CHEM 208

Module 2:

List the steps of the scientific method:

Scientific Method: performing a study in organized step: (led to enunciation of Law of Conversation of Mass and Matter)

1. **Performing experiments:**
    a. An <u>experiment</u> is a set of steps (procedures) that are performed under controlled conditions to propose or test a hypothesis
2. **Making observations**
    a. <u>Observations:</u> a key factor in scientific studies as well as in everyday life, can be:
        i. <u>Qualitative:</u> Observation *does not* use numbers
            - Ex: flower is purple, bubbles are produced when lemon juice is added to baking soda
        ii. <u>Quantitative:</u> Observation is *measurement*. It has 2 components:
            - *A measured quantity (numerical value) with an appropriate unit*
            - Ex: Observed temperature (measured using a thermometer) – the measured quantity is temperature and the appropriate unit is *C
3. **Prosing a hypothesis** (based on the observations)
    a. <u>Hypothesis:</u> tentative explanation to account for the observations of an experiment. It is valid provided that one's assumptions to explain the observations of an experiment can be tested
4. **Confirming the hypothesis** (repeated experimentation)
    a. The validity of the hypothesis needs to be confirmed via repeated and controlled experiments
    b. In order to accept a hypothesis, there must be no inconsistencies between the hypothesis and the experimental observations – if there are any inconsistencies, step 1-4 must be repeated
5. **Proposing a scientific law**
    a. By repeatedly performing experiments and modifying the hypothesis to account for the observations from these experiments, one is able to propose a scientific law.

Numerical Values: certain rules apply to the expression of the numerical value and the system of units used in scientific measurements, include:

Scientific notation: very convenient way of expressing very large or very small numbers and at the same time provides a method of increasing efficiency in scientific calculations

- Mathematical expression in which a number is expressed as :

$$\mathbf{N \times 10^{\pm n}}$$

    - **N:** contains only one nonzero digit to the left of the decimal – can be any number other than zero, and can only have one digit before the decimal
        - Can be 4.56 but no 78.54
    - **n:** integer: represents how many times you are multiplying N by 10, how many times you move the decimal one place – whole number, can be positive or negative
        - Can be 7 but not 7.4
    - **Power of 10:** the number 10 raised to a certain power
    - **Ex:** $598.46 = 5.9846 \text{x} 10^2$

Measurement	Scientific Notation
1.23 g	1.23×10^{0} g
1000.4 mL	1.0004×10^{3} mL
0.0012 mg	1.2×10^{-3} mg
28.32×10^{15} L	2.832×10^{16} L
21.0 million kg	2.10×10^{7}kg

Significant figures (digits): (sig.fig.) in a measured number include all certain digits and one uncertain digit

- Used to indicate the precision (reproducibility (closeness) of measurements) of a measurement or that of a calculated result using such measurements
- <u>Does not apply to counted items such as exact numbers (5 cars, 3 houses) and defined conversion factors (1 inch = 2.54 cm)</u>
- **Rules for determining significant figures:** every measurement and thus all calculations performed by using such measurements always have some uncertainty (a characteristic of any measurement).
    - This is due to experimental error
    - In the case of measuring devices, the last digit of measurement is uncertain
    - If Patrick has 3 cars in his driveway → number has an infinite number of sig.fig.

- **Rule 1: All nonzero digits are significant figures:**

Measurement	No. of significant figures
135.62g	5
23.6cm	3
17.9871kg	2

- **Rule 2: Counting begins from the left with the first nonzero number –** (leading zeros are not significant)

Measurement	No. of significant figures
0.056 mL	2
0.000356 L	3
0.0225 kg	3

- **Rule 3: Zeros between nonzero digits are counted as significant figures**

Measurement	No. of significant figures
1.056 g	4
30.78 cm	4
300.5 mL	4

- **Rule 4: Terminal zeros (zeros to the right of a number):**
    - Always significant if the value contains a decimal point : 2.3700 g → 5 (sig.fig)
    - Terminal zeros in other cases may or may not significant: a measured value given as 200cm does not signify anything regarding the number of significant figures. Sometimes a decimal

point makes the zero significant, but mostly scientific notation is used:

Measurement	No. of significant figures
200 cm	3
2.00×10^2 cm	3
2.0×10^2	2
2×10^2	1

Rules for Significant Figures in Chemical Calculations: In calculation involving measured values (with a certain number of significant figures), the number of significant figures in the final answer depends on the operation performed.

- **Rules 1:** In **addition** and **subtraction** of measured quantities, the final answer contains the same number of decimal places as are in the measurement with **least number of decimal places** (least precise measurement)
  - **Ex:** 25.0g + 22.41g +1.234g = 48.6g → measurement 25.0g has the least number of decimal places → therefore the answer has one decimal place
  - **Ex:** 4.84 mL – 0.02mL = 4.83 mL → both measurements have 2 decimals.

- **Rule 2:** In **multiplication** and **division** of measured quantities, the final answer contains the same number of significant figures as are in the measurement with the **least number** of **significant figures**
  - **Ex:** 9.20g (3 sig.fig.) x 2.450 (4 sig.fig) = 22.5g → 9.20g has the least sig.fig
- **Rule 3:** In the final answer of calculation involving exact numbers, unit conversion factors and constants, the number of significant figures is dictated by the measured quantity involved
  - Exact numbers, unit conversion factors and constants have no effect on the number of significant figures
  - **Ex:** if the mass of steel ball is 2.35g (measured quantity), what will be the total mass of 8 (exact number) identical steel balls?
    - The total mass of 8 identical steel balls is: 2.35g x 8 = 18.8g
  - **Ex:** How many centimeters are there in 31.10 inches (measured quantity)?
    - The unit conversion factor is 1.0 in = 2.54 cm, 31.10 in x 2.54 = 78.99 cm
  - **Ex:** Using the formula: c = a x b, calculate the value of b when a = 22.65 cm (measured quantity) and c = 1.2 cm/s (constant)
    - 1.2 cm/s = 22.65 cm x b → b = 1.2/22.65 s^{-1}

Rules for Rounding off Chemical Calculations:
- **Rule 1: For multi-step calculations**
  - All numbers are carried to the final result, which is then rounded off to give the correct number of significant figures
- **Rule 2: If the digit being rounded off is > 5:** You round the digit up by one
  - Ex: 8.379 g, rounded to 3 significant figures, 8.38g
- **Rule 3: If the digit being rounded off is < 5:** the preceding digit remains the same
  - Ex: 6.371 g rounded to 3 sig.fig. , 6.37g
- **Rule 4: If the digit being rounded off is equal to 5:** the preceding digit is rounded up if it is odd and remains the same if it is even
  - Ex: 16.75 mL → 7 is odd → 16.8 mL
  - Ex: 16.65 mL → 6 is even → 16.6 mL

SI System of Units (based on the metric system):

- **SI:** Systeme International d'Unites, based upon the metric system, uses SI base unites.
  - Some Si based and Si derived units include
    - Length　　Temperature
    - Volume　　Time
    - Mass
    - Density
  - **Length:** The SI unit for length is the meter (m)
  - **Volume:** The SI derived unit for volume is cubic meter (m^3). The most common unit for volume is liter (L) and millimeter (mL)
    - Volume depends on the size of the object.
    - Common types of equipment used to measure liquid volume: Laboratory buret and household measuring cups

$$\mathbf{1\ L = 1 dm^3 = 10^{-3} m^3}$$
$$\mathbf{1\ mL = 1\ cm^3 = 10^{-3} dm^3 = 10^{-3} L = 10^{-6} m^3}$$

  - **Mass:** The SI unit of mass is kilogram (kg).
    - Mass refers to the amount of matter an object contains, thus mass is a constant. Like volume, mass also depends on the size of an object
  - **Density:** The SI unit for density is the kilogram per cubic mater (kg/m^3).
    - Generally, in chemistry, density is expressed in the units of g/mL or g/L
    - Density is a physical property of substance

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

    - Larger or smaller units than the base unit can be obtained by combining the base unit with one of the prefixes
      - Ex: Combining the base unit second with the prefix nano gives nanoseconds (symbol: ns = 10^{-9}s)
      - Ex: Combining Liter with one of the prefix milli gives milliliter **(symbol mL=10^{-3}L) → (1 L = $(10cm)^3$ = $10^3 cm^3$, 1 mL = 1 cm^3)**
    - **SI units of other physical quantities:**
      - Area = d x d = m x m = m^2 (d = length)
      - Volume = d x d x d = m x m x m = m^3 (d=length)

Dimensional Analysis: Solving Numerical Problems

- **Dimensional analysis:** method of solving numerical problem using both the numerical value and the unit
  - Involves the use of conversion factors to change the units that a measured quantity is expressed in
  - In this method the identical units are multiplied or canceled
  - Ex: 1.0 cm x 2.1 cm = 2.1 cm^2
  - Ex: 10.4 ~~g~~ / 2.0 ~~g~~ = 5.2
- **Step 1:** Write the equations relating the units
  - **Ex:** (1.0 g = 10^3mg), to convert a mass from grams to milligrams, the equation must show the relationship between 1gram and the number of milligrams in 1 gram
  - **Ex:** (1.0 L = $10^3 \mu$L)

- **Step 2:** State the relationship as a faction (also called conversion factor) – 2 ways to be expressed:

$$\frac{1.0\ g}{10^3\ mg} \text{ and } \frac{10^3\ mg}{1.0\ g}$$

$$\frac{1.0\ L}{10^6\ \mu L} \text{ and } \frac{10^6\ \mu L}{1.0\ L}$$

- **Step 3:** Multiply the measured quantity by the conversion factor that cancels unwanted units and gives the final answer in required unit
  - A conversion factor changes the units of the measured quantity, but the actual quantity does not change
  - **Ex 1:** To convert from mg to g, the initial unit of quantity is the denominator of the conversion factor and the unit of quantity we are converting to (g) is in the numerator. By multiplying the measured quantity by this conversion factor, the units of the measured quantity change from mg to g.

Example (1)

$$\cancel{mg} \times \frac{g}{\cancel{mg}} = g \quad \text{and} \quad \cancel{g} \times \frac{mg}{\cancel{g}} = mg$$

Example (2)

$$\cancel{L} \times \frac{\mu L}{\cancel{L}} = \mu L \quad \text{and} \quad \cancel{\mu L} \times \frac{L}{\cancel{\mu L}} = L$$

- **Example 1:** How many Canadian Dollars would you have to pay to buy a souvenir in Delhi, costing 251.10 Indian Rupees
  - **Step 1: The equation relating the units:**
    - \$ 1.00 = 31.00 Rupees
  - **Step 2: The conversion factors are:**
    - $\frac{\$\ 1.00}{31.00\ \text{Rupees}}$ and $\frac{31\ \text{Rupees}}{\$1.00}$
  - **Step 3: Choose the conversion factor that cancels the units of Rupees and gives the answer in dollars (the conversion factor with units of dollars in the numerator):**
    - 251.10 Rupees x $\frac{\$\ 1.00}{31.00\ \text{Rupees}}$ = \$ 8.10
- **Example 2:** What is the mass of an object in milligrams that weighs 10.31 lbs?
  - **Step 1: The equation relating the units:**
    - 1.0 lb = 453.6 g
      1.00 g = 1.00 x 10^3mg
  - **Step 2: The conversion factors are:**
    - $\frac{1.00\ lb}{453.6\ g}$ and $\frac{453.6\ g}{1.00\ lb}$
      $\frac{1.00\ g}{1.00\ x\ 10^3 mg}$ and $\frac{1.00\ x\ 10^3 mg}{1.00\ g}$
  - **Step 3: Choose the conversion factor that cancels the units of pounds and gives the answer in mg**
    - 10.31 lb x $\frac{453.6\ g}{1.00\ lb}$ x $\frac{1.00\ x\ 10^3 mg}{1.00\ g}$ = 4.677 x 10^6mg

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Slides: scientific method: approach that scientists use to formulate and establish scientific laws includes a set of steps: Explaining the experimental observations, subsequent measurements in terms of a model, rectifying the model with additional experiments.
Note: although it is known as the scientific method, scientists do not use this practice alone. We all (through instinct) use and experience the same approach in our daily lives.

- Ex: in a standard visit to a medical clinic, tests are performed (**making observations**) that help the physician in diagnosing the probable cause of the problem (**proposing a hypothesis**). Next, the **medication is prescribed** and finally, the ailment is checked in due course to ensure that the problem has been solved (**testing the hypothesis by experiment**).

Key Terms

Law of Conservation of Mass: A law which states that "mass is neither created nor destroyed in chemical reactions". This essentially means that the same number of atoms of each type occurring on the reactant side must also appear on the product side.

Integer: A whole number (i.e., a number that contains no decimal or fractions).

Uncertainty: A characteristic of every measurement, indicating the error in a reported value. This is generally specified using significant figures.

Significant figures: A method for reporting the results of a measurement that includes all known or certain digits and a final digit that is considered uncertain.

SI unit: The International System of units that is based on the metric system and uses SI derived units for length, mass, time, etc. and prefixes to indicate magnitude (i.e., powers of 10).

Density: The physical property of a substance that corresponds to the ratio of its mass to volume. This property does not depend on the amount of substance.

Dimensional analysis: A method for solving numerical problems that uses 'conversion factors' to cancel out unwanted units and results in an answer with the desired units (i.e., dimensions).

Some SI Base Units

Physical Quantity	Unit	Symbol
Length	Meter	m
Mass	Kilogram	kg

Common SI Prefixes

Prefix	Multiple	Symbol
Giga	10^{9}	G
Mega	10^{6}	M
Kilo	10^{3}	k
Centi	10^{-2}	c
Milli	10^{-3}	m
Micro	10^{-6}	μ*
Nano	10^{-9}	n

*Greek letter pronounced as 'mew'

Module 3:

What is Matter?

Matter:

Describes things which

- Occupy space (hence possess volume and mass)
- Are perceivable by our senses

Can be classified in terms of its:

- Physical state: solid (wood), liquid (water) or gas (air)
- Chemical composition: element (copper), compound (sugar) or mixture (antifreeze)

Elements, Compounds and Mixtures:

Element: composed of one type of atom : classified as: ex: Pure carbon

- Metal
- Nonmetal
- Metalloid (semi-metal)

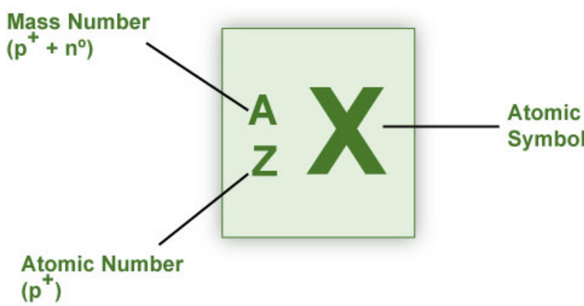
Compounds: combination of elements in a definite proportion.

- The atoms of each of the individual elements are chemically combined to form the compound
- The properties of the compound are different from those of the individual elements that is comprised of.
- A chemical change can break down a compound into its individual elements

Mixture: A non-pure substance made of 2 or more elements or compounds that can be separated by physical, as opposed to chemical, procedures

Atoms: composed of electrons, protons, and neutrons

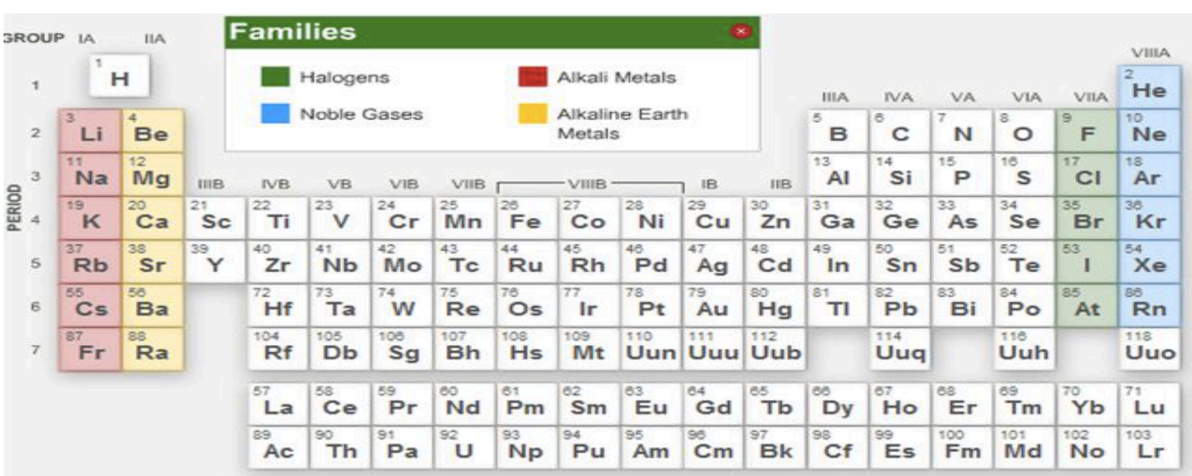
- The **nucleus** of the atom **contains :** the neutrons and protons
- The **electrons** surround the nucleus and are equal to the number or protons in a neutral atom
- **Protons:** have a positive charge (p^+)
- **Neutrons:** have no charge (n^o)
- **Electrons:** have a negative charge (e^-)
- **Atomic number** of an element **(Z)** is equal to :
  - The number of protons in the nucleus of its atoms
  - Each element has a different atomic number
- **Mass number (A)** is equal to:
  - The number of protons and neutrons in the atom
  - **Number of neutrons = Mass number (A) – Atomic Number (Z)**
- Every element has an **atomic symbol (X)**.
- Both the atomic and mass numbers are included with the atomic symbol



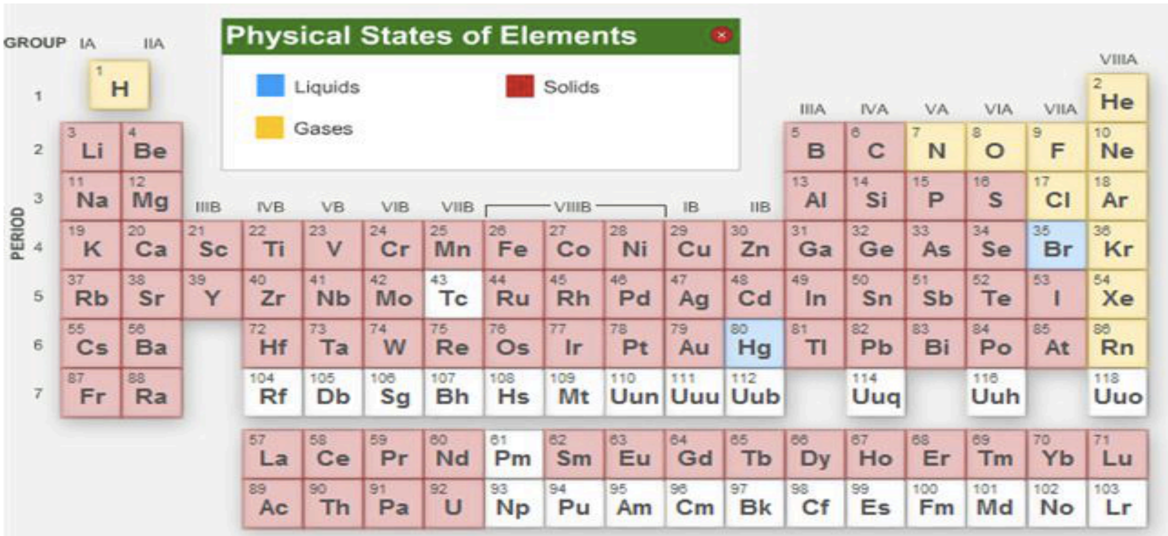
- **Molecules:** Combination of atoms in definite proportion (molecule of water) H_2O; hydrogen :oxygen= 2:1
- **Ions:** Are charged species formed by loss or gain of electrons from an atom
- **Cation:** Cation generation: M (g) → M^+(g) + e^- (loss of e^-)
    - M = metal, g = gaseous state
    - M^+, a positively charged ion is the cation
- **Anion:** Anion generation: X (g) + e^- → X^- (g) (gain of e^-)
    - X = nonmetal
    - X^-, a negatively charged ion is the anion
    - Ex: NaCl (table salt) a compound, consists of Na^+ (cation) and Cl^- (anion) ions

The Periodic Table: after the discovery of a large number of elements, it became necessary to find some way of arranging them in terms of their similarities. Many attempts were made by a number of scientists all over the world, and the result was a chart called the Periodic Table, proposed by the Russian scientis: Dmitri Mendeleev

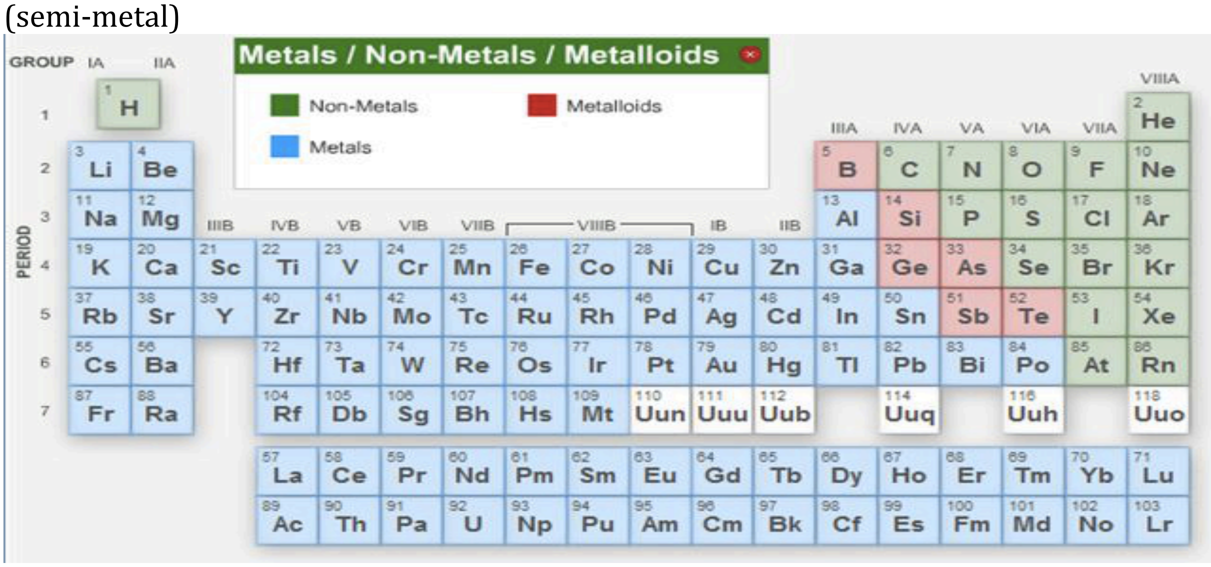
- **Elements:** are arranged into:
    - Vertical columns called **groups** or families, and horizontal rows called **periods**
    - Elements in the same group (or family) have similar chemical properties.
    - Elements in a period have different chemical properties
    - For each element, the atomic number is given at the top of the symbol and the atomic mass at the bottom
    - Period designated with numbers 1 to 7, contrary to groups, elements in a period have different properties
- **Groups:** are designated with the number 1 to 8, and a letter A or B.
    - The elements in the groups with the letter A designation are referred as the main group elements
    - Those with the letter B designation as the transition elements.
    - The inner transition elements (lanthanides and actinides) are located between groups 3B and 4B
    - Elements in the same group or family have similar chemical properties
    - Traditional families:
        - Alkali metals: group 1A
        - Alkaline earth metals: group 2A
        - Halogens: group 7A
        - Noble, rare, inert gases: 8A



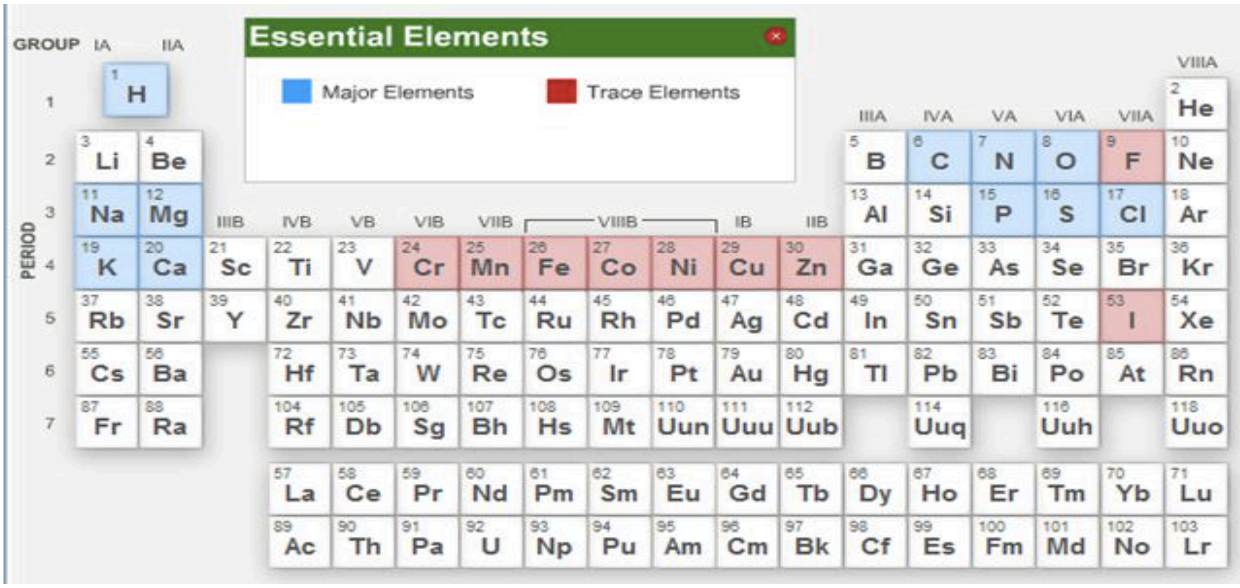
- **Physical states of elements include:**
  - **Solid :** metals are solids at (25°C) → with the exception of mercury (Hg) which is a liquid at (25°C).
    - Metals lose electrons to form cations and the nonmetals gain electrons to form anions
  - **Liquid**
  - **Gas**



- **Metals / Non – Metals / Metalloid:** Elements are classified as Metal, Non-metal and Metalloid (semi-metal)



- **Essential Elements:**
  - Major elements: most abundant in living systems and are essential for life
  - Trace elements: are found in trace amounts in the living system



Classifying Compounds
Chemical compounds are classified as:

- **Organic:** Compounds that contain at least one carbon atom
- **Inorganic:** Compounds that contain atoms other than carbon
  - **Exception:** Inorganic compounds such as: CARBON MONOXIDE (CO), CARBON DIOXIDE (CO_2)
  - For the purpose of **nomenclature** (system naming), inorganic compounds are classified as:
    - Containing a metal and nonmetal (ionic compounds)
    - Containing two nonmetals (covalent compounds)

Ionic compounds are composed of ions.

- They are formed by the electrostatic attraction between a positive ion (cation from the metal) and a negative ion (anion from the nonmetal).
- Electrons are transferred from the atoms of one element to the atoms of another element.
- In general, metals lose electrons and nonmetals gain electrons.
- Ionic compounds are charge neutral (zero net charge).
- Charges on ions involved in an ionic compound must be known in order to name the compound formed between them or to derive the formula for the compound.

Rules for naming and deriving formulas of ionic compounds:

- **Rule 1:** For metals of groups 1A, 2A, and 3A, <u>the charge on the monatomic cation is equal to the group number</u>
  - The cation has the same name as that of the metal and is always named first

Elements Symbol	Cation	Cation Name
Li	Li^+	Lithium
Mg	Mg^{2+}	Magnesium
Al	Al^{3+}	Aluminum

- **Rule 2:** For nonmetals of groups 5A,6A,7A, the charge on the formed monatomic anion is equal to the group number minus 8
    - The anion is comprised of the root of the nonmetal name and ends with the suffix-ide

Elements Symbol	Anion	Anion Name
N	N^{3-}	Nitride
O	O^{2-}	Oxide
F	F^{-}	Fluoride

Rule 3: Many metals (particularly transition metals) form more than one cation, thus forming more than one ionic compound with a particular anion.

- To indicate which ionic compound is formed, the charge on the metal ion is specified by using Roman numerals within parentheses following the name of the metal ion.
- Ex: Ion forms Fe^{2+} and Fe^{3+} ions. Thus iron can form 2 different ionic compounds with chlorine: $FeCl_2$ (iron(II) chloride) and $FeCl_3$ (iron (III) chloride)

Main group metal cations	Transition metal cations
Sn^{2+} and Sn^{4+}	Fe^{2+} and Fe^{3+}
Pb^{2+} and Pb^{4+}	Cu^{+} and Cu^{2+}
	Cr^{2+} and Cr^{3+}

- **Rule 4:** Ions that consist of two or more atoms bonded covalently possessing a net positive or negative charge are referred to as **polyatomic ions**
    - **Ammonium ion (NH_4^+) or Carbonate ion (CO_3^{2-})**
    - For more than 1 polyatomic ion, the ion is written in parentheses with a subscript indicating how many there are of the same polyatomic ion:
    - Calcium nitrate contains Ca^{2+} and 2 NO_3^- ions → $Ca(NO_3)_2$

- **Rule 5:** The name, formula and the charge for a number of commonly encountered polyatomic (ions with 2 or more atoms joined together) anions can be derived from common acids (H^+ donors) and bases (producing OH^-). When an acid dissociates (breaks up) it produces H^+ ions and a counter ion of equal but opposite charge.
    - In case of sulfuric acid (H_2SO_4) it can do so in 2 ways
        - Loses one H^+ ion leaving $HS0_4^-$ ion (notice 1 positive charge and 1 negative charge cancel out)
        - Lose 2 H^+ ion leaving the SO_4^{2-} ion (2 positive charges and 2 negative charges cancel out)

Common Polyatomic Ions

Name	Formula
Ammonium	NH_4^+
Bicarbonate	HCO_3^-
Carbonate	CO_3^{2-}
Hydroxide	OH^-
Nitrate	NO_3^-
Phosphate	PO_4^{3-}
Bisulfate	HSO_4^-
Sulfate	SO_4^{2-}

Example 1: $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$ [aq = aqueous]
(nitric acid) (nitrate ion)

Example 2: $H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$
(sulfuric acid) (hydrogen sulfate ion) (bisulfate ion)

$H_2SO_4(aq) \rightarrow 2H^+(aq) + SO_4^{2-}(aq)$
(sulfate ion)

Example 3: $H_2CO_3(aq) \rightarrow H^+(aq) + HCO_3^-(aq)$
(carbonic acid) (hydrogen carbonate ion) (bicarbonate ion)

$H_2CO_3(aq) \rightarrow 2H^+(aq) + CO_3^{2-}(aq)$
(carbonate ion)

Example 4: $H_3PO_4(aq) \rightarrow H^+(aq) + H_2PO_4^-(aq)$
(phosphoric acid) (dihydrogen phosphate ion)

$H_3PO_4(aq) \rightarrow 2H^+(aq) + HPO_4^{2-}(aq)$
(hydrogen phosphate ion)

$H_3PO_4(aq) \rightarrow 3H^+(aq) + PO_4^{3-}(aq)$
(phosphate ion)

Example 5: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
(sodium hydroxide) (hydroxide ion)

Rules for Naming Covalent Compounds

1. Name the first element in the formula list
2. Name the second element as an anion (-id)
3. Use prefixes to indicate the numbers of atoms of each element
4. Do not use the prefix mono- for the first element

Note: Some compounds are still known by their old names (and not with nomenclature)

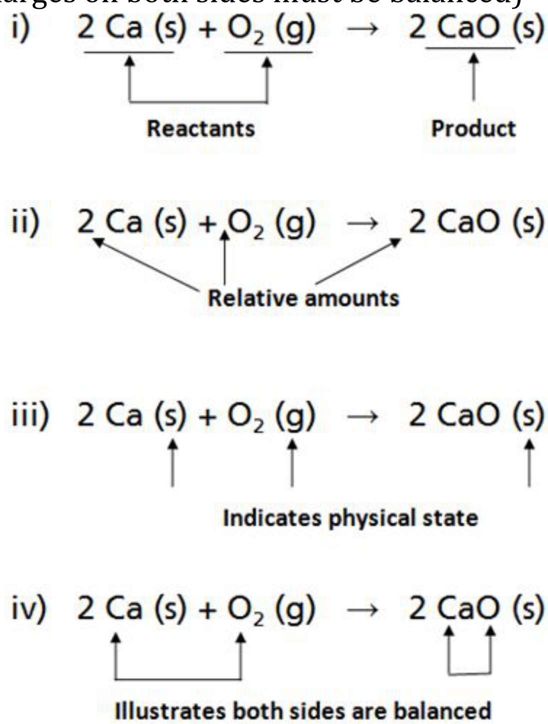
- Water (H_2O) [not dihydrogen monoxide]
- Ammonia (NH_3) [not nitrogen trihydride]
- Hydrogen sulfide (H_2S) [not dihydrogen monosulfide]
- Hydrogen peroxide (H_2O_2) [not dihydrogen dioxide]
- Methane (CH_4) [not carbon tetrahydride]
- Ozone (O_3) [not trioxygen]

Prefixes used in Naming Compounds:

Prefix	Number Indicated
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6

Representation of Chemical Reactions

- **Balancing:** Chemical reactions are expressed as an equation (chemical equation) using chemical formulas.
  - **Ex:** the reaction between calcium metal and oxygen gas is expressed as the balanced chemical equation: $2\ Ca\ (s) + O_2\ (g) \rightarrow 2\ CaO\ (s)$
    - **Step 1:** Identify the reactant (s) and product (s)
    - **Step 2:** Shows the relative amount of the reactants and products
    - **Step 3:** Indicate the physical state of the reactant (s) and product (s)
      - S: Solid
      - L : liquid
      - G: Gas
      - Aq: Aqueous
    - **Step 4:** The balanced equation complies with the law of conservation of mass and matter (atoms and charges on both sides must be balanced)



Writing Chemical Equations:

- Fractional coefficients are avoided
- Smallest whole-number coefficients are written
- **Coefficients:** number written in front of each formula $\rightarrow$ indicates the moles of reactants and products. It multiplies <u>every atom in the formula</u>
- **Subscripts:** shows the number of each atom present
- **Ex:** $2\ Al_2O_3$ represents 4 aluminum atoms and 6 oxygen atoms
- **Ex: $Pb(NO_3)_2\ (aq) + 2\ KI\ (aq) \rightarrow PbI_2\ (s) + 2\ KNO_3(aq)$**
  - **Coefficient:** 1,2,1,2
  - **Subscripts:**
    - Pb : 1, N : 1, O : 3
    - Bracket around (NO) and the subscript outside the bracket multiplies the subscript on both N and O by that amount.
    - The real subscript: N would 1 x 2 = 2, O 2 x 3 = 6
    - $2\ KNO_3$ represents 2 potassium atoms, 2 nitrogen atoms and 6 oxygen atoms

Representation of Chemical Reactions:

- A main point of a balanced equation is that there are the same numbers of atoms appearing in the reactants (left hand side) and the products (right hand side)
- Ex: $K + H_2O \rightarrow KOH + H_2$

  K:1 K:1
  H:2 H:3
  O:1 O:1

  - With 2 on the LHS and 3 on the RHS, H is not balanced
  - Note: K and H_2 are in elemental form (are not combined with any other element)
  - Multiplying each H_2O and KOH by 2

    $K + 2\ H_2O \rightarrow 2\ KOH + H_2$

    K:1 K:2
    H:4 H:4
    O:2 O:2
  - Both H and O are now balanced, but K is not → Multiply K on LHS by a factor to balance the overall equation

    $\mathbf{2\ K + 2\ H_2O \rightarrow 2\ KOH + H_2}$

    K:2 K:2
    H:4 H:4
    O:2 O:2

Using Balanced Equations for Chemical Calculations

- Balanced chemical equations are used for chemical calculations. As atoms, molecules or ions are the basic units, the coefficients in such equations denote their number
  - Ex: $N_2\ (g) + 3\ H_2\ (g) \rightarrow 2\ NH_3\ (g)$: 1 mole of nitrogen molecules consumes 3 moles of hydrogen molecules to yield (produce) 2 moles of ammonia molecules
- Under a set of optimum conditions, solid calcium carbonate decomposes to produce sold calcium oxide and gaseous carbon dioxide, how many moles of carbon dioxide gas will be produced (under similar conditions) from 5 moles of calcium carbonate
  - $CaCO_3\ (s) \rightarrow CaO(s) + CO_2\ (g)$
  - According to the balanced chemical equation → 1 mole of $CaCO_3$ produces 1 mole of CO_2 (g). Therefore 5 moles of carbon dioxide gas will be produced from 5 moles of solid calcium carbonate
- **Mole:** considered a "packaging unit" for atoms, molecules, ions etc.
  - **1 mole = 6.022×10^{23} Avogadro's number**
- **Atomic masses:** Mass of 1 mole of atoms or mass of 6.022×10^{23} atoms
  - Molecular mass (mass of 1 mole of molecules or mass of 6.022×10^{23} molecules) and Formula masses (term used for ionic compounds) can be calculated
    - **Molecular mass = sum of atomic masses**
    - **Molecular mass of SF_6 → (1 x 32.06 g/mol) + (6 x 19.00 g/mol) = 146.06 g/mol**
    - **Formula mass of $Ca(NO_3)_2$ → (1 x 40.08) + (2x14.01)+(6x16) = 164.10 g/mol**
  - **Mole =** $\dfrac{\textbf{Mass in grams}}{\textbf{Molar mass (or formula mass)}}$

Key Terms

Matter: Anything that possesses mass, occupies space, and is perceivable by our senses. It is generally classified in terms of a physical state and a chemical composition

Mixture: A non-pure substance made of two or more elements or compounds that can be separated by physical, as opposed to chemical, procedures. For example, a mixture of sodium chloride salt (an ionic compound) and water (a covalent compound) can be separated by removing the water through evaporation, leaving behind solid sodium chloride.

Elements: A substance composed of only one type of atom. Elements typically classified as being metal, nonmetal or semi-metal (metalloid).

Compound: substance composed of more than one atom. The elements are bound to one another in fixed numbers and can be broken down only through chemical reaction. The properties of a compound differ from those of the individual elements.

Atoms: The smallest unit of any element. It is composed of a positively charged nucleus, containing protons and neutrons, surrounded by a cloud of negatively charged electrons. The number of protons and electrons are identical, resulting in an overall neutral charge.

Electrons: A negatively charged sub-atomic particle (i.e., a particle that is smaller than an atom) that is found outside of the nucleus. The sharing or exchange of electrons between two different atom types is responsible for the formation of compounds.

Protons: A positively charged sub-atomic particle (i.e., a particle that is smaller than an atom) found in the nucleus of an atom.

Neutrons: An uncharged sub-atomic particle (i.e., a particle that is smaller than an atom) found in the nucleus of an atom.

Nucleus: It is the center of an atom where the protons and neutrons are located and which the electron(s) surround.

Atomic Number: A number that is unique to each type of atom and indicates the number of protons in the nucleus of each individual atom.

Mass number: Equal to the sum of the number of protons and neutrons in the nucleus of an atom.

Molecules: A discrete unit consisting of two or more atoms of the same or different elements bound together

Ions: An atom or molecule that has gained or lost electrons and is thus charged. Cations are positively charged (i.e., having lost electrons) while anions are negatively charged (i.e., having gained electrons).

Cation: An atom or molecule that has lost one or more electrons and thus possesses a positive charge. Such species are generated by the removal of electrons and are typically metals.

Anion: An atom or molecule that has gained one or more electrons and thus possesses a negative charge. Such species are generated by the capture of electrons and are typically non-metals.

Alkali metals: Any element belonging to Group IA (located in the left-most column) of the periodic table, with the exception of hydrogen. These Group IA metals all form monovalent cations (i.e., a singly charged cation).

Alkaline earth metals: Any element belonging to Group IIA of the periodic table. All of the Group IIA metals form divalent cations (i.e., cations with a charge of +2).

Halogens: Any element belonging to Group VIIA of the periodic table. Halogens are non-metals.

Noble gases: Any element belonging to Group VIIIA of the periodic table. Noble gases are also referred to as a rare or inert gases

Ionic Compounds: A compound consisting of a cation and an anion (generally a metal and a non-metal) where the bond is formed through the electrostatic interaction between the cation and the anion.

Polyatomic Ions: An ion (i.e., a species possessing a net overall charge) that is composed of two or more atoms that are held together through a covalent bond.

Acids: A chemical substance that produces protons (H^+ ions) when dissolved in water. The pH of the resulting solution is less than 7.

Bases: A substance that produces hydroxide ions (OH^-) when dissolved in water. The pH of the resulting solution is greater than 7.

Mole: A convenient unit of measure, used by chemists, indicating the number of items of a particular substance present. It is defined as the number of carbon atoms in 12 g (1 mole) of pure 12C and corresponds to 6.022x1023 (Avogadro's number) atoms.

Atomic Masses: The weighted average mass of the naturally occurring isotopes (i.e., an atom containing the same number of electrons and protons but differing in the number of neutrons) of an element. The weighting of each isotope mass in the weighted average mass correspond to the percentage of each isotope occurring in nature.