Name ________________________ **AP CHEM** ____/____/____

Chapter 1 Chemical Foundations

Metric Conversions

All measurements in chemistry are made using the metric system. In using the metric system you must be able to convert between one value and another. You must memorize the factors, prefixes and symbols in the chart below.

There are numerous ways to do metric conversion. I present one method below that I have used in my Honors and Chemistry I classes.

Example: Convert 1.83×10^{-1} kilograms to centigrams.

Step 1. Subtract the power of ten value you are solving for from the power of ten value you are given.
In this problem you are given 1.83 kilograms and you are converting to centigrams. Kilograms have a power of ten value of 3; centigrams have a power of ten value of -2. 3-(-2) = 5

Step 2. Write your result as the power of ten value of your answer.
In this problem our answer is 1.83×10^{-1} $\mathbf{\times 10^5}$.

Step 3. Put your answer in proper scientific notation.
$1.83 \times 10^{-1} \times 10^5$ is converted to $\mathbf{1.83 \times 10^4}$. Remember that when multiplying using powers of ten you add.

Factor	Prefix	Symbol
1×10^{12}	tera-	T
1×10^{9}	giga-	G
1×10^{6}	mega-	M
1×10^{3}	kilo-	k
1×10^{2}	hecto-	h
1×10^{1}	deca-	D
1×10^{0}	---	---
1×10^{-1}	deci-	d
1×10^{-2}	centi-	c
1×10^{-3}	milli-	m
1×10^{-6}	micro-	µ
1×10^{-9}	nano-	n
1×10^{-10}	angstrom	Å
1×10^{-12}	pico-	p

Uncertainty in Measurement

- Depends on the precision of the measuring device
  - For example a measurement of 1.682956 grams is a more precise measurement than 1.7 grams
- Reliability in Measurements
  - Accuracy – the closeness to the actual scientific value
  - Precision – getting repeated measurements in repeated trials
- Types of Errors
  - Random: error in measurement has equal probability of being high or low
  - Systematic: errors all occur in the same direction

General Rules for determining if a Number is Significant

1. Draw a box around all nonzero digits beginning with the leftmost nonzero digit and ending with the rightmost nonzero digit in the number.
2. If a decimal is present, draw a box around any trailing zeros **to the right of the original box**.
3. Consider any all boxed digits significant.

Example 1: 20406: 5 significant digits
Example 2: 0.0045: 2 significant digits
Example 3: 4000: 1 significant digit
Example 4: 4000. : 4 significant digits
Example 5: 0.002500: 4 significant digits
Example 6: 3.00: 3 significant digits

Addition or Subtraction using Significant Figures

The answer can only be as precise as the least precise measurement.

Example
2.8701 (precise **four** places to the right of the decimal)
0.0673 (precise **four** places to the right of the decimal)
<u>+ 301.520</u> (precise **three** places to the right of the decimal)
304.4574 → rounds off to 304.457 (answer must be precise **three** places after the decimal)

Multiplication or Division using Significant Figures

The answer can have no more total significant figures than there are in the measurement with the smallest total number of significant figures.

Example:
$$
\begin{array}{rl}
12.257 & \text{(5 total significant figures)} \\
\times\ \underline{1.162} & \text{(4 total significant figures)} \\
14.2426 \rightarrow & \text{rounds off to } 14.24 \quad \text{(4 total significant figures)}
\end{array}
$$

As a general rule, if you are unsure how many significant figures to us on the AP exam, use 3 significant figures. This may not always work but it will work most times. However you should always pay close attention to using the correct number of significant figures in all calculations.

Scientific Notation

In chemistry, we often use numbers that are either very large (1 mole = 602 200 000 000 000 000 000 000 atoms) or very small (the mass of an electron = 0.000 000 000 000 000 000 000 000 000 000 910 939 kg). Writing numbers with so many digits would be tedious and difficult. To make writing very large and small numbers easier, scientists use an abbreviation method known as scientific notation. In scientific notation the numbers mentioned above would be written as 6.022×10^{23} and 9.10939×10^{-31}.

Converting a number to or from scientific notation

- If you move the decimal place to the left, the power of 10 value increases.
- If you move the decimal place to the right, the power of 10 value decreases.

Example 1: Look at the first number from above: 602 200 000 000 000 000 000 000
To put this number in scientific notation you would move your decimal place until there is one number to the left of the decimal. To do this, we must move our decimal 23 places to the left. When you move the decimal to the left, the power of 10 value increases. It increases from 0 to 23. Thus, the answer is $\mathbf{6.022 \times 10^{23}}$

Look at the second number from above: 0.000 000 000 000 000 000 000 000 000 000 910 939
To put this number in scientific notation we must move our decimal 31 places to the right. **REMEMBER: You should always have one digit to the left of the decimal when writing numbers in scientific notation.** Since we are moving our decimal to the right, we must decrease our power of 10 value. It decreases from 0 to –31. The answer is $\mathbf{9.10939 \times 10^{-31}}$

Rules for multiplying & dividing using scientific notation:

- **When multiplying two numbers in scientific notation, ADD their power of 10 values**.

  For example: $(3.45 \times 10^6)(4.3 \times 10^5) = 14.835 \times 10^{11}$. But, we must also remember to express our answer in significant figures. Thus, the final answer is $\mathbf{1.5 \times 10^{12}}$
- **When dividing numbers in scientific notation, SUBTRACT the denominator's power of 10 value from the numerator's power of 10 value**.

  For example: $(2.898 \times 10^{12}) \div (3.45 \times 10^{15}) = 0.840 \times 10^{-3}$ (I had to add the zero at the end to get the three significant figures needed.) I got 10^{-3} because $12 - 15 = -3$. Make sure your answer is in proper scientific notation (one number to the left of the decimal). In this problem we have to move the decimal one place to the right. When we move our decimal to the right, we decrease our power of 10. –3 decreases by 1 to –4. Our final answer is: $\mathbf{8.40 \times 10^{-4}}$

Dimensional Analysis

- Used to convert a number from one system of units to another.
- **Understanding dimensional analysis is crucial. This process will help you when performing difficult calculations later in the year.**
- You will find a complete listing of English/Metric Equivalents on the last page of your text book. Conversion factors are also available from the side "tools" menu at: http://www.sartep.com/chem. Conversion factors **do not** need to be memorized.

Example: Calculate how many kilometers there are in 5 miles.
Solution: (Needed Equivalents: 1 mile = 1760 yards, 1 meter = 1.094 yards)

$$\frac{5 \text{ miles}}{} \times \frac{1760 \text{ yards}}{1 \text{ mile}} \times \frac{1 \text{ meter}}{1.094 \text{ yards}} \times \frac{1 \text{ kilometer}}{1000 \text{ meters}} = 8.04 \text{ (rounded to) } 8 \text{ kilometers}$$

Example: Convert 55.0 miles/hour to meters/second
Solution: (Needed Equivalents: 1 mile = 1760 yards, 1 meter = 1.094 yards, 1 hour = 60.0 minutes, 1 minute = 60.0 seconds)

$$\frac{55.0 \text{ miles}}{\text{hour}} \times \frac{1760 \text{ yards}}{1 \text{ mile}} \times \frac{1 \text{ meter}}{1.094 \text{ yards}} \times \frac{1 \text{ hour}}{60.0 \text{ minutes}} \times \frac{1 \text{ minute}}{60.0 \text{ seconds}} = 24.5785 \text{ (rounded to) } 24.6 \text{ m/s}$$

General Rules for Rounding Numbers in Chemistry

- **Rule 1. If the digit following the last significant figure is less than 5, the last significant figure remains unchanged. The digits after the last significant figure are dropped.**
  Example: Round 23.437 to three significant figures.
  Answer: 23.4
  Explanation: 4 is the last significant figure. The next number is 3. 3 is less than 5. Thus, 4 remains unchanged and 37 is dropped.

- **Rule 2. If the digit following the last significant figure is 5 or greater, then 1 is added to the last significant figure. The digits after the last significant figure are dropped.**
  Example: Round 5.383 to two significant figures.
  Answer: 5.4
  Explanation: 3 is the last significant figure. The next number is 8. 8 is greater than 5. Thus, 1 is added to 3 making it 4. The 83 is dropped.

- As a rule, when performing a series of calculations, **wait until the very end to round off to the proper number of significant figures instead of rounding off each intermediate result unless you are changing from addition /subtraction to multiplication/division or vice versa.**

  Example 1: 10.82 + 2.5 + 2.64 =
  WRONG: 10.82 + 2.5 = 13.32 (rounded to 13.3) 13.3 + 2.64 = 15.94 (rounded to) 15.9
  CORRECT: 10.82 + 2.5 + 2.64 = 15.96 (rounded to) 16.0 (precise to 1 place after the decimal)

  Example 2: (12.00 – 10.00) ÷ 12.00 =
  In this case you would subtract and ROUND to the proper number of significant figures and then divide because you are changing between significant figure rules.
  CORRECT: (12.00-10.00) = 2.00 (precise to 2 places after the decimal)
  2.00 ÷ 12.00 = 0.167 (rounded to 3 total significant figures)

Temperature Conversion

C: Celsius or Centigrade
K: Kelvin (named for Lord Kelvin, a.k.a. William Thomson)
Fahrenheit is not used in AP Chemistry so we will ignore it.

K = °C + 273 **°C = K – 273** **(These formulas must be memorized)**

Example 1: Convert 29.0 °C to Kelvin.
K = 29.0 + 273
K = 302

Example 2: Convert 888 K to °C.
°C = 888 – 273
°C = 615

Density

Density (d) is the ratio of the mass (m) of a substance to the volume (v) occupied by the substance. Pure water is used as the standard in measuring density. **The density of pure water is 1.0 g/mL**. In more precise calculations the actual density of water will differ slightly. If a substance has a density less than water, it will float; if a substance has a density greater than water, it will sink.

$$\boxed{d = \frac{m}{v}}$$

Mass is expressed in grams (g). Volume is expressed in liters (L), milliliters (mL) or cubic centimeters (cm^3). A mL is the same as a cm^3. Thus, density of a liquid or solid can be expressed as g/mL or g/cm^3. The density of a gas is expressed in g/L.

Example: A piece of wood has a volume of 3350 cm^3. If the density of the wood is 0.512 g/mL, what is its mass?

d = 0.512 g/mL
v = 3350 cm^3
m = x

$0.512 = \frac{X}{3350}$ X = 1715.2 **1.72×10^3 grams (3 significant digits)**

Percent Error

The accuracy of your measurements can be checked by calculating the percent error. In a percent error calculation you will compare your experimental value to the accepted scientific value (referred to as the theoretical value). Since you are taking the absolute value of the subtraction**, your percent error will always be a positive number**. Remember to use significant figures in all % error calculations. In most quantitative lab experiments you will be expected to calculate your % error.

$$\boxed{\textbf{\%Error} = \frac{|\ \textbf{theoretical yield} - \textbf{experimental yield}\ |}{\textbf{theoretical yield}} \ \textbf{x 100.}}$$

Example: The theoretical density of aluminum is 2.70 g/mL. In an experiment a student measures the mass of an aluminum bar to be 12.52 grams and finds the volume of the bar to be 4.71 mL. Calculate the student's percent error.

First find the experimental density: $d = \frac{12.52}{4.71}$ d = 2.66 g/mL

Now use the experimental value to find % error. $\%E = \frac{|2.70 - 2.66|}{2.70} \times 100. \rightarrow \frac{0.04}{2.70} \times 100. \rightarrow$ **%E = 1%**

Classification of Matter

Chemistry is the study of matter

- Matter – anything that has mass and takes up space
    - Matter exists in states
        - Solid – rigid, fixed shape and volume
        - Liquid – definite volume, takes the shape of the container
        - Gas – no fixed volume or shape, very compressible
    - Most matter exists as mixtures
        - Example: wood, wine
        - Homogeneous mixtures
            - Visibly indistinguishable parts
            - Called solutions
            - Example: air, brass, salt water
        - Heterogeneous mixtures
            - Visibly distinguishable parts
            - Example: sand in water, oil & water
    - Can be separated into pure substances through physical changes
        - Physical changes do not change the chemical composition of the matter
        - Ways to separate mixtures
            - Distillation – separation by boiling point
            - Filtration – separation method used with a solid & liquid mixture where a barrier blocks solid particles from passing through
            - Chromatography – a series of methods that employ a system with two phases(states of matter)
                - Mobile phase (liquid or gas)
                - Stationary phase (solid)
                - Types of Chromatography
                    - Paper Chromatography
                    - Gas Chromatography
                    - HPLC Chromatography
    - Pure substances
        - Matter made up of only one type of element or compound
            - Compound – substance with constant composition that can be broken down into elements by chemical means
            - Elements – a substance that cannot be decomposed into simpler substances by chemical or physical means
                - Atoms – the most basic unit of matter
                    - Proton – positively charged particle
                    - Neutron – neutral particle
                    - Electron – negatively charged particle