We begin with a quick review of the math skills you will need for this course.

MATH SKILLS NEEDED FOR CHEMISTRY

SIGNIFICANT FIGURES

In Science, measured values are reported in terms of significant figures. **Significant figures** in a measurement consist of all the digits known with certainty plus one final digit, which is somewhat uncertain or is estimated (when measuring in the lab). For example, if a nail length is 6.36 cm, the last digit 6 is uncertain. All the digits, including the uncertain one, are significant. All contain information and are included in the reported value. Therefore, the term significant does not mean certain. In any correctly reported measured value, the final digit is significant but not certain. Insignificant digits are never reported. As a Chemistry student, you will need to use and recognize significant figures when you work with measured quantities and report your results, and when you evaluate measurements reported by others.

RULES

1. All nonzero digits **are** significant.
2. All zeros between two nonzero digits **are significant.**
3. Zeros to the right of a nonzero digit, but to the left of an understood decimal point, **are not** significant **unless** specifically indicated as significant by a bar placed above the rightmost such zero that is significant.
4. All zeros to the right of a decimal point but to the left of a nonzero digit **are not** significant.
5. All zeros to the right of a decimal point and to the right of a nonzero digit **are** significant.

Practice Problems:

In each of the following measurements, (a) determine the number of significant figures, and (b) identify the applicable rules.

1. 1.030 cm
2. 2,074,000.0 s
3. 0.00080 kg
4. 147 g
5. 1,570,500cm/s

ROUNDING OFF NUMBERS

Rounding depends on the identity of the next digit (i.e., the digit after the **cutoff** point).

RULES

1. If the next digit is less than 5, then the previous digit remains the same.
2. If the next digit is greater than 5 or 5 followed by non-zeros, then the previous digit is increased by one.
3. If the next digit is 5 or 5 followed by all zeros, then the previous digit remains the same, if it is even or increased by one if it is odd (i.e., keep or make the previous digit even).

Practice Problems:

Round off the following numbers to 2 and 3 significant figures respectively.

1. 1.294
2. 0.9946
3. 0.999
4. 1.2951
5. 1.325001
6. 1.285
7. 1.295
8. 1.22

9. 500

SCIENTIFIC NOTATION

Scientific notation is a way to easily handle very large numbers or very small numbers.
The following websites and videos will provide a review on how to convert from scientific notation to standard form as well as how to perform basic mathematical operations using these numbers
http://www.chem.tamu.edu/class/fyp/mathrev/mr-scnot.html
http://www.purplemath.com/modules/exponent3.htm
http://www.youtube.com/watch?v=H578qUeoBC0

After reviewing, complete the following questions.

Practice Problems:

Converting from scientific notation and standard form.

1) $6.781 \times 10^{6}=$

2) $1.177 \times 10^{3}=$

3) $3.4377 \times 10^{-7}=$

4) $2.231 \times 10^{-7}=$

Converting from standard form to scientific notation.

5) 0.000096949

6) 1220000000000000

7) 0.000002834

8) 12.221

9) 78.21

10) .00000000000000000000005

Multiplying and dividing with scientific notation (your answer must also be in scientific notation!)

11) (9.18×10^{-6}) x (1.15×10^{-5})

12) (1.89×10^{4}) x (5.14×10^{7})

13) $\dfrac{(1.22 \times 10^{-8})}{(8.15 \times 10^{9})}$

14) $\dfrac{(6.2 \times 10^{10})}{(9.1 \times 10^{14})}$

CALCULATING PERCENTS

Calculating the Percent of Experimental Error

The Percent of Experimental Error (PEE, or PE) is the difference (absolute value) between your experimental value (the value you get in a lab) and the actual, standard, value. The difference is then turned into a percent by dividing it by the actual/standard value, and multiplying by 100 (round to 0.1%)

$$PEE = \left| \frac{\text{Actual value} - \text{Experimental Value}}{\text{Actual Value}} \right| \times 100$$

Calculating Percent Change

Percent Change is the difference (absolute value) between an original number and a new number, divided by the original number and turned into a percent (x 100), rounded to the 0.1 %.

Practice Problems:

Percent Error

1. The weight of a rock is 16. 8 g, but in the lab it weighed 17.2 g. Find the Percent Error

Percent Change

2. Due to evaporation a crystal grows from 2.6 g to 3.6 g. Find the Percent Change

ALGEBRAIC MANIPULATION OF VARIABLES

Using your previous algebra 1 knowledge, isolate the indicated variable in each one of the following formulas (remember to use the opposite operation to get rid of a variable on any side):

1. $D = m/v$; solve for m
2. $C = wf$; solve for f
3. $E = hf$; solve for h
4. $PV = nRT$; solve for R
5. $V_1/T_1 = V_2/T_2$; solve for T_2

Practice Problems:
Use the formulas from above to solve using the given values:

1. Find the value for v in ml; if D = 2.4 g/ml and m = 3.5 g

2. Find the value for the wavelength (w) if $C = 3X10^8$ m/s and v = 45 sec^{-1}

3. Find the value for frequency (f) if h = $6.26 \text{ X } 10^{-34}$ Js and E = 450 J

4. Find the value for R if P = 200 atm, v = 28 L, n = 4 moles and T = 300K

5. Find the value for T_1 given: T_2 = 308K; V_1 = 25ml and V_2 = 67ml

DIMENSIONAL ANALYSIS

Dimensional Analysis (also called Factor-Label Method or the Unit Factor Method) is a problem-solving method that uses the fact that any number or expression can be multiplied by one without changing its value. It is a useful technique. The only danger is that you may end up thinking that chemistry is simply a math problem - which it is not.

Unit factors may be made from any two terms that describe the same or equivalent "amounts" of what we are interested in. For example, we know that

1 inch = 2.54 centimeters

We can make two-unit factors from this information:

Now, we can solve some problems. Set up each problem by writing down what you need to find with a question mark. Then set it equal to the information that you are given. The problem is solved by multiplying the given data and its units by the appropriate unit factors so that only the desired units are present at the end.

(1) How many centimeters are 6.00 inches?

(2) Express 24.0 cm in inches.

You can also string many unit factors together.

(3) How many seconds are in 2.0 years?

$$? \text{ seconds} = \frac{2.0 \text{ years}}{1} \times \frac{365 \text{ days}}{1 \text{ year}} \times \frac{24 \text{ hours}}{1 \text{ day}} \times \frac{60 \text{ minutes}}{1 \text{ hour}} \times \frac{60 \text{ seconds}}{1 \text{ minute}}$$

$$= 63{,}000{,}000 \text{ seconds}$$

Always put the unit that you are starting with on top of your fraction and then cancel it out by using your conversion factors. To cancel out the original unit, it must be on the bottom of the conversion factor that you are using. If you look at the problem below it is the same as the one above but highlighted to show how the units cancel each other out. The problem above is highlighted to show how this is done.

Scientists generally work in metric units. Common prefixes used are the following:

Prefix	**Abbreviation**	**Meaning**	**Example**
mega-	M	10^6	1 megameter (Mm) = 1 x 10^6 m
kilo-	k	10^3	1 kilogram (kg) = 1 x 10^3 g
centi-	c	10^{-2}	1 centimeter (cm) = 1 x 10^{-2} m
milli-	m	10^{-3}	1 milligram (mg) = 1 x 10^{-3} g
micro-		10^{-6}	1 micrometer (g) = 1 x 10^{-6} g
nano-	n	10^{-9}	1 nanogram (ng) = 1 x 10^{-9} g

(4) Convert 50.0 mL to liters. (This is a very common conversion.)

(5) What is the density of mercury (13.6 g/cm^3) in units of kg/m^3?

Practice Problems:

1. How many seconds are there in 1 day?
2. If you have 8 dozen apples, how many total apples do you have?
3. Express 1,250 centimeters in kilometers.
4. How many centimeters are in 10 feet?
5. If a car is traveling 60 miles per hour, how many meters per second is going?

1.1 Moles and Molar Mass

Molar Mass, and Molecular Weight

- Molar mass (mm) the mass of a mole
- Molecular weight (mw) sum of all atomic weights
- **Pay attention to significant figures!**

The Mole

- **Avogadro's number** 6.02×10^{23}**, the number of particles in a mole of anything**

EXAMPLE PROBLEMS BELOW

Example

Aluminum is a metal with a high strength to mass ratio and a high resistance to corrosion, thus it is often used for structural purposes. Compute both the number of moles of atoms and the high number of atoms in a 10.0 g sample of aluminum.

Example

Calcium carbonate is the principal mineral found in limestone, marble and the shells of marine animals such as clams.

A. Calculate the molar mass of calcium carbonate. Show work.

B. A certain sample of calcium carbonate contains 4.86 moles. What is the mass of grams in this sample? What is the mass of the CO_3^{2-} ions present? Show work.

Example

A sucrose packet contains 4.0 g of sucrose, how many moles of sucrose does the packet contain?

1. Find the molar mass of sucrose

2. Write a mole-mass conversion factor with grams in the denominator
3. Multiply given quantity by the conversion factor

1.2 Mass Spectroscopy of Elements

ISOTOPES

- Atoms having the same atomic number but a different number of neutrons
- Most elements have at least two stable isotopes
- 0 neutrons hydrogen
- 1 neutron deuterium
- 2 neutrons tritium

Atomic Masses

- Atomic mass is the average of all natural isotopes in the element

Mass Spectrometer

- Determines atomic masses of an atom
- Determine isotope composition of an element

Average Atomic mass

- (Mass 1) (%) + (Mass 2) (%)

IONIZATION

- Atoms or molecules are passed into a beam of high speed electrons. This knocks electrons off the atoms or molecules transforming them into cations.

Example

A certain isotope X contains 23 protons and 28 neutrons. What is the mass number of this isotope? Identify the element. Justify your reasoning.

Example

An element consists of 62.20 % of an isotope with mass 186.956 amu and 37.40 % of an isotope with mass 184.953 amu. Calculate the average atomic mass and identify the element. Justify.

1.3 Elemental Composition of Pure Substances

The Law of Definite Proportions : a given compound always contains exactly the same proportions of elements by mass.

Example

Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

Example

What is the percent composition by mass of N_2O_4?

Empirical formula steps

1. Base calculation on 100 grams
2. Determine moles of each element
3. Divide by smallest value
4. Multiply by an integer to obtain whole numbers

Example

11.0 g of a certain compound contains 2.82 g of magnesium and 8.18 g of chlorine. What is the empirical formula?

Example

A 0.200 gram sample of a compound (vitamin C) composed of only C, H, and O is burned completely with excess O2 0.2998 of CO_2 and 0.0819 g of H_2O are produced. What is the empirical formula?

Molecular formula

- The actual molecule

- Need the molecular weight
- Multiply ratio by the empirical formula

Example

Determine the empirical and molecular formulas for a compound that gives the following analysis (in mass percents)

71.65% Cl   24.27% C   4.07% H   The molar mass is known to be 98.96 g/mol.

Example

A white powder is analyzed and found to contain 43.64% phosphorus and 56.36% oxygen by mass. The compound has a molar mass of 283.88 g/mol. What are the compound's empirical and molecular formulas?

1.4 Composition of Mixtures

Hydrates

- an ionic compound that contains water molecules in its structures

Anhydrate

- Is the substance that remains after the water is removed from a hydrate

Example

A calcium chloride hydrate has a mass of 4.72 g. After heating for several minutes the mass of the anhydrate is found to be 3.56 g. Use this information to determine the formula for the hydrate.

Example

A 2.4 g mixture of calcium chloride, and sodium chloride is found to contain 0.12 g Na. What percent of this sample is NaCl?

Example

A 5.0 g sample of a mixture $CaCO_3$ and SiO_2 contains 1.5 g of Ca. What percent of $CaCO_3$ is in the mixture?

1.5 Atomic Structure and Electron Configuration

Heisenberg uncertainty principle - there is a fundamental limitation to just how precisely we can know both the position and momentum of a particle.

Orbitals - probability function describing the possibility that an electron will be found in a region of space.

Principle quantum number

- Has integral values : 1,2,3 ……
- Higher numbers = greater distance
- Greater distance = less tightly bound = higher energy

Orbital Shapes and Energies :

- Size of orbital
- Defined as the surface that contains 90% of the total electron probability
- Orbitals of the same shape grow larger as it increases

1. Aufbau Principle - fill orbitals with the lowest energy first
2. Hund's Rule - always fill each orbital before pairing electrons to minimize repulsions
3. Pauli Exclusion Principle - in a given atom no two electrons can have the same set of four quantum numbers

Example

Write the electron configurations for the following atoms

A. Cs
B. Ni^{2+}
C. Se
D. Ag
E. Cl^{1-}
F. Ca^{2+}

Example

Write the expected last terms of the electron configurations for each of the following :

A. Cl
B. Sb
C. Sr
D. W
E. Pb
F. Mn

1.6 Photoelectron Spectroscopy

Quantum effect - must add enough energy to completely remove the electron

Photoelectron Spectroscopy

- Atoms bombarded with photons
- Relative amount of energy it takes to remove different types of electrons
- This creates peaks that show up on the PES diagram
- Location of peak = how much energy it takes to remove electron
- Size of peak = number of electrons based on relative size

1.7 Periodic Trends

Coulomb's Law

- The subatomic particles in an atom are charged and interact with one another so that they each experience an electrostatic force.
- Same charge = repulse
- Different charge = attract

- Larger charges produce a lower force; smaller distances produce a higher force

Shielding

- Electrons attracted to nucleus
- Repulsed by other electrons

Ionization Energy - energy required to remove an electron from an atom. Increases for successive electrons.

Atomic Radius

- Decreases across a period
- This decrease can be explained in terms of the increasing effective nuclear charge going from left to right. The valence electrons are drawn closer to the nucleus, decreasing the size of the atom.
- Increases down a group
- This can be explained due to the increasing orbital size in successive principle quantum levels.

Example

In terms of atomic structure, explain why the atomic radius of gallium is smaller than calcium?

Example

A student claims that the first ionization energy for F is greater than that of Br. Do you agree? Explain why or why not in terms of atomic structure and Coulomb's law.

1.8 Valence Electrons and Ionic Compounds

Valence Electrons

- Known as outer shell electrons
- The likelihood that two elements will form a chemical bond is determined by the interaction between valence electrons

Ion charges

- Elements in the same column tend to form analogous compounds

- Typical charges of atoms in ionic compounds are determined by their location on the periodic table and number of valence electrons.

Example

A student claims that a compound with sodium and fluorine and a compound with sodium and oxygen will have the same formula. Do you agree? Explain why or why not.

Example

Calcium reacts with a certain element to form a compound with the general formula CaX_2. what would be the most likely formula for a compound formed between sodium and element X?

PRACTICE MULTIPLE CHOICE BELOW GOODLUCK!

Directions: Select the letter that best completes the statement or answers the question presented.

1. How many moles of sulfate ions are in 100 mL of a solution of 0.0020 M $Fe_2(SO_4)_3$?

   a. $2.0 \cdot 10^{-4}$

   b. $6.0 \cdot 10^{-4}$

   c. $2.0 \cdot 10^{-1}$

   d. $6.0 \cdot 10^{-1}$

2. Which of the following equal volume solutions has the lowest conductivity?

   a. 0.1 M $CuSO_4$

   b. 0.1 M KOH

   c. 0.1 M $BaCl_2$

   d. 0.1 M HF

3. A 0.10 M aqueous solution of sodium sulfate, Na_2SO_4, is a better conductor of electricity than a 0.10 M aqueous solution of sodium chloride, NaCl. Which of the following best explains this observation?

   a. Na_2SO_4 is more soluble in water than NaCl is.

   b. Na_2SO_4 has a higher molar mass than NaCl has.

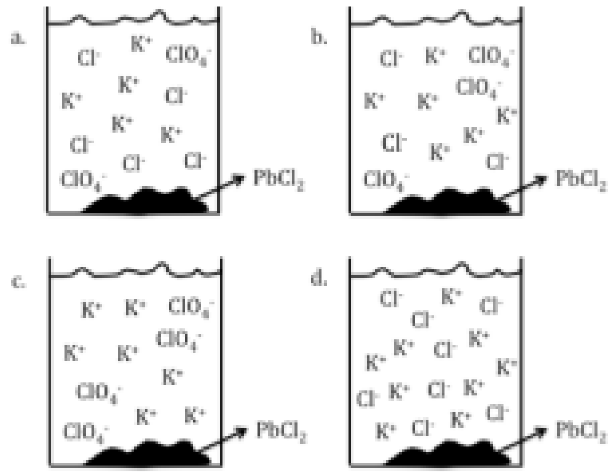
   c. To prepare a given volume of 0.10 M solution, the mass of Na_2SO_4 needed is more than twice the mass of NaCl needed.

d. More moles of ions are present in a given volume of 0.10 M Na_2SO_4 then in the same volume of 0.10 M NaCl.

4. How many moles of ions are present in 250 mL of a 4.4 M solution of sodium sulfate, Na_2SO_4?
   a. 1.1
   b. 2.2
   c. 3.3
   d. 13

5. If 200. mL of 0.60 M $MgCl_2$(aq) is added to 400. mL of distilled water, what is the concentration of Mg^{2+}(aq) in the resulting solution?
   a. 0.20 M
   b. 0.30 M
   c. 0.40 M
   d. 0.60 M

6. A chemist mixes a dilute solution of lead (II) perchlorate with an excess of potassium chloride, forming solid lead (II) chloride. Which of the following particulate views represents the experiment after the reactants are mixed thoroughly?



7. When this equation is balanced using the smallest possible integers, what is the sum of the coefficients?

$$\_\_(NH_4)_3PO_4 + \_\_CaCl_2 \rightarrow \_\_Ca_3(PO_4)_2 + \_\_NH_4Cl$$

   a. 8
   b. 9
   c. 11
   d. 12

Questions 8 - 10 should be answered on the basis of the equations below.

a. $NaCl(s) + H_2O(l) \rightarrow NaCl(aq)$

b. $HCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + HNO_3(aq)$ c. $2HCl(aq) + Mg(OH)_2(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$ d.

$2HCl(aq) + NaOCl(aq) \rightarrow NaCl(aq) + Cl_2(g) + H_2O(l)$

8. Which equation represents an oxidation-reduction reaction?

9. Which equation represents a precipitation reaction?

10. Which equation represents an acid base reaction?

11. How many H^+ ions are required when the number below is balanced with the smallest whole number coefficients? $Cu + NO_3^- + H^+ \rightarrow Cu^{2+} + NO + H_2O$

a. 2
b. 4
c. 6
d. 8

12. What is the oxidation number of rhenium in $Ca(ReO_4)_2$?

a. +1
b. +3
c. +6
d. +7

13. In which case does chromium undergo reduction?

a. $CrO_3 \rightarrow CrOF_3$

b. $Cr^{3+} \rightarrow Cr(OH)_{4^-}$

c. $2CrO_4^{2-} \rightarrow Cr_2O_{7^{2-}}$

d. $Cr^{3+} \rightarrow CrO_{4^{2-}}$

14. In the reaction,

$ClO_3^- + 5Cl^- + 6H^+ \rightarrow 3Cl_2 + 3H_2O$ the oxidizing and reducing agents are, respectively,

a. Cl^- and ClO_{3^-}
b. ClO_3^- and Cl
c. ClO_3^- and

https://drive.google.com/file/d/0B29R3FbvvOUOWm5lT21pMzQ4R1E/view?resourcekey=0-77O5YrMV-haufByOaxJJ_A **(ANSWER KEY) (COMPLETE TEST FIRST. DO NOT TRICK YOURSELF)**