})$
   e. $(1.51 \times 10^{-3})^2/(1.074 \times 10^{-7})$
   f. $(1.89 \times 10^{2})/[(7.01 \times 10^{-3})(4.1433 \times 10^{4})]$

7. Make the indicated temperature conversions.
   a. – 50.1°C to Fahrenheit degrees
   b. – 30.7 °F to Celsius degrees

8. Given the following mass, volume, and density information, calculate the missing quantity.
   a. mass = ? g; volume = 124.1 mL; density = 0.821 g/mL
   b. mass = ? g; volume = 4.51 L; density = 1.15 g/cm^3

Name ____________________________ Section ________ Date __________

Chapter 5: Challenge Review Worksheet

1. Can uncertainty ever be completely eliminated in experiments? Explain.

2. Why is reporting the correct number of significant figures so important in science? Without consulting the text, summarize the rules for deciding whether or not a particular digit in a number used in a calculation is "significant." Summarize the rules for rounding off numbers. Summarize the rules for doing arithmetic with the correct number of significant figures.

3. For each of the following, make the indicated conversion.
   a. 0.0000009814 to standard scientific notation
   b. 14.2×10^0 to ordinary decimal notation

4. Make the indicated temperature conversions.
   a. 351 K to Fahrenheit degrees
   b. 72 °F to kelvins

5. At which temperature does the number of Celsius degrees equal the number of Fahrenheit degrees? Prove your answer.

6. Given the following mass, volume, and density information, calculate the missing quantity.
   a. mass = 142.4 g; volume = ? mL; density = 0.915 g/mL
   b. mass = 4.2 lb; volume = ? cm^3; density = 3.75 g/cm^3

7. You measure the mass of an object and find it to be 128.1 ± 0.1 g. You measure the volume of this object and report it as 24.3 ± 0.01 mL. Calculate the density and report it as ___ ± ___ g/mL.

2. A polyatomic is an ion containing more than one atom. Parentheses are used in writing formulas containing polyatomic ions to indicate unambiguously how many of the polyatomic ions are present in the formula, while making certain that there is no mistake as to what is meant by the formula. For example, consider the calcium phosphate. The correct formula for this substance is $Ca_3(PO_4)_2$, which indicates that three calcium ions are combined for every two phosphate ions (check the total number of positive and negative charges to see why this is so). If we did not write the parenthesis around the formula for the phosphate ion, that is, if we had written Ca_3PO_{42}, people reading this formula might think that there were 42 oxygen atoms present.

3. a. chromium(III) sulfide; b. copper(I) sulfide; c. iron(III) oxide; d. gold(III) iodide; e. manganese(IV) oxide; f. cobalt(II) bromide

4. a. correct; b. incorrect, should be CsCl; c. incorrect, should be Ba_3P_2; d. correct

5. a. diboron trioxide; b. tetraphosphorus decoxide; c. dinitrogen trioxide

6. a. H_2S; b. $Ba_3(PO_4)_2$; c. $Mg(ClO_4)_2$; d. $MnCl_2$

Chapter 5: Basic Review Worksheet

1. Uncertainty in measurement means that we can never take an exact measurement (except for counting). The last digit of a recorded measurement is estimated, and therefore uncertain.

2. A unit tells us what scale or standard is being used to represent the results of the measurement.

3. Dimensional analysis is a method of problem solving which pays particular attention to the units of measurements and uses these units as if they were algebraic symbols that multiply, divide, and cancel. Consider the following example. A dozen of eggs costs $1.25. Suppose we want to know how much one egg costs, and also how much three dozen eggs will cost. To solve these problems, we need to make use of two equivalence statements:

   1 dozen eggs = 12 eggs

   1 dozen eggs = $1.25

   The first of these equivalence statements is obvious: everyone knows that 12 eggs are "equivalent" to one dozen. The second statement also expresses an equivalence: if you give the grocer $1.25, he or she will give you a dozen eggs. From these equivalence statements, we can construct the conversion factors we need to answer the two questions. For the first question (what does one egg cost) we can set up the calculation as follows

   $$\frac{\$1.25}{12\text{ eggs}} = \$0.104 = \$0.10$$

   as the cost of one egg. Similarly, for the second question (the cost of 3 dozen eggs), we can set up the conversion as follows

$$3 \text{ dozen} \times \frac{\$1.25}{1 \text{ dozen}} = \$3.75$$

as the cost of three dozen eggs. See section 5.6 of the text for how we construct conversion factors from equivalence statements.

4. a. $122.4 \times 10^5 = (1.224 \times 10^2) \times 10^5 = 1.224 \times 10^7$
   b. $5.993 \times 10^{-4} = 0.0005993$

5. a. $6.0 \text{ pt} \times \frac{1 \text{ qt}}{2 \text{ pt}} \times \frac{1 \text{ L}}{1.0567 \text{ qt}} = 2.8 \text{ L}$
   b. $6.0 \text{ pt} \times \frac{1 \text{ qt}}{2 \text{ pt}} \times \frac{1 \text{ gal}}{4 \text{ qt}} = 0.75 \text{ gal}$
   c. $5.91 \text{ yd} \times \frac{1 \text{ m}}{1.0936 \text{ yd}} = 5.40 \text{ m}$
   d. $62.5 \text{ mi} \times \frac{1 \text{ km}}{0.62137 \text{ mi}} = 101 \text{ km}$
   e. $88.5 \text{ cm} \times \frac{10 \text{ mm}}{1 \text{ cm}} = 885 \text{ mm}$

6. a. $10.20 + 4.1 + 26.001 + 2.4 = 42.701 = 42.7$ (one decimal place)
   b. $[1.091 - 0.991] + 1.2 = 1.3$ (one decimal place)
   c. $(4.06 + 5.1)(2.032 - 1.02) = (9.16)(1.012) = (9.2)(1.01) = 9.3$
   d. $(67.21)(1.003)(2.4) = 161.8 = 1.6 \times 10^2$ (only 2 significant figures)

7. a. $541 \text{ K} - 273 = 268°\text{C}$
   b. $221°\text{C} + 273 = 494 \text{ K}$

8. a. density $= 212.4 \text{ g}/42.4 \text{ cm}^3 = 2.86 \text{ g/cm}^3$
   b. $0.721 \text{ lb} = 327 \text{ g}$
   density $= 327 \text{ g}/241 \text{ cm}^3 = 1.36 \text{ g/cm}^3$

Chapter 5: Review Worksheet

1. Whenever a scientific measurement is made, we always employ the instrument or measuring device we are using to the limits of its precision. On a practical basis, this usually means that we *estimate* our reading of the last significant figure of the measurement. Scientists appreciate the limits of experimental techniques and instruments, and always assume that the last digit in a number representing a measurement has been estimated.

2. We need to include a unit in order to be able to communicate. If we report, for example, that we have "1.4 of water" this is not very helpful. Is this amount 1.4 cups? 1.4 liters? The unit is just as important as the number.

3. Answers will vary.

4. a. $0.0004321 \times 10^4 = (4.321 \times 10^{-4}) \times 10^4 = 4.321 \times 10^0 = 4.321$
   b. $5.241 \times 10^2 = 524.1$

5. a. $16.0\,\text{L} \times \dfrac{1\,\text{qt}}{0.94633\,\text{L}} \times \dfrac{32\,\text{fl. oz.}}{1\,\text{qt}} = 541\,\text{fl. oz.}$
   b. $5.25\,\text{L} \times \dfrac{1\,\text{gal}}{3.7854\,\text{L}} = 1.39\,\text{gal}$
   c. $8.25\,\text{m} \times \dfrac{1.0936\,\text{yd}}{1\,\text{m}} \times \dfrac{36\,\text{in.}}{1\,\text{yd}} = 325\,\text{in}$
   d. $4.25\,\text{kg} \times \dfrac{2.2046\,\text{lb}}{1\,\text{kg}} = 9.37\,\text{lb}$
   e. $4.21\,\text{in} \times \dfrac{2.54\,\text{cm}}{1\,\text{in}} = 10.7\,\text{cm}$

6. a. $[(7.815 + 2.01)(4.5)]/(1.9001) = 23$
   b. $(1.67 \times 10^{-9})(1.1 \times 10^{-4}) = 1.837 \times 10^{-13} = 1.8 \times 10^{-13}$
   c. $(4.02 \times 10^{-4})(2.91 \times 10^{3}) / (9.102 \times 10^{-1}) = 1.29$
   d. $(1.04 \times 10^2 + 2.1 \times 10^1) / (4.51 \times 10^3) = 2.77 \times 10^{-2}$
   e. $(1.51 \times 10^{-3})^2 / (1.074 \times 10^{-7}) = 21.2$
   f. $(1.89 \times 10^2) / [(7.01 \times 10^{-3})(4.1433 \times 10^4)] = 0.651$

7. a. $1.80(-50.1°\text{C}) + 32 = -58.2°\text{F}$
   b. $[(-30.7°\text{F}) - 32] / 1.80 = -34.8°\text{C}$

8. a. mass $= 124.1\,\text{mL} \times 0.821\,\text{g.mL} = 102\,\text{g}$
   b. $4.51\,\text{L} = 4{,}510\,\text{cm}^3$
   mass $= 4{,}510\,\text{cm}^3 \times 1.15\,\text{g/cm}^3 = 5.19 \times 10^3\,\text{g}$

Chapter 5: Challenge Review Worksheet

1. Since the last significant figure in every measurement is assumed to be estimated, it is never possible to exclude uncertainty from measurements. The best we can do is to try to improve our techniques and instruments so that we get more significant figures for our measurements.

2. Scientists are careful about reporting their measurements to the appropriate number of significant figures to indicate to their colleagues the precision with which the experiments were performed. That is, the number of significant figures reported indicates how "carefully" measurements were made. Suppose you were considering buying a new home, and the real estate agent told you that a prospective new house was "between 1000-2000 square feet of space: and had "five or six rooms, more or less" and stood on "maybe an acre or two of land". Would you buy the house or would you look for another real estate agent? When a scientist says that a sample of material "has a mass of 3.126 grams" he or she is narrowing down the limits as to the actual, true mass of the sample: the mass is clearly slightly more than half way between 3.12 and 3.13 grams.

   The rules for significant figures are covered in the text. In brief, these rules for experimentally measured numbers are as follows: (1) nonzero integers are *always* significant; (2) leading zeroes are *never* significant, captive zeroes are *always* significant, and trailing zeroes *may* be significant (if a decimal point is indicated); (3) exact numbers (e.g., definitions) have an infinite number of significant figures.

When we have to round off an answer to the correct number of significant figures (as limited my whatever measurement was least precise), we do this in a particular manner. If the digit to be removed is equal to or greater than 5, the preceding digit is increased by 1. If the digit to be removed is less than 5, the preceding digit is not changed. To perform a series of calculations involving a set of data, retain all of the digits in the intermediate calculations until arriving at the final answer, and then round off to the appropriate number of significant figures.

When doing arithmetic with experimentally determined numbers, the final answer is limited by the least precise measurement. In doing multiplication or division calculations, the number of significant figures in the result should be the same as the measurement with the fewest significant figures. In performing addition or subtraction, the number of significant figures in the result is limited by the measurement with the fewest decimal places.

3. a. $0.0000009814 = 9.814 \times 10^{-7}$
   b. $14.2 \times 10^{0} = 14.2$

4. a. $351\text{ K} - 273 = 78°\text{C}$
   $1.80\,(78°\text{C}) + 32 = 172.4 = 172°\text{F}$
   b. $[(72°\text{F}) - 32] / 1.80 = 22.2°\text{C}$
   $22.2°\text{C} + 273 = 295.2 = 295\text{ K}$

5. $-40°\text{C} = -40°\text{F}$; $1.8C + 32 = F$; if $C = F$, $1.8C + 32 = C$, thus $32 = -0.8C$ and $-40 = C$

6. a. volume = 142.4 g/0.915 g/mL = 156 mL
   b. $4.2\text{ lb} = 1.9 \times 10^{3}\text{ g}$
   volume = $1.9 \times 10^{3}\text{ g}/3.75\text{ g/cm}^3 = 507\text{ cm}^3 = 5.1 \times 10^{2}\text{ cm}^3$

7. 5.27 ± 0.03 g/mL

   128.1 g/24.3 mL = 5.27 g/mL. We can get the extremes of the densities by dividing the maximum mass by the minimum volume, and the minimum mass by the maximum volume. Doing so we get: 128.2 g/24.2 mL = 5.30 g/mL and 128.0 g/24.4 mL = 5.25 g/mL.

Chapter 6: Basic Review Worksheet

1. The average atomic mass of an element represents the weighted average mass, on the relative atomic scale, of all the isotopes of an element. Average atomic masses are usually given in terms of atomic mass units

2. The molar mass of a compound is the mass in grams of one mole of the compound (6.022×10^{23} molecules of the compound), and is calculated by summing the average atomic masses of all the atoms present in a molecule (or empirical formula unit for an ionic substance) of the compound. For example, a unit of the compound K_2O contains two potassiums and one oxygen: the molar mass is obtained by adding up the average atomic masses of these atoms: molar mass K_2O = 2(39.10 g) + 1(16.00 g) = 94.20g