Summaries of Review Topics for AP Chemistry

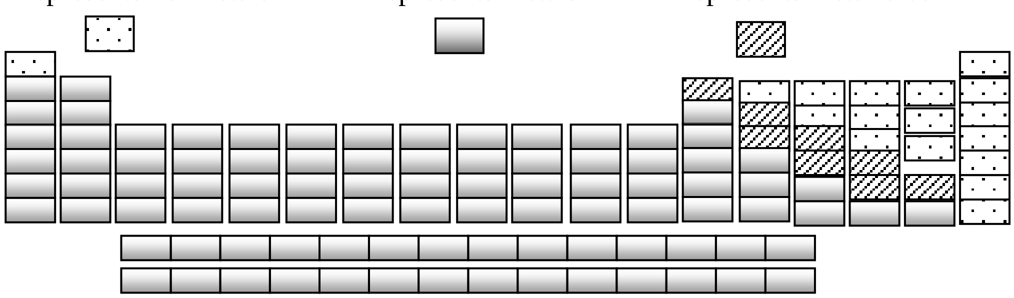
Classification of Matter

- Matter
  - Mixtures
    - Homogeneous
    - Heterogeneous
  - Pure substances
    - Elements
      - Metals
      - Semimetals
      - Nonmetals
    - Compounds → Ionic
      - Molecular Covalent
      - Network Covalent

(1) **<u>Element is a pure substance that cannot be broken down into other substances by chemical or physical means; it contains only one kind of atom.</u>** Elements are represented by the chemical symbols. The symbols for most elements consist of the first one or two letters of the name of the element. *Ex: H, O, He.*

▪ Elements can be classified *as metals, nonmetals, and metalloids (or semimetals).* In the **diagram** below you can find if an element is *metal, nonmetal, or metalloid (or semimetal).*

Represents nonmetals   Represents metals   Represents metalloids



▪ *Metals* are generally shiny solids. They are *good conductors of heat and electric current*. ***<u>Some of them (Fe, Co, Ni, Mn) are magnetic.</u>***

▪ *Nonmetals* are generally either gases or brittle solids at room temperature. Majority of nonmetals *do not conduct electricity and heat. Exception is graphite (C): it conducts heat and electricity.*

▪ *Metalloids (or semimetals*) have some of the characteristics of both metals and nonmetals.

(2) Elements are arranged in the periodic table in <u>7 horizontal rows called periods.</u>

(3) The elements in the <u>same column have similar properties; they are called a group or family</u>

<u>(4) Representative elements</u> - groups 1A through 8A , <u>transition elements</u> – B groups.

(5) 2 rows below the main part of PT are <u>lanthanides</u> (similar to lanthanum) and <u>actinides</u> (similar to actinium.)

(6) Some groups of elements have the following names:

Group <u>1A</u> elements: <u>alkali metals, the most reactive metals.</u>	**Group <u>7A</u> elements : <u>halogens, the most reactive nonmetals.</u>**
Group <u>2A</u> elements : <u>alkaline earth metals, a little bit less reactive than alkali metals.</u>	**Group <u>8A</u> elements : <u>noble (or inert) gases, unreactive elements</u>**

(2) <u>Each box on the Periodic Table contains certain information about the element (see the example below).</u>

<u>Groups (columns)</u>

6A ← <u>Column's (group's)</u> number

<u>Row's (period's)</u> number

2

<u>(horizontal)</u>

8 ← Element's (Atomic) number

O ← Element's symbol

<u>Oxygen</u> ← Element's name

15.999 ← Element's (Atomic) mass (in atomic mass Units, amu)

(3) A **Compound** is a <u>neutral substance</u> which consists of two or more different elements chemically bonded. Examples: $CuSO_4$.

(4) **An ion** is <u>a charged particle</u> with a charge shown in the upper right corner. Examples: P^{3-}, HSO_4^-

(5) According to the **Law of Definite Composition (proportions)**, each chemical compound has **a definite composition,** a composition that is always the same wherever that compound is found. A chemical formula tells several things about a compound: **it tells <u>what elements make up a compound (each element starts with a capital letter: the number of elements in a compound equals the number of capital letters in the formula).</u>**

(7) The formulas of compounds also contain numbers called **subscripts.** A subscript applies to the symbol that precedes it in the formula. **By convention, the subscript 1 is not written in formulas. The subscripts in a formula show the ratios of the atoms of the elements that combine to form the compound.** Thus, there are 2 atoms of H for 1 atom of S, and for 4 atoms of O in the compound H_2SO_4.

Type of Bonding: ionic, covalent, and metallic

Type of bonding	**Composition**	**Examples**
Ionic	Consist of metal and nonmetals	Ex: NaCl, $Al(OH)_3$, $Cu(NO_3)_2$
Covalent	Consist of nonmetals	Ex: HCl, H_2
Metallic	Unique for metals	Ex: all metals, pure elements such as Cu, and Alloys such as Cu_xZn_y (brass)

Writing Formulas of Ionic Compounds

(1) The simplest ratio of the ions represented in an ionic compound is called a **formula unit.**

(2) If an ion consists of only one atom, it's called **a monatomic ion.** The charge (or oxidation number/state) of a monatomic ion is equal to the number of electrons that were transferred from an atom. Ex: Ca^{2+} (the Calcium atom lost 2 valence electrons) and S^{2-} (S gained 2 electrons). An ion that contains more than one atom, is called a **polyatomic ion**. The charge of a polyatomic ion applies to the entire group of atoms. Example: MnO_4^-, PO_4^{3-}.

(4) In the formula of an ionic compound, the symbol of a cation is written before that of an anion. Subscripts, or small numbers written to the lower right of the chemical symbols, show the numbers of ions of each type present in a formula unit. If there is more than one polyatomic ion in a formula unit, parentheses are written around the formula of the ion, and a subscript is written after the final parenthesis.

(5) The easiest way to write the formula for an ionic compound is to use the **crisscross method**:

(To write the formula for an ionic compound you need to remember that the overall charge of the formula unit must be zero).

Example: Write the formula for calcium nitrate

STEP # 1

Calcium-ion has a charge (+2) because it's located in group 2A.

Nitrate- ion has a charge (-1) (find it in the Ions Chart.) Ca^{2+} NO_3^{1-}

STEP # 2 Switch numbers between cation and anion dropping signs (+) and (-). Around polyatomic ions (which consist of more than one atom) put parentheses if there is more than one. Answer: $Ca_1(NO_3)_2$, which is written $Ca(NO_3)_2$

Ion Charges for select Groups in the Periodic Table

	Group 1A	Group 2A	Group 3A	Group 4A	Group 5A	Group 6A	Group 7A	Group 8A
Charge on MonatomicIon	**1+**	**2+**	**3+**	**4+ metals**	**3- non-metals**	**2-**	**1-**	**0**
Examples of Ions	Li^+ Na^+	Mg^{2+} Ca^{2+}	Al^{3+}	Pb^{4+}	N^{3-}	O^{2-} S^{2-}	F^- Cl^-	
Examples of Ionic compounds	Li_2O Na_2O NaCl KCl	CaO $MgCl_2$	Al_2O_3 $AlCl_3$	PbO_2 $PbCl_4$	Mg_3N_2 NH_3	H_2S SO_2 SO_3	LiF KCl NaBr	
Examples of Covalent compounds					NH_3	H_2O	HCl	

IUPAC Naming Rules for Covalent Compounds

(IUPAC stands for International Union of Pure and Applied Chemistry)

Rule #1: Identify and name acids: acids are covalent compounds which formulas start with H (except H_2O and H_2O_2). Find their name in the "Names and Formulas of Acids" below. If the acid is made with a polyatomic ion, change the ending of the ion from –ate to –ic, or from –ite to –ous and add acid to the name. If the acid is made with a monatomic ion, use the prefix hydro + root + ic.

Rule #2: Identify and name binary compounds (which consist of two elements only).

♦ The first element in the formula is named first using the name of the element (from the Periodic Table, PT)

♦ For the second element in the formula, use the root from its original name plus the suffix "ide".

Roots: H- hydr-; B- bor-; C- carb; N- nitr-; O-ox-; F- fluor-; Si- silic-;P- phosph-; S-sulf- ;Cl- chlor-; Se- selen;Br-brom-; I - iod-.

♦ If there is more than one atom of an element in the compound, Greek prefixes are used to indicate how many atoms of each element are here in a compound: mono - for 1; di - for 2; tri – for 3; tetra – for 4 ; penta - for 5 ; hexa – for 6 ; hepta – for 7; octa – for 8 ; nano – for 9 ; and deca – for 10. !!!Mono- is not used as a prefix for the name of the first element.

!!! Use prefixes just for binary covalent compounds, NOT ionic compounds!!!

♦ Note: The ammonium ion (NH_4^+) is an example of a polyatomic cation made of all non-metals. Ammonium compounds should be named using the rules for ionic compounds.

Names and Formulas of Acids

Chemical formula	Name of the acid
H_2S	Hydrosulfuric acid
HF/ HCl /HBr / HI	Hydrofluoric acid/ Hydrochloric acid/ Hydrobromic acid/ Hydroiodic acid
H_2CO_3/ H_2SiO_3	Carbonic acid/ Siliconic acid
HNO_3 / HNO_2	Nitric acid / Nitrous acid
H_3PO_4	Phosphoric acid
H_2SO_4 / H_2SO_3	Sulfuric acid / Sulfurous acid

Examples of using the IUPAC nomenclature to write the names of covalent compounds
H_3PO_4 phosphoric acid , HCl hydrochloric acid, N_2O_3 dinitrogen trioxide

Examples of using the IUPAC nomenclature to names covalent compounds
nitric acid , HNO_3 ; Diphosphorus trioxide P_2O_3

Examples of using the IUPAC nomenclature to write the names of covalent compounds

▪ *Use the crisscross method to fill in the table and name each compound (write the name of a cation followed by the name of an anion).*
▪ *!!! For transition metals (B-groups elements) show their charges in parenthesis in roman numerals.*

	N^{3-} (nitride)	OH^{1-} (hydroxide)	NO_3^{1-} (nitrate)	SO_4^{2-} (sulfate)	CO_3^{2-} (carbonate)
Cu^{2+}	Cu_3N_2 and name is copper (II) nitride	$Cu(OH)_2$ and name is copper (II) hydroxide	$Cu(OH)_2$ and name is copper (II) nitrate	$Cu(OH)_2$ and name is copper (II) sulfate	$Cu(OH)_2$ and name is copper (II) carbonate

CHEMICAL REACTIONS AND BALANCING CHEMICAL EQUATIONS

(1) Statements called **equations** are used to represent *chemical change* / *"chemical reaction*

(2) Equations indicate **reactants**, or starting substances, and **products**, or substances formed during the reaction.

(3) Equations can be in the form of **word equations**, in which the reactants and products are indicated by their names. Word equations can be replaced by **skeleton equations**, which use chemical formulas rather than words to identify reactants and products.

Word equation: Solid lithium reacts with chlorine gas to produce a solid lithium chloride.

Skeleton equation: $Li(s) + Cl_2(g) \rightarrow LiCl(s)$

(4) The reactants are written to the left of an arrow, and the products are written to the right of the arrow:

(5) Plus signs are used to separate the different reactants or products.

(6) An arrow in a reaction stands for "yields" and points from the reactants toward the products.

reactants → products

(7) During a chemical change elements can combine to form compounds, compounds can be broken down into smaller compounds and /or elements, and compounds can react to form new compounds.

(8) **Law of Conservation of Matter for a Chemical Change: during chemical changes atoms are conserved, but rearranged forming new substances (with new physical and chemical properties).**

(9) When words in a word equation are replaced by chemical formulas, we get unbalanced (or skeleton) equation. The unbalanced equation needs to be balanced to make it complete. To balance an equation, chemists use numbers written in front of formulas. They are called coefficients. A coefficient multiplies the number of each atom in a chemical formula.

♦ $2HNO_3$ means 2H, 2N, and 6O (oxygen).

♦ $Mg(OH)_2$ means 1 Mg, 2O, and 2H because $(OH)_2$ means 2O and 2H.

$2\,Mg(OH)_2$ means 2Mg, 4O, and 4H because $(OH)_2$ already means 2O and 2H.

♦ $Mg(NO_3)_2$ means 1 Mg, 2N, and 6O because subscript outside of parenthesis multiplies the number of atoms of elements inside of parenthesis only; $2\,Mg(NO_3)_2$ means 2 Mg, 4N, and 12 O.

Example

$$? \boxed{HCl} + ? \boxed{Ca(OH)_2} \rightarrow ? \boxed{CaCl_2} + ? \boxed{H_2O}$$

Step#1: For this particular equation, start to balance Ca or Cl **(!!! any element except H and O!!!).** Compare the number of their atoms to the left of the arrow with ones to the right of the arrow. You have one atom of Ca on both sides. Thus, Ca is already balanced. To balance Cl, multiply HCl by the coefficient of 2.

Step#2: Balance number of hydrogen atoms: you have 4 atoms of hydrogen to the left of the arrow (two in 2HCl + two in $Ca(OH)_2$) and 2 atoms of hydrogen to the right of the arrow (in H_2O). To balance hydrogen, multiply H_2O by two.

Step#3: Balance number of oxygen atoms: you have 2 atoms of oxygen to the left of the arrow (two in $Ca(OH)_2$) and 2 atoms of oxygen to the right of the arrow (in $2H_2O$). Thus, oxygen is also balanced.

The balanced equation is: $2 \boxed{HCl} + 1 \boxed{Ca(OH)_2} \rightarrow 1 \boxed{CaCl_2} + 2 \boxed{H_2O}$

TYPES OF CHEMICAL REACTIONS

(1) *COMBINATION (SYNTHESIS)* : *Elements or compounds are combined to form only one compound as a product*

A + B → AB   or   AB + C → ABC   (reaction has only one product AB or ABC)

(2) *DECOMPOSITION* : *A single compound breaks down into two or more elements, or new compounds*

ABC → AB + C   (reaction has only one reactant ABC)

(3) *SINGLE –REPLACEMENT*: *Atoms of one element replace the atoms of another element in a compound.*

A + BC → AC + B (A and B are elements, BC and AC are compounds)

(4) *DOUBLE -REPLACEMENT* : *two compounds exchange their parts forming two new compounds*

AB + CD → AD + CB   (AB, CD, AD, and CB are compounds)

(5) *COMBUSTION (burning)*: *substance reacts with oxygen producing heat and light*

A + O_2 → compound or AB + O_2 → one or more compound   (one of the reactants is O_2)

DIMENSIONAL ANALYSIS AND MOLE CONVERSIONS

Dimensional Analysis, or Unit Convertion Method, or Factor Label Cancellation Method

Two basic rules are associated with the unit conversion method:

Rule # 1: Always write the unit and the number associated with the unit.

Rarely in chemistry will you have a number without a unit. Pi (= 3.14) is the major exception. Ex: 22.02 m, 22.02 s

Rule # 2: Carry out mathematical operation with the units, canceling them until you end up with the unit you want in the final answer.

Units are a part of all problems. If the units cancel out correctly, you know you have set up the problem correctly.

A word that is used a lot with units is "per". The term "per" can be denoted by a single line. For example, miles per gallon can be written as miles / gallon. The term "per" also gives a clue that division is involved here.

Sample problem: Convert 7 days to milliseconds

(a) Begin by drawing a long line and the equal sign.

(b) Place the unit that the question asks for to the right of the equals sign

(c) Start with given, 7 days. Put this data in the numerator to the first blank of the problem.

(d) Add the rest of the information needed to end up with the correct units in the answer:

1 day = 24 hr; 1 hr = 3,600 s; 1 s = 1,000 ms. Represent these relationships as conversion factors.

$$\frac{7\,\cancel{\text{day}}}{} \,\Bigg|\, \frac{24\,\cancel{\text{hr}}}{1\,\cancel{\text{day}}} \,\Bigg|\, \frac{3{,}600\,\cancel{\text{s}}}{1\,\cancel{\text{hr}}} \,\Bigg|\, \frac{1{,}000\,\text{ms}}{1\,\cancel{\text{s}}} = \qquad \text{ms}$$

Cancel out the same units if they appear on the opposite sides of a long line.

(e) Multiply all numbers above the long line, then all numbers below this line, and finally divide them. This is your answer. (Note: The vertical lines are a shorthand way of writing parentheses indicating multiplication.)

$(7 \cdot 24 \cdot 3{,}600 \cdot 1{,}000) \div (1 \cdot 1 \cdot 1) = 604{,}800{,}000$ ms

The Mole and Mole Conversions (Y- diagram)

(1) <u>One mole of a substance is the amount whose mass equals the molecular or atomic mass (weight) of a substance expressed in grams</u>. The mass of 1 mole of a substance is called its <u>molar mass, MW.</u>

<u>Unit of a molar mass is g/mol.</u>

Example: Molecular mass of Na_2O = (23.0 •2) + 16.0 = 62.0 amu; (MW) Na_2O = 62.0 g/mol

(2) <u>A mole is the amount of a substance that contains $6.02 \cdot 10^{23}$ particles.</u>

Distinguish the following <u>representative particles in a sample of matter: atoms, molecules, or formula units.</u>

▪ Particles are called <u>atoms</u> if they consist of a <u>single element</u>. *Example: Fe*

▪ Particles are called <u>molecules</u> if a compound consists of <u>nonmetals only</u>. Those compounds are called <u>covalent</u>. *Example: HClO, Cl_2*

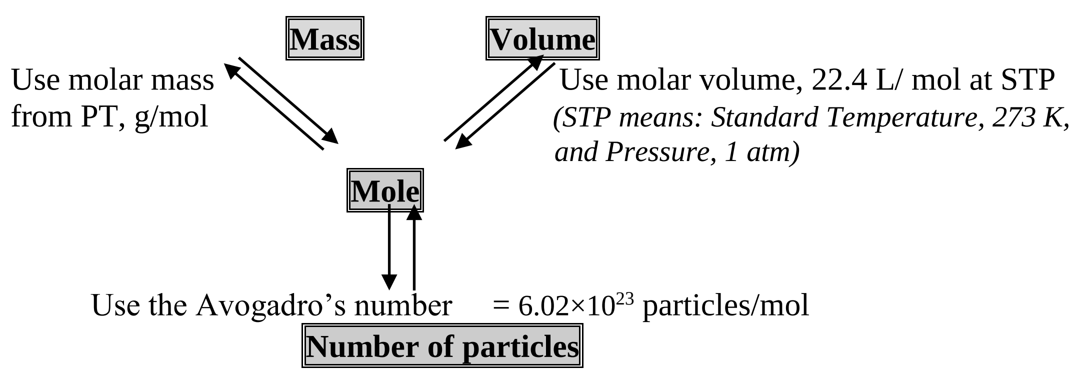
▪ Particles are called <u>formula units</u> if a compound consists of <u>metals combined with nonmetals</u>. Those compounds are called <u>ionic</u>. *Example: FeO, KClO*

▪ 1 mole contains $6.02 \cdot 10^{23}$ particles. $6.02 \cdot 10^{23}$ particles is the Avogadro's number.

▪ Molar weight (mass) is a mass of 1 mole of a substance = sum of atomic masses

▪ Molar volume is a volume of 1 mole of a gaseous substance = 22.4 L/mol at STP

The diagram below will help you to do <u>conversions between mass, volume, number of particles, and moles.</u>



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Examples

Mass to Mole Conversions

♦ If the given quantity is the mass of a substance in grams, multiply the given by a conversion factor which has grams in a denominator and 1 mol in a numerator.

Example#1 Determine the number of moles in 54.0 g of H_2O.

Given: 54.0 g of H_2O; molar mass of H_2O is 18.0 g/mol; conversion factor for this problem is (1 mol/18.0 g)

$$54.0 \text{ g of } H_2O \bullet \frac{1\text{mol of } H_2O}{18.0\text{g } H_2O}$$

Ans: 3.00 mol H_2O

Mole to Mass Conversions

♦ If given is a number of moles of a substance, multiply the given by conversion factor which has moles in a denominator and grams in the numerator.

Example # 2 Determine the number of grams in 5.00 moles of H_2O

Given: 5.00 mol of H_2O ; MW H_2O is 18.0 g/mol; conversion factor for this problem is (18.0 g/ 1 mol)

$$5.00 \text{ mol of } H_2O \bullet \frac{18.0\text{g } H_2O}{1\text{mol of } H_2O}$$

Ans: 90.0 g H_2O

Mole to Particle Conversions

♦ If given is the number of moles of a substance, multiply the given by the conversion factor 6.02×10^{23} particles/1mol. (Look at Example #3 below)

♦ If given is a number of particles of a substance, multiply the given by conversion factor 1 mol/ 6.02×10^{23} particles. (Look at Example #4 below)

Example # 3 How many molecules are there in 5 moles of H_2O?

$$5 \text{ mol } H_2O \bullet \frac{6.02\times10^{23} \text{ molec. } H_2O}{1 \text{ mol } H_2O} = 30.1\times10^{23} \text{ molec.} H_2O = 3.01\times10^{24} \text{ molecules } H_2O$$

(in proper scientific notation)

Example # 4 How many moles are there in 1.204×10^{24}atoms of Na?

$$1.204\times10^{24}\text{at.Na} \bullet \frac{1\text{mole Na}}{6.02 \bullet 10^{23}\text{at.Na}}$$

Ans: 2.0 mol Na

DIMENSIONAL ANALYSIS & STOICHIOMETRY

Stoichiometry

♦ **Stoichiometry is the calculation of amounts of substances involved in a chemical reaction.**

♦ **Theoretical number of moles of reactants and/or products equal to the coefficients before these substances.**

Example

Use the balanced equation below to determine the following mole/mole ratios in this reaction.

$$3\ Ag_{(s)} + 4\ HNO_{3\,(ag)} \rightarrow 3\ AgNO_{3\,(ag)} + 1\ NO_{(g)} + 2\ H_2O_{(l)}$$

Example: Moles of $AgNO_3$ produced per moles of HNO_3 used: Answer: **3:4**

Moles of HNO_3 used per moles of Ag used : Answer: **4:3**

To solve stoichiometric problems using dimensional analysis method, follow the following steps:

(1) Write down and balance the chemical equation or use the given equation.

(2) Identify given and unknown: the number of moles, mass, volume (of gases), or the number of particles.

(3) Write the information that is given and wanted above the corresponding chemical formulas in the equation.

(4) Make the appropriate conversions from the given data to the requested quantity using the factor label cancellation method: convert given (in grams, or liters, or particles) to moles of this substance.

(a) To convert g ↔ mol, use the molar mass (MW) from the PT as a conversion factor.

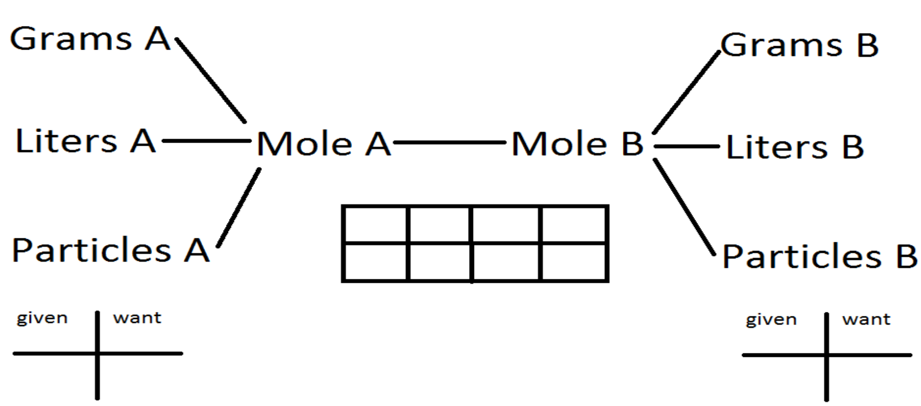
(b) To convert mol ↔ particles (atoms, molecules, formula units, electrons, etc.), use the Avogadro's number ($6.02 \cdot 10^{23}$ particles/mol) as a conversion factor.

(c) To convert **liters of a gas** ↔ mol of a gas, use the molar volume (22.4L/mol) as a conversion factor.

(5) Next conversion factor is mole-mole ratio: use coefficients from the balanced chemical equation (ratio of coefficients) as a conversion factor.

(6) Next conversion factor converts moles of unknown to g, or liters, or particles of unknown.

Extended Mole Map



Stoichiometry Examples

(1) How many moles of ethane, C_2H_4, need to burn completely in order to obtain 3.60 g of H_2O ?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X mol}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow 2CO_2 + \underset{\textbf{Given}}{\overset{\textbf{3.60 g}}{\underline{2\,H_2O}}}$$

$$3.60\text{ g }H_2O \cdot \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }H_2O}$$

X= 0.100 mol C_2H_4

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(2) How many grams of ethane, C_2H_4, need to burn completely in order to obtain 3.60 g of H_2O ?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X g}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow 2CO_2 + \underset{\textbf{Given}}{\overset{\textbf{3.60 g}}{\underline{2\,H_2O}}}$$

$$3.60\text{ g }H_2O \cdot \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }H_2O} \cdot \frac{28.0\text{ g }C_2H_4}{1\text{mol }C_2H_4}$$

X= 2.80 g C_2H_4

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(3) How many molecules of ethane, CH_4, need to burn completely in order to obtain 3.6 g of H_2O ?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X molecules}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow 2CO_2 + \underset{\textbf{Given}}{\overset{\textbf{3.60 g}}{\underline{2\,H_2O}}}$$

$$3.60\text{ g }H_2O \cdot \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }H_2O} \cdot \frac{6.02 \bullet 10^{23}\text{ molecules/mol}C_2H_4}{1\text{mol }C_2H_4}$$

X= $6.02 \bullet 10^{22}$ molecules C_2H_4

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(4) How many liters of ethane, C_2H_4, need to burn completely in order to obtain 3.60 g of H_2O ?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X L}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow 2CO_2 + \underset{\textbf{Given}}{\overset{\textbf{3.60 g}}{\underline{2\,H_2O}}}$$

$$3.60\text{ g }H_2O \cdot \frac{1\text{ mol }H_2O}{18.0\text{ g }H_2O} \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }H_2O} \cdot \frac{22.4\text{ L }C_2H_4}{1\text{mol }C_2H_4}$$

X= 2.24 L C_2H_4

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(5) How many moles of ethane, C_2H_4, need to burn completely in order to obtain 100. mol of H_2O ?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X mol}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow 2CO_2 + \underset{\textbf{Given}}{\overset{\textbf{100. mol}}{\underline{2\,H_2O}}}$$

$$100.\text{ mol }H_2O \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }H_2O}$$

X= 50.0 mol C_2H_4

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(6) How many liters of C_2H_4 need to burn in order to obtain 50.0 L of CO_2 at STP (273 K,1atm)?

$$\underset{\textbf{Unknown}}{\overset{\textbf{X L}}{\underline{1C_2H_4}}} + 3O_2 \rightarrow \underset{\textbf{Given}}{\overset{\textbf{50.0L}}{\underline{2CO_2}}} + 2\,H_2O$$

$$50.0\text{ L }CO_2 \cdot \frac{1\text{ mol }CO_2}{22.4\text{ L }CO_2} \cdot \frac{1\text{ mol }C_2H_4}{2\text{ mol }CO_2} \cdot \frac{22.4\text{ L }C_2H_4}{1\text{mol }C_2H_4}$$

X= 25.0 L C_2H_4

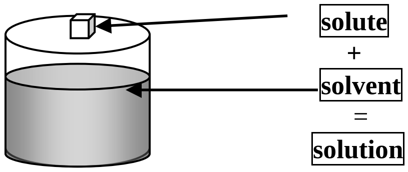
!!!!(Notice that the 22.4 L/mol conversions are cancelled out, therefore the ratio of liters of gases at the same temperature and pressure is the same as the ratio of moles).

CONCENTRATION OF SOLUTIONS

(1) **A solution** is a mixture of two or more substances that are uniformly mixed with one another.

(2) **A solute**- substance that is being dissolved.

(3) **A solvent**- substance in which the solute is dissolved.



Mass of solute + Mass of solvent = Mass of solution

(4) **A <u>concentrated</u> solution** (versus **<u>dilute</u>**) contains a <u>large amount of solute.</u>

(5) <u>The concentration of a solution is a measure of the amount of solute that is dissolved in a given quantity of solvent or solution.</u>

(6) Substances that dissolves in a solvent is said to be **<u>soluble</u>** <u>(versus</u> **<u>insoluble).</u>**

MOLARITY or MOLAR CONCENTRATION

Molarity (M) equals the number of moles of solute completely dissolved in 1 liter of solution.

Formula Equation:

$$M = \frac{\text{\# moles of solute}}{\text{\# liters of solution}}$$

Units**: **M which is abbreviation for mole/1L; *The volume* ***MUST BE IN LITERS*** *(1 L = 1000 mL)*

When you see M you may interpret it as the word MOLAR: 0.5 M HCl is read as a "0.5 Molar solution of HCl".

Conversions between grams, liters, and molarity (mol/L)

Many times you will not be given the ***number of moles directly.*** *You will have to calculate the number of moles and then apply it to the molarity equation.*

$$\text{Number of moles} = \frac{\text{given mass}}{\text{Molar mass (MW)}}$$

You also can use the Dimensional Analysis to convert grams to moles.

Example. A student made a 5.0 L solution by dissolving 80.0 grams of NaOH into water. Calculate the molarity of the solution.

<u>Step 1</u>: Calculate the number of moles from grams

Na + O + H
23 + 16 + 1

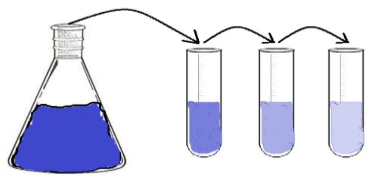
$$80.0\text{ g NaOH} \times \frac{1\text{ mol NaOH}}{40.0\text{ g NaOH}} = \boxed{2.0\text{ mol NaOH}}$$

<u>Step 2</u>: Calculate the molarity using the formula

$$M = \frac{\text{\# moles of solute}}{\text{\# liters of solution}} = \frac{2.0\text{ moles}}{5.0\text{ L}} = \boxed{0.40\text{ M}}$$

MAKING DILUTIONS

Stock solutions are big volumes of solutions of known concentration. Typically they have a high molarity. Solutions that we use in lab are weakened so we don't get hurt. **You can prepare less concentrated (dilute) solution from more concentrated by adding solvent. The number of moles of solute does not change when a solution is diluted.**



$$\boxed{\textbf{Formula equation: } M_1 \cdot V_1 = M_2 \cdot V_2}$$

In this equation:

M_1 and V_1 represent the molarity and volume of the concentrated solution;

M_2 and V_2 represent the molarity and volume of the dilute solution.

Example. A student used 0.50 L of 8.0 M stock solution of NaOH to prepare 5.0 L of a dilute solution. Calculate the molarity of the dilute solution.

$$M_2 = (M_1 \cdot V_1)/\ V_2 = (8.0\ M \cdot 0.50\ mL)/\ 5.0\ L = \underline{0.80\ M}$$

ATOMIC STRUCTURE and ELECTRON CONFIGURATIONS

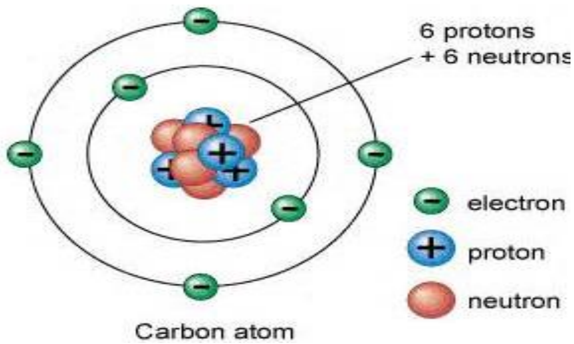
Some discoveries are described in the table below:

Year	Scientist(s)	Discovery
1803	John Dalton- the "father" of modern chemistry	The main points of his theory: ♦ All matter is composed of atoms. ♦ An atom is the smallest particle of an element that maintains the properties of that element. ♦ Atoms of one element are different from atoms of other elements. ♦ Atoms cannot be created, divided into smaller particles, or destroyed. ♦ In a chemical reaction, atoms are separated, combined, or rearranged
1897	J.J.Thomson	♦ Studied "cathode rays". Proposed the "plum pudding model" of the atom. Identified the first subatomic particle, the electron. Measured the charge-to-mass ratio of an electron by the deflection of the beam of electrons with an electric field.
1898	Rutherford	♦ Studied radiations emitted from uranium and thorium and named them alpha and beta. Discovered the nucleus. Proposed the nuclear model of the atom.
1909	Robert Milikan	Determined the charge of electron ("oil drop experiment").
1900	Max Planck	Used the idea of quanta (discrete units of energy) to explain hot glowing matter.
1905	Albert Einstein	♦ Explained the photoelectric effect. ♦ Published the famous equation $E = mc^2$ m which means that a loss or gain in mass accompanies any reactions that produces or consumes energy.
1914	H.G.Moseley	Determined the charges of the nuclei of most atoms and found that the atomic number of an element is equal to the number of protons in the nucleus. This work was used to reorganize the periodic table based upon atomic number instead of atomic mass.
1922	Niels Bohr	Proposed the atomic model with successive energy levels of electrons.
1923	De Broglie	Discovered that electrons had a dual nature-similar to both particles and waves. Supported Einstein.
1927	Heisenberg	Proposed Principle of Uncertainty - you cannot know both the position and velocity of a particle.
1930	Schrodinger	Viewed electrons as having wave-like properties. Introduced Orbitals, cloud-like regions in which there is a probability of finding the electron. Introduced "wave mechanics" as a mathematical model of the atom.

Table below summarizes information about the three <u>fundamental subatomic particles</u>.

(Amu is used to express the mass of these particles because it's more practical than using the extremely small values expressed in grams: 1 amu = 1.66×10^{-24} g or 1 g = 6.022×10^{23} amu).

Subatomic particle	Abbreviation	Location	Charge of each	Mass of each
Proton	p^+	***Nucleus***	***+1***	***1 amu***
Neutron	n^0	***Nucleus***	***0***	***1 amu***
Electron	e^-	***Electron cloud or an orbital***	***-1***	***1/1840 amu***



(1) The nucleus is the positively charged central core of an atom composed of protons and neutrons. Because protons and neutrons have a much greater mass than electrons (see the table above) *almost all of the mass of an atom is concentrated in a tiny and very dense nucleus.*

(2) The electrons surround the nucleus and occupy most of the volume of the atom. *Space occupied by electron cloud (its volume) is approximately 10,000 times bigger than the space occupied by the nucleus.*

(3) For a neutral atom, **Number (#) of p^+ = Number (#) of e^- = Atomic Number**

(4) *Protons, not neutrons, determine an element's identity*. Although all atoms of an element have the same number of protons, most elements have atoms with different numbers of neutrons. <u>Atoms with the same number of protons but different number of neutrons and mass number are called ***isotopes***.</u> An isotope of an element is identified by its mass number; the mass of electrons is negligible and can be ignored.

Mass number = Number of protons + Number of neutrons

(5) Recall that the **<u>number of protons</u>** in an atom's nucleus **<u>equals to its atomic number</u>**. Substituting atomic number for the number of protons gives the following relationship.

Number of neutrons = Mass number – Atomic number

(6) There are two ways of writing the atomic symbols: <u>symbolic notations.</u>

(a) When writing the symbol for an isotope, write the **atomic number on the lower left corner (subscript) and mass number on the upper left (superscript) as shown in the following examples:**

H atom with one proton and <u>no neutrons</u> and H atom with one proton and <u>one neutron</u>

mass number→ $_1^1$**H *<u>(protium)</u>*** **mass number→** $_1^2$**H (*<u>deuterium</u>*)**

(b) Sometimes the name of the element is written out **followed by a dash and the mass number of the isotope.** For example, hydrogen's isotopes can be written as **H-1 or hydrogen-1** (protium), **H-2** or **hydrogen-2** (deuterium), and **H-3** or **hydrogen-3** (tritium).

(1) Atomic Number = number (#) of p^+ = number (#) of e^-

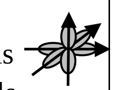
(2) Number of energy levels = the number of a period (row)

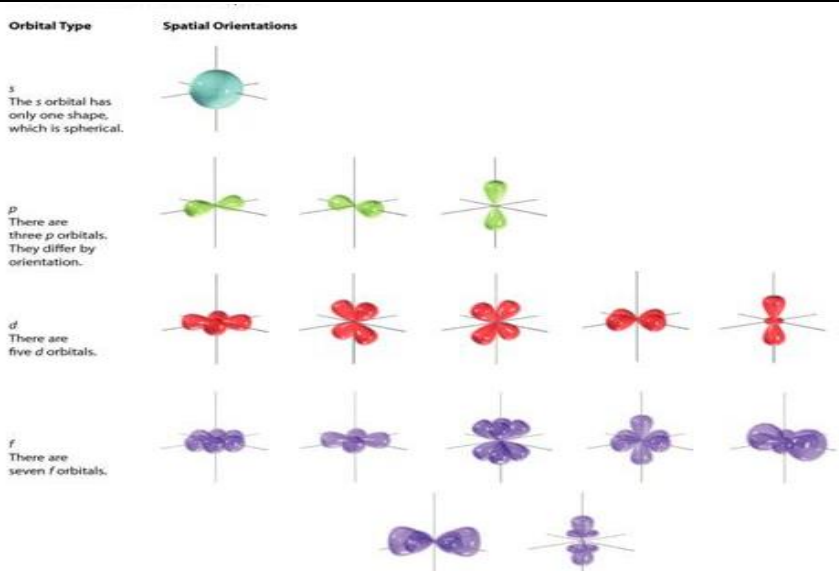
(3) Number of valence electrons (outermost electrons) = the number of a group (column)

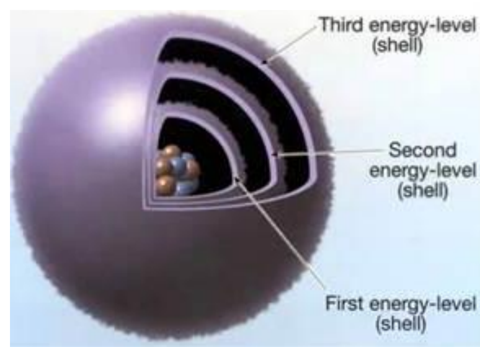
Electron Configurations

The electron configuration of an atom describes the arrangement of electrons in levels, sublevels, and orbitals (according to Aufbau Rule below).

1s→2s→2p→3s→3p→4s→3d→4p→5s→4d→5p→6s→4f→5d→6p→7s→5f→ 6d→ 7p

Energy level (orbit or shell)	Energy sublevels on this level	**Electrons on sublevels:** s - maximum 2 electrons, 1 orbital  p - maximum 6 electrons, 3 orbitals  d - maximum 10 electrons, 5 orbitals f - maximum 14 electrons, 7 orbitals
1 level	S	2 e^- on s-sublevel
2 level	s , p	(2 e^- on s-sublevel) + (6 e^- on p-sublevel) = 8 e^-
3 level	s, p , d	(2 e^-on s-subl) + (6 e^- on p-subl) + (10 e^- on d-subl) = 18 e^-
4 level	s , p , d, f	(2 e^-on s-subl) + (6 e^- on p-subl) +(10 e^- on d-subl) + (14 e^- on f-subl) = 32 e^-





A typical electron configuration consists of numbers, letters, and superscripts:

(a) A number indicates the energy level (this number is called principal quantum number).

(b) A letter indicates the type of orbital: s,p,d,f

(c) The number and letter together refer to the energy sublevel: 1s, 2p, etc.

(d) A superscript indicates the number of electrons in the orbital.

Ex.$1s^2$ means that there two electrons in the *s*-orbital of the first energy level. The element is He.

♦♦♦♦ *To write an electron configuration* ♦♦♦♦

(1) Use the Aufbau diagram (**1s→2s→2p→3s→3p→4s→3d→4p→5s→4d→5p→6s→4f→5d→6p→7s→5f→ 6d→ 7p)** to fill the orbitals with electrons. The Aufbau rule requires that electrons fill the lowest energy orbitals first: e.g. atoms are built from the ground upwards.

(2) The sum of the superscripts should equal the total number of electrons.

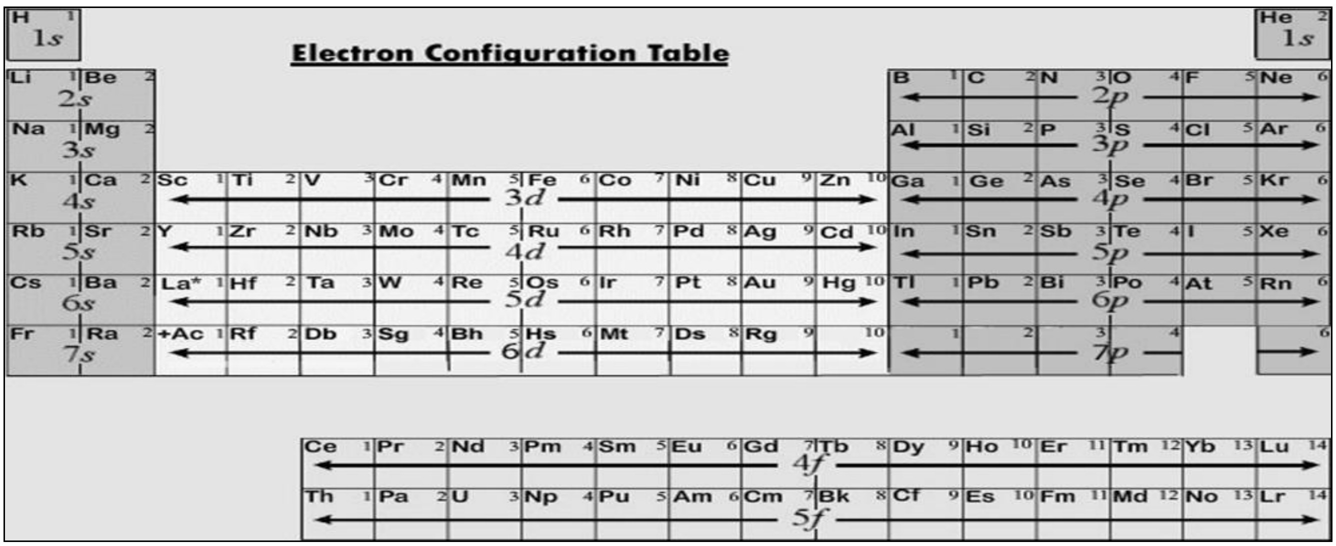
Ex: $1s^2$ $2s^2 2p^6$ $3s^2 3p^4$ 16 electrons, therefore S

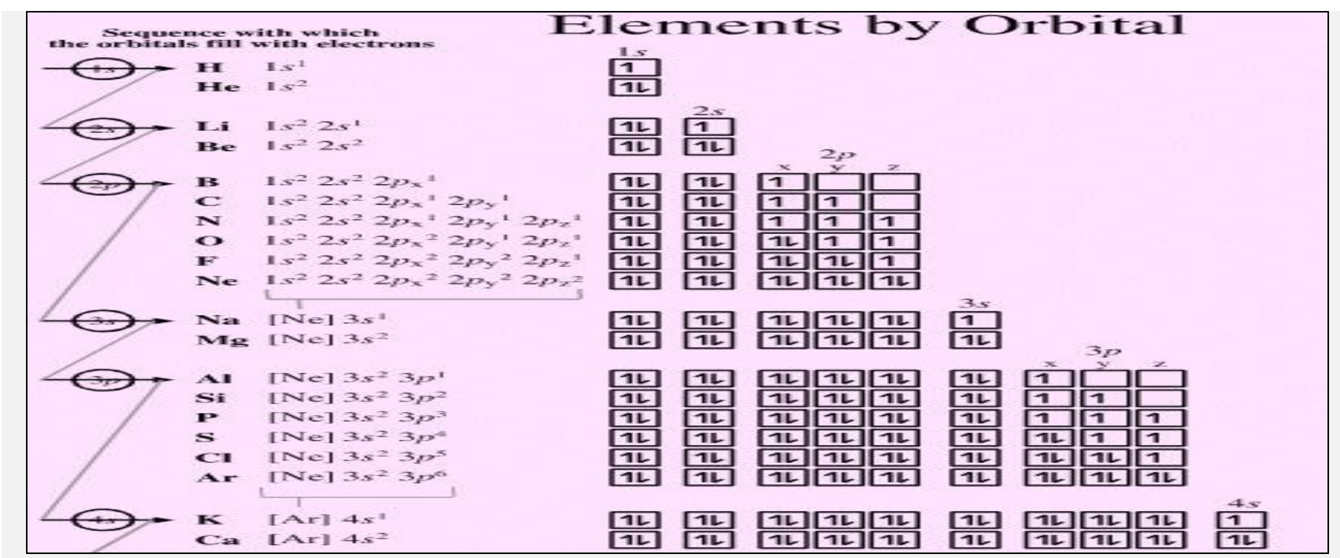
(3) Remember that electrons occupy the lowest energy orbital first.

(4) Remember that orbitals are considered to be in the same energy level if they have the same first number (no matter in what order filling is done).

(5) Remember that the outer shell is the highest numbered shell which has electrons in it. Only s and p orbitals are part of the outer shell.

(6) Electron configurations for which electrons occupy the lowest energy orbital first are called ground state electron configurations. If any lower energy level orbitals are not completely filled, and electrons are in higher energy levels or sublevels, this is referred to as an excited state electron configuration.





International System of Measurement (SI System)

Chemists must be able to communicate their measurements to other chemists all over the world, so they need to speak the same measurement language. This language is ***the SI system of measurement*** (from the French " Systeme International d'Unites") commonly referred as the ***metric system***. There are actually minor differences between the SI and metric system, but for the most part, they are interchangeable. The SI system is ***a decimal system***. There are base units for mass, length, volume, and so on, and there are ***prefixes that modify the base units***. For example, kilo- means 1,000; a kilogram is 1,000 grams, and a kilometer is 1,000 meters.

Units of Measurement

Physical quantity	**Unit/ abbreviation**	**Other units**	**English/ SI conversion**
Length (distance)	Meter/ m	km, cm, mm, dm,etc.	1 mile (mi) = 1.61 km 1 yard (yd) = .914 m 1 inch (in) = 2.54 cm 1 foot (ft) = 30 cm
Mass / weight	Gram/ g or kilogram/kg	kg, cg, mg, dg, etc.	1 pound (lb) = 454 g 1 kg = 2.2 lb 1 ounce (oz) = 28.4 g 1 pound = 16 oz 1 carat (car) = 200 milligrams (mg)
Volume (amount of space)	Cubic meter/m^3 or liter/ L 1 L = $(0.1m)^3$	km^3 , cm^3, mm^3,dm^3, etc. or kL, cL, mL, dL, etc.	I quart (qt) = .946L 1 pint (pt) =.473 L 1 fluid ounce (fl oz) = 29.6 mL 1 gallon (gal) = 3.78 L 1mL = $1cm^3$ -the same amount of space
Temperature	Degree Kelvin /K or Celsius / ^{0}C	K, ^{0}C	**Celsius to Fahrenheit: $^0F = 1.8\ ^0C + 32$** **Fahrenheit to Celsius: $^0C = 5/9\ (^0F - 32)$** **Celsius to Kelvin: $K = ^0C + 273$**

Base Units of Measurement (SI System)

Physical quantity	***Base unit/ abbreviation***
Length (distance)	Meter/ m
Mass / weight	Kilogram/kg
Temperature	Degree Kelvin / K
Time	Second/ s

Derived Units of Measurement (SI System)

(a) are those formed by combining base units or multiples of base units. m/s, m^2 etc.
(b) multiples of base units: kilometer, millimeter etc.

Prefixes that Modify the Base Units (m, kg, J, s, K, mole)

Shaded area is used for prefixes which make units bigger

Prefix	**Abbreviation**	**Meaning**	**Examples**
Giga-	G	10^9	1 gigameter (Gm) = 1,000,000,000 m or 10^9 m 1 m = 0.000 000 001 Gm or 10^{-9} Gm
Mega-	M	10^6	1 megameter (Mm) = 1,000,000 m or 10^6 m 1 m = 0.000 001 Mm or 1 m = 10^{-6} Mm
Kilo-	K	10^3	1 kilometer (km) = 1,000 m or 10^3 m 1 m = 0. 001 km or 1 m = 10^{-3} km
Deci-	D	1/10 or 10^{-1}	1 meter = 10 dm or 10^1 dm 1dm = 10^{-1} m
Centi-	C	1/100 or 10^{-2}	1 meter = 100 cm or 10^2 cm 1cm = 10^{-2} m
Milli-	M	1/1000 or 10^{-3}	1 meter = 1,000 mm or 10^3 mm 1mm = 10^{-3} m
Micro-	M	10^{-6}	1 meter = 1,000,000 μm or 10^6 μm 1 μm = 10^{-6} m
Nano-	N	10^{-9}	1meter= 1,000,000,000 nm or 10^9 nm 1 nm = 10^{-9} m

Significant Figures

(1) Significant figures (sig.fig.) are the number of digits that you report in the final answer of the mathematical problem you are calculating. The number of digits that a person reports in his or her final answer is going to give a reader some information about how accurately the measurements were made. The number of sig. fig. is limited by the accuracy of the measurement.

(2) If you are asked about number of automobiles in your household, your answer might be 0 or 2, but you would know exactly how many autos you have. Those are what are called *counted numbers.*

(3) If you are asked how many inches there are in a foot, your answer will be 12. That is an *exact number*. In both exact and counted numbers, there is no doubt what the answer is.

(4) If you make measurements, these measured numbers always have some error associated with them. You need to report all the numbers you know with certainty (where the scale has lines) PLUS one more figure of estimation (between the lines.)

(5) Rules of sig. fig. in a measured number:

All non-zero numbers are always significant *The zeroes are the only numbers that you have to worry about*	567.8907	7 s.f.
Zeros between non-zero numbers are always significant	606	3 s.f.
All zeroes to the left of the first nonzero digit are not significant	0.00241	3 s.f.
Zeroes to the right of the last nonzero digit are significant if there is a decimal point in the number	2098.000 0.00200	7 s.f. 3 s.f.
Zeroes to the right of the last nonzero digit are not significant if there is no decimal point in the number	2098000	4 s.f.
In scientific notation (**a** × **10**$^{\mathbf{b}}$) look at **a** to determine the number of sig.figs	18.06×10^{23}	4 s.f.

(6) Reporting the Correct Number of Sig. Fig. after Calculations

(a) Addition and Subtraction

In addition or subtraction, your answer should be reported to the number of decimal places used in the number that has the fewest decimal places.

Ex: 13.001 + 2.01 + 3.11111 = 18.12

Your calculator will show 18.123111, but you are going to round off to the hundredth place based on the 2.01, because it has the fewest number of decimal places. (All the numbers that are added have a significant figure in the hundredth place.) You then round the figure off to the 18.12

(b) Multiplication and Division

When you multiply or divide numbers, your answer must have the same number of s.f. as the measurement with the fewest significant figures.

$$2.2 \times 30000 = 66000$$

2.2 has 2 s.f., 30000 has 1 s.f., so round the answer to 1 s.f. The final answer is 70000 or 7×10^4