<table>
<thead>
<tr>
<th>Topic 2</th>
<th>Stoichiometry: Text: Chapters 1-3</th>
<th>(5 lectures)</th>
</tr>
</thead>
<tbody>
<tr>
<td>2.0</td>
<td>Classification of Materials: Mixtures, Elements and Compounds</td>
<td></td>
</tr>
<tr>
<td>2.1</td>
<td>Dalton’s Atomic Theory and Mass Law</td>
<td></td>
</tr>
<tr>
<td>2.2</td>
<td>Atoms, molecules and ions</td>
<td></td>
</tr>
<tr>
<td>2.3</td>
<td>Atoms and molecules, atomic and molecular masses</td>
<td></td>
</tr>
<tr>
<td>2.4</td>
<td>Isotopes and the Mass Spectrometer introduction</td>
<td></td>
</tr>
<tr>
<td>2.5</td>
<td>The mole</td>
<td></td>
</tr>
<tr>
<td>2.6</td>
<td>Relationships between FW's, mass and moles</td>
<td></td>
</tr>
<tr>
<td>2.7</td>
<td>Empirical and molecular formulae</td>
<td></td>
</tr>
<tr>
<td>2.8</td>
<td>Inorganic formulae and nomenclature (brief introduction)</td>
<td></td>
</tr>
<tr>
<td>2.9</td>
<td>Writing and balancing chemical equations</td>
<td></td>
</tr>
<tr>
<td>2.10</td>
<td>Mole and mass relationships from balanced chemical equations</td>
<td></td>
</tr>
<tr>
<td>2.11</td>
<td>Yields; percent yields; limiting reagents</td>
<td></td>
</tr>
<tr>
<td>2.12</td>
<td>Solutions; morality, volumetric analysis</td>
<td></td>
</tr>
<tr>
<td>2.13</td>
<td>Analysis of mixtures</td>
<td></td>
</tr>
</tbody>
</table>
Classification of Materials:

- **Matter** is anything that occupies space and has mass
- A **Substance** has a constant composition and distinct properties
- A **Mixture** is a combination of two or more substances that retain identities
  - Homogeneous - constant composition
  - Heterogeneous - composition not uniform
- Substances can be either **Elements** or **Compounds**
  - An **Element** cannot be separated into a simpler substance chemically
  - A **Compound** contains two or more elements united in fixed proportions
Three States of Matter:
All Substances can exist in 3 States, Solid, Liquid, and Gas that can be interconverted without changing the composition of the substance

- **Solid**
  - rigid
  - slight expansion on heating
  - slight compressibility

- **Liquid**
  - Flows and assumes container shape
  - slight expansion on heating
  - slight compressibility

- **Gas**
  - Fills container completely
  - Expansion infinitely
  - Easily compressibility

Solid $\rightarrow$ heat $\rightarrow$ Liquid $\rightarrow$ heat $\rightarrow$ Gas
Chemistry 121: Topic 2 - From Atoms to Stoichiometry

Physical and Chemical Properties:

- **Physical Property** measured and Observed *without* changing Composition
  - ie., Melting and Boiling point
  - Helium less dense than air (lighter than air)

- **Chemical Property** is measured or observed *with* a chemical change
  - Flammability
  - Reactivity

- An **Extensive Property** is a property that depends on how much matter is being examined ie., mass

- An **Intensive Property** is a property that does not depend on how much matter is being examined ie., density
Dalton’s Atomic Theory (1808):

- Elements are composed of extremely small particles called atoms. All atoms of a given element are identical, having the same size, mass, and chemical properties. The atoms of one element are different from the atoms of all other elements.

- Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.

- A chemical reaction involves only the separation, combination, or rearrangement of atoms; it does not result in their creation or destruction.

Essentially these statements are still correct
Dalton’s Atomic Theory (1808):

The first hypothesis:
- Atoms of one element are different from atoms of all other elements.
- Different properties explained by assuming atoms are not the same type.

The second hypothesis:
- To form a certain compound, need atoms of right elements, and proportion.
- Extension of a law published in 1799 by Joseph Proust.
- i.e., analyze samples of carbon dioxide gas different sources, find the same ratio by mass of carbon to oxygen.

- The law of multiple proportions. Two (or more) elements can combine to form more than one compound. Different compounds made up of the same elements differ in the number of atoms of each kind that combine.
Dalton’s Atomic Theory (1808):

The Third hypothesis:

- A restating of the law of conservation of mass, which is that matter can be neither created nor destroyed.

If we have 44 grams of Carbon Dioxide (CO₂) and we chemical change this material, to its elements or other compounds, we must still have 44 grams total.

If we have exactly 18 grams of water (H₂O) and convert all of this to its elements Hydrogen and Oxygen and nothing else. If we can demonstrate that we produced exactly 2 grams of Hydrogen then we can conclude that we produced 16 grams of Oxygen by the law of conservation of mass.
Atoms, molecules and ions:
From Dalton’s Theory, an Atom is the basic unit of an element that can enter into a chemical combination (i.e., Carbon, Oxygen and Hydrogen discussed earlier) However, it has subsequently been demonstrated that an Atom has a definite structure and is composed of smaller subatomic particles.

Electron
- very small particle \( (9.1 \times 10^{-28}\text{ grams}) \)
- fixed charge of \(-1.6022 \times 10^{-19}\text{ Coulombs} \) (-1 charge units)

Protons
- Larger than electron \( (1.7 \times 10^{-24}\text{ grams}) \)
- has a fixed charge of \(+1.6022 \times 10^{-19}\text{ Coulombs} \) (+1 charge units)

Neutrons
- Same mass as the Proton \( (1.7 \times 10^{-24}\text{ grams}) \)
- has no charge
Atomic Number, Mass Number, and Isotopes

- Atoms can be identified by the number of protons and neutrons they contain.

- The atomic number (Z) is the number of protons in the nucleus of each atom of an element.

- In a neutral atom the number of protons is equal to the number of electrons.

- The chemical identity of an atom can be determined solely from its atomic number. For example, the atomic number of nitrogen is 7. This means that each neutral nitrogen atom has 7 protons and 7 electrons. In other words, every atom that contains 7 protons is correctly named "nitrogen".

- The mass number (A) is the total number of neutrons and protons present in the nucleus of an atom of an element. Except for hydrogen, all atomic nuclei contain both protons and neutrons. In general the mass number is given by mass number = number of protons + number of neutrons.
Atomic Number, Mass Number, and Isotopes

- The number of neutrons in an atom is equal to the difference between the mass number and the atomic number, or \( A - Z \).
  - For example, the mass number of fluorine is 19 and the atomic number is 9 (indicating 9 protons in the nucleus). Thus the number of neutrons in an atom of fluorine is \( 19 - 9 = 10 \). The atomic number, number of neutrons, and mass number all must be positive integers.

- Atoms of a given element do not all have the same mass. Most elements have two or more isotopes, atoms that have the same atomic number but different mass numbers.
  - For example, there are three isotopes of hydrogen. One, simply known as hydrogen, has one proton and no neutrons. The deuterium isotope contains one proton and one neutron, and tritium has one proton and two neutrons.

\[
\begin{array}{ccc}
\text{mass #} & \to & A \\
\text{atomic #} & \to & Z \\
X & \equiv & 1 \ H \\
\text{hydrogen} & \equiv & 1 \ H \\
\text{deuterium} & \equiv & 2 \ H \\
\text{tritium} & \equiv & 3 \ H
\end{array}
\]
Atomic Number, Mass Number, and Isotopes
Determine the number of protons, neutrons and electrons in the following:

<table>
<thead>
<tr>
<th>Atomic Number</th>
<th>Mass Number</th>
<th>Elements</th>
</tr>
</thead>
<tbody>
<tr>
<td>17</td>
<td>199</td>
<td>Oxygen-17</td>
</tr>
<tr>
<td>8</td>
<td>80</td>
<td>Mercury-199</td>
</tr>
<tr>
<td>200</td>
<td>80</td>
<td>Mercury-200</td>
</tr>
<tr>
<td>200</td>
<td></td>
<td>Mercury-200 (cation)</td>
</tr>
</tbody>
</table>

- **Oxygen-17** (8p, 9n, 8e)
- **Mercury-199** (80p, 119n, 80e)
- **Mercury-200** (80p, 120n, 80e)
- **Mercury-200 (cation)** (80p, 120n, 78e)

> Isotopes have nearly identical chemical properties, but different mass
The Periodic Table: Elements are arranged by Atomic Number
Horizontal Rows called Periods: Vertical Columns are called groups or families
Classification of Elements:

- **Metal**: Good conductor of heat and electricity
- **Nonmetal**: Poor conductor of heat and electricity
- **Metalloid**: Properties intermediate between metal and nonmetal

![Periodic Table]

**FIGURE 2.10** The modern periodic table. The elements are arranged according to the atomic numbers above their symbols. With the exception of hydrogen (H), nonmetals appear at the far left of the table, while metals are located at the far right. Metalloids, which exhibit characteristics of both metals and nonmetals, are found in the central region of the table. The periodic table is a classification system for the chemical elements, listing them in order of increasing atomic number. It provides a framework for understanding the properties and behavior of elements, their compounds, and the chemistry of the universe.
### Group classification:

- **Group 1A: Alkali Metals**: (Li, Na, K, Rb, Cs, and Fr)
- **Group 2A: Alkaline Earth Metals**: (Be, Mg, Ca, Sr, Ba and Ra)
- **Group 7A: Halogens**: (F, Cl, Br, I, and At)
- **Group 8A: Nobel Gasses**: (He, Ne, Ar, Kr, and Rn)
Molecules and Ions:
A Molecule is a aggregate of at least two atoms in a definite arrangement held together by chemical forces (chemical bonds)

Diatomic Molecules contain two atoms, Polyatomic Molecules more than 2

- An ion is an atom or group of atoms with a net charge
  - Cation has a positive (+) charge
    - \( \text{Mg}^{+2}, \text{Fe}^{+3} \) Monatomic Cations
    - \( \text{NH}_4^+ \) Polyatomic Cation
  - Anion has a negative (-) charge
    - \( \text{Cl}^-, \text{S}^{-2} \) Monatomic Anions
    - \( \text{OH}^-, \text{CN}^- \) Polyatomic Anions

Common monatomic ions have a characteristic charge. Ions from similar “Groups” have a similar charge state. Why is that?
Empirical and molecular formulae:

**Molecular Formula:** Shows the *exact number* of atoms of each element in the smallest unit of a substance

- $H_2$, $H_2O$, $O_2$, $O_3$, $C_2H_6O$, $H_2O_2$

Molecular Formula’s can be represented in numerous ways;
Empirical and molecular formulae:
With Molecular formula’s you cannot tell connectivity
Structural Formula’s indicate connectivity (Note: mixed C₂H₅OH).

Empirical formula: describes the simplest whole-number ratio of the atoms in
a molecular formula ie., H₂O₂ has the empirical formula of HO, N₂H₄ is ?

What is the empirical formula of Caffeine, C₈H₁₀N₄O₂?
Atomic and molecular masses:

By international agreement, atomic mass (sometimes called atomic weight) is the mass of the atom in atomic mass units (amu). One atomic mass unit is defined as a mass exactly equal to one-twelfth the mass of one carbon-12 atom.

- Carbon-12 is the carbon isotope that has six protons and six neutrons.
- For Example; A hydrogen atom is 8.400 % as massive as a carbon-12 atom.

  \[
  \therefore \text{the atomic mass of hydrogen must be;}
  \]

  \[
  0.084 \times 12.00 \text{ amu} = 1.008 \text{ amu}.
  \]

Similar calculations show the atomic mass of oxygen is 16.00 amu.
Average Atomic Mass:

- Carbon has two naturally occurring isotopes carbon-12 with a natural abundance of 98.90% and carbon-13 with a natural abundance of 1.10%.
- The atomic mass of carbon-13 is 13.00335 amu. What is the average atomic mass of Carbon?

\[
(0.9890)(12.00000\text{ amu}) + (0.0110)(13.00335\text{ amu}) = 12.01\text{ amu}
\]

- Therefore, it is important to note that the atomic mass values reported in typical periodic tables is an average value.
- There is no carbon atom with an atomic mass of 12.01.
- Atomic mass units (amu) is a relative scale for the masses of the elements.
The Mole and Molar Mass:

- Chemists measure atoms and molecules in moles.

- In the SI system the mole (mol) is the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope. The actual number of atoms in 12 g of carbon-12 is determined experimentally.

- Avogadro’s number \((N_A)\), The currently accepted value is

\[
N_A = 6.0221367 \times 10^{23} \quad \text{(rounded to} \ 6.022 \times 10^{23})
\]

- Since 1 mole of carbon-12 atoms has a mass of exactly 12 grams and contains \(6.022 \times 10^{23}\) atoms. The molar mass \((M)\) of carbon-12 is exactly 12 grams.

- The molar mass of an element is numerically equal to its atomic mass in amu. i.e., The atomic mass of sodium (Na) is 22.99 amu, its molar mass is 22.99 g; the atomic mass of phosphorus is 30.97 amu, its molar mass is 30.97 g; and so on. The atomic mass of an element is numerically equal to its molar mass.
The Mole and Molar Mass:

**Calculation Example:** What is the mass in grams of a single carbon-12 atom.

- Carbon: \( \frac{12.00 \text{ grams}}{\text{mol}} : \frac{6.022 \times 10^{23} \text{ atoms}}{\text{mol}} \)

\[
\frac{12.00 \text{ grams}}{\text{mol}} = \frac{1.993 \times 10^{-23} \text{ grams}}{\text{atom}}
\]

\[
1.993 \times 10^{-23} \text{ grams} \times \frac{1 \text{ Carbon atom}}{12 \text{ amu}} = 1.661 \times 10^{-24} \text{ grams} \text{ amu}^{-1}
\]

- Inverse is equal to \( 6.022 \times 10^{23} \text{ amu/gram} \)
Molecular Mass:

- If we know the atomic masses, it is then a trivial matter to determine the Molecular Mass.

**Example:** What is the Molecular mass of H₂O? Caffeine (C₈H₁₀N₄O₂)?

**Answer:**

\[2(1.008) + 16.00 = 18.00 \text{ amu} = 18.00 \text{ grams/mol (the Molar Mass)}\]

\[8(12.01) + 10(1.008) + 4(14.01) + 2(16.00) = 194.20 \text{ amu} = 194.20 \text{ grams/mol}\]
**Molecular Mass:**

**Example:** How many Hydrogen atoms are present in 25.6 grams of Urea \((\text{NH}_2)_2\text{CO}\). First find the Molar Mass of Urea:

\[
2(14.01) + 4(1.008) + 12.01 + 16.00 = 60.06 \text{ amu} = 60.06 \text{ grams/mol}
\]

How many moles?

\[
\frac{25.6 \text{ grams}}{60.06 \text{ grams/mol}} = 0.426 \text{ mol}
\]

\[
0.426 \text{ mol} \times 6.022 \times 10^{23} \frac{\text{Molecules}}{\text{mol}} = 2.57 \times 10^{23} \text{ molecules}
\]

\[
(2.57 \times 10^{23} \text{ molecules}) \times 4 \text{ H atoms/molecule} = 1.03 \times 10^{24} \text{ H atoms}
\]
Propagation of Uncertainty (errors):

**Absolute Error:**
- The actual magnitude of the error
  - ie., for the reported value 200 ± 10, the Absolute error is ±10

- **When adding values add absolute error**
  - (18.6 ± 0.1) + (126.0 ± 0.5) = 144.6 ± 0.6
  - More exact error is the square root of the sum of the variances

**Relative Error (Percent Error):**
- The magnitude of the error expressed as a percentage of the value
  - ie., for the reported value 200 ± 10, the Relative error is ± 5%

- **When multiplying or dividing add relative error**
  - (126.0 ± 0.4%) ÷ (18.6 ± 0.5%) = 6.77 ± 0.9% = 6.77 ± 0.06
  - Powers i.e. $A^2$ or $A^3$ are simply multiple multiplications
Chemistry 121: Topic 2 - From Atoms to Stoichiometry

Problem Review: 12) A solution containing 72.7 mL of liquid A and 59.4 mL of liquid B, has a density of 1.024 g/mL. On adding a further 60.0 mL of liquid B to the mixture, the density increases to 1.1301 g/mL. Assuming that volumes are additive, calculate the densities of both liquids.

Initially, we have (know);
➢ 72.7 mL A + 59.4 mL B = 132.1 mL total solution
➢ 132.1 mL x 1.024 g/mL = 135.27 g

Add 60 mL of B
➢ 132.1 mL + 60 mL = 192.1 mL total solution
➢ 192.1 mL x 1.1301 g/mL = 217.09 g total mass

Determine mass of B added and Density of B
➢ 217.09 g - 135.27 g = 81.82181 g
➢ 81.82181 g ÷ 60 mL = 1.36369 g/mL Density of B

Determine mass of B added, Volume of A and Mass of A
➢ 59.4 mL + 60 mL = 119.4 mL B added in total
➢ 119.4 mL x 1.36369 g/mL = 162.83 g total mass B
➢ 217.09 g - 162.83 g = 54.26 g of A originally
➢ 54.26 g ÷ 72.7 mL = 0.7464 g/mL Density of A

Check:
(72.7 mL x 0.7464 g/mL) x (119.4 mL x 1.36369 g/mL) = 217.09 grams
Problem Review:

Bromine is a reddish-brown liquid. Calculate its density (in g/mL) if 586 g of the substance occupies 188 mL.

\[ 586 \text{ g} \div 188 \text{ mL} = 3.12 \text{ g/mL} \]

Mercury is the only metal that is a liquid at room temperature. Its density is 13.6 g/mL. How many grams of mercury will occupy a volume of 95.8 mL?

\[ 13.6 \text{ g/mL} \times 95.8 \text{ mL} = 1.30 \times 10^3 \text{ g} \]

Lithium is the least dense metal known (density: 0.53 g/cm$^3$). What is the volume occupied by 1.20 X $10^3$ g of lithium?

\[ 1.20 \times 10^3 \text{ g} \div 0.53 \text{ g/cm}^3 = 2.3 \times 10^3 \text{ cm}^3 \]
Problem Review:

The atomic masses of $^{35}\text{Cl}$ (75.53 percent) and $^{37}\text{Cl}$ (24.47 percent) are 34.968 amu and 36.956 amu respectively. Calculate the average atomic mass of chlorine. The percentages in parentheses denote the relative abundances.

$(34.968 \text{ amu} \times 0.7553) + (36.956 \times 0.2447) = 35.45 \text{ amu}$

Which of the following has more atoms: 1.10 g of hydrogen atoms or 14.7 g of chromium atoms?

For hydrogen: $1.10 \text{ g H} \times \frac{1 \text{ mol H}}{1.008 \text{ g H}} \times \frac{6.022 \times 10^{23} \text{ H atoms}}{1 \text{ mol H}} = 6.57 \times 10^{23} \text{ H atoms}$

For chromium: $14.7 \text{ g Cr} \times \frac{1 \text{ mol Cr}}{52.00 \text{ g Cr}} \times \frac{6.022 \times 10^{23} \text{ Cr atoms}}{1 \text{ mol Cr}} = 1.70 \times 10^{23} \text{ Cr atoms}$

There are more hydrogen atoms than chromium atoms.
How many molecules of ethane ($\text{C}_2\text{H}_6$) are present in 0.334 g of $\text{C}_2\text{H}_6$?

- $2 \times (12.01 \text{ g/mol}) + 6 \times (1.008 \text{ g/mol}) = 30.07 \text{ g/mol}$

- $0.334 \text{ g} \div 30.07 \text{ g/mol} = 1.11 \times 10^{-2} \text{ mol}$

- $(1.11 \times 10^{-2} \text{ mol}) \times (6.022 \times 10^{23} \text{ molecules/mol}) = 6.69 \times 10^{21} \text{ molecules}$
**Formulae for Inorganic Substances:**

- Ionic solids are held together by Ionic bonds (electrostatic attraction)
- No discrete molecular units
- ∴ can only report empirical formula

- Relative relationship of Cations and Anions indicated by their charge.
  - i.e., Na\(^+\) and Cl\(^-\) → NaCl ; Zn\(^{+2}\) and I\(^-\) → ZnI\(_2\)
  - What about Al\(^{+3}\) and O\(^{-2}\) → ?
    - Al\(_2\)O\(_3\)

- Recall that ions can also be polyatomic, i.e., NH\(_4\)\(^+\). Same rules apply
  - i.e., NH\(_4\)\(^+\) and Cl\(^-\) → NH\(_4\)Cl
  - Zn\(^{+2}\) and CrO\(_4\)\(^{-2}\) → ZnCrO\(_4\)
Nomenclature (Naming) of Ionic Substances:

Organic molecules contain Carbon and Hydrogen, Oxygen, Nitrogen or Sulfur (Some exceptions; CO, CO$_2$, CS$_2$, CN$^-$, CO$_3^{2-}$, HCO$_3^-$ etc.,)

All other compounds are considered **Inorganic compounds**

For the purpose of Inorganic Nomenclature we will consider 4 Inorganic types;

- Ionic Compounds
- Molecular Compounds
- Acids and Bases
- Hydrates

Metals are also Inorganic Compounds but hold either the elemental name or a common name (for mixtures of metals).
Ionic Compounds:

Ionic Compounds have both a **Cation** and an **Anion**.
Almost all Cations are derived from metal atoms (NH$_4^+$ is an exception).

**Metal Cations have names derived from their element:**

<table>
<thead>
<tr>
<th>Element</th>
<th>Cation</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Sodium (Na)</td>
<td>Na$^+$</td>
<td>Sodium ion (or Sodium Cation)</td>
</tr>
<tr>
<td>Magnesium (Mg)</td>
<td>Mg$^{+2}$</td>
<td>Magnesium ion (or Magnesium Cation)</td>
</tr>
</tbody>
</table>

Many Ionic Compounds are **binary compounds**, formed from 2 elements
ie., NaCl
In this case, the Anion is named by taking the first part of the element name,
removing the ending and adding “-ide”  (See Table 3.1 Text )

<table>
<thead>
<tr>
<th>Element</th>
<th>Anion</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>Chlorine (Cl)</td>
<td>Cl$^-$</td>
<td>Chloride ion</td>
</tr>
<tr>
<td>Bromine (Br)</td>
<td>Br$^-$</td>
<td>Bromide ion</td>
</tr>
</tbody>
</table>

What would a **ternary compound** be? consists of three elements
(See Table 3.3 Text)
Ionic Compounds:

Some metals can form more than one type of cation i.e.;

\[ \text{Fe}^{+2}, \text{Fe}^{+3} : \text{Mn}^{+2}, \text{Mn}^{+3}, \text{Mn}^{+4} \]

In an older naming system, the cation with less charge was identified by “-ous” whereas the larger charge ion was given the ending “-ic”

ie., Ferrous chloride = ? : Ferric chloride = ?

What about those elements with 3 or more? Use the Stock system (Roman numeral based system)

- \[ \text{MnO} \rightarrow \text{Manganese (II) Oxide} \quad \text{(Manganese-two oxide)} \]
- \[ \text{Mn}_2\text{O}_3 \rightarrow \text{Manganese (III) Oxide} \]
- \[ \text{MnO}_2 \rightarrow \text{Manganese (IV) Oxide} \]
Molecular Inorganic Compounds:

Molecules are discrete units. Inorganic molecules are usually composed of nonmetallic elements. (see periodic table figure 2.16 Text).

For binary molecules, name first element in formula first and then add “-ide” to the end of the second element.

- HCl = Hydrogen Chloride: SiC = Silicon Carbide

Use Greek Prefix to clarify compounds such as CO and CO₂
- CO = Carbon Monoxide: CO₂ = Carbon Dioxide

Greek Prefixes

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Number</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mono-</td>
<td>1</td>
</tr>
<tr>
<td>Di-</td>
<td>2</td>
</tr>
<tr>
<td>Tri-</td>
<td>3</td>
</tr>
<tr>
<td>Tetra-</td>
<td>4</td>
</tr>
<tr>
<td>Penta-</td>
<td>5</td>
</tr>
<tr>
<td>Hexa-</td>
<td>6</td>
</tr>
<tr>
<td>Hepta-</td>
<td>7</td>
</tr>
<tr>
<td>Octa-</td>
<td>8</td>
</tr>
<tr>
<td>Nona-</td>
<td>9</td>
</tr>
<tr>
<td>Deca-</td>
<td>10</td>
</tr>
</tbody>
</table>

2.8 Inorganic Nomenclature
Chemistry 121: Topic 2 - From Atoms to Stoichiometry

**Acids and Bases:**

An Acid is a substance that yields hydrogen ions (H\(^+\)) when dissolved in water.

Anions whose names end in “-ide” form acids with a “hydro-“ prefix and an “-ic” ending.

- HCl, Hydrogen Chloride as a molecule becomes hydrochloric acid

**Oxyacids** are acids that contain Hydrogen, oxygen and another element (which is the central element.

To write the molecular formula, H first, then central element, then Oxygen

ie., HNO\(_3\) → Nitric Acid; H\(_2\)CO\(_3\) → Carbonic Acid; H\(_2\)SO\(_4\) → Sulphuric Acid

Could have other acids with same elements but different number of Oxygen

Set of rules required to systematic name compounds
Acids and Bases:

An base can be described as a substance that yields hydroxide ions (OH\(^-\)) when dissolved in water.

These are often named as regular ionic substances;

- NaOH → Sodium Hydroxide
- KOH → Potassium Hydroxide

Note: A base is a substance that yields hydroxide ions (OH\(^-\)) when dissolved in water, not one that must contain hydroxide.

NH\(_3\) reacts with water to form NH\(_4^+\) and OH\(^-\) or stated another way;

\[ \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^- \] (chemical reaction)

Therefore, NH\(_3\) is a base by the definition given above.
**Hydrates:**

Hydrates are compounds that have a specific number of water molecules attached to them.

For example, in its normal state, each unit of copper(II) sulfate (CuSO₄) has five water molecules associated with it. CuSO₄• 5H₂O

CuSO₄• 5H₂O called copper(II) sulfate pentahydrate

The water molecules can be removed driven off by heating. The resulting compound is CuSO₄, which is sometimes called anhydrous copper(II) sulfate

- BaCl₂ • 2H₂O  barium chloride dihydrate
- LiCl • H₂O   lithium chloride monohydrate
- MgSO₄ • 7H₂O  magnesium sulfate heptahydrate

Many inorganic compounds have common names (see table 3.3 Text)
Writing and balancing chemical equations:

A chemical reaction is a process that changes one substance into one or more substances. Recall previous example with NH₃.

A chemical equation uses chemical symbols to show what happens in a chemical reaction.

Writing Chemical Equations:

When hydrogen gas (H₂) burns in air (which contains oxygen, O₂) to form water (H₂O).

\[ \text{H}_2 + \text{O}_2 \rightarrow \text{H}_2\text{O} \]

This symbolic expression can be read: "Molecular hydrogen reacts with molecular oxygen to yield water."

The reaction is assumed to proceed from left to right as the arrow indicates. This Equation is not complete because there are twice as many oxygen atoms on the left side of the arrow (two) its on the right side (one)
Writing and balancing chemical equations:

To conform to the law of conservation of mass, there must be the same number of each type of atom on both sides of the arrow (same # atoms before as after)

$$2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$$

This balanced chemical equation shows that "two hydrogen molecules can combine or react with one oxygen molecule to form two water molecules"

The equation can also be read as "2 moles of hydrogen molecules react with 1 mole of oxygen molecules to produce 2 moles of water molecules."

We know the mass of a mole of each of these substances, so we can also interpret the equation as "4.04 g of H$_2$ react with 32.00 g of O$_2$ to give 36.04 g of H$_2$O."
Example:
If you have 600mL of a 2:1, Hydrogen: Oxygen gas mixture. Determine how many moles of hydrogen and Oxygen you have and how many mL of water that would be produced if these combined completely. Assume STP and ideal gas conditions where 1 mole of gas = 22.4 L at STP.

Let \( x \) = volume of Oxygen: \( \therefore \) Volume of Hydrogen = 2x

\[ x + 2x = 600 \text{ mL} \]
\[ x = 200 \text{ mL} = \text{volume of Oxygen} \]
Volume of Hydrogen = 2x = 400 mL

for Oxygen:
\[ 0.200 \text{ L} \div 22.4 \text{ L/mol} = 0.008928 \text{ mol} = 8.93 \times 10^{-3} \text{ mol} \]

for Hydrogen
\[ 0.400 \text{ L} \div 22.4 \text{ L/mol} = 0.01786 \text{ mol} = 1.79 \times 10^{-2} \text{ mol} \]

\[ 1.79 \times 10^{-2} \div 8.93 \times 10^{-3} \text{ mol} = 1.9996 = 2.0 \]
Example cont.

Recall:

\[ 2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

From Oxygen:

1 mol of \( \text{O}_2 \) \( \rightarrow \) 2 mol of \( \text{H}_2\text{O} \)

\[ 8.93 \times 10^{-3} \text{ mol} \times 2 = 1.79 \times 10^{-2} \text{ mol} \text{H}_2\text{O} \] produced

Molar Mass \( \text{H}_2\text{O} = 2(1.008) + 16.00 = 18.02 \text{ g/mol} \)

\[ 18.02 \text{ g/mol} \times 1.79 \times 10^{-2} \text{ mol} = 0.322 \text{ grams H}_2\text{O} \]

Density of \( \text{H}_2\text{O} = 1.00 \text{ g/mL} \)

\[ \therefore 0.332 \text{ g} \div 1.00 \text{ g/mL} = 0.322 \text{ mL H}_2\text{O} \] produced

How many Liters of Hydrogen gas would you need to blow up to drown?
Example:

Aluminum is oxidized to produce $\text{Al}_2\text{O}_3$. How many moles of $\text{Al}_2\text{O}_3$ are produced when 0.50 grams of Aluminum are oxidized? How many grams of oxygen ($\text{O}_2$) are consumed?

$$\text{Al} + \text{O}_2 \rightarrow \text{Al}_2\text{O}_3$$

$$2\text{Al} + \frac{3}{2}\text{O}_2 \rightarrow \text{Al}_2\text{O}_3 : 4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$$

$\text{Al} = 0.50\text{g} \div 26.98\text{g/mol} = 0.0185\text{ mol}$

$$\frac{4}{0.185} = \frac{2}{x} \quad \Rightarrow \quad 4x = 2(0.185) : x = 0.0093\text{ mol} \text{ Al}_2\text{O}_3 \text{ Produced}$$

$$\frac{4}{0.185} = \frac{3}{x} \quad \Rightarrow \quad 4x = 3(0.185) : x = 0.14\text{ mol} \text{ O}_2 \text{ Consumed}$$

$0.14\text{ mol} \times 32.00\text{ g/mol} = 4.4\text{ g} \text{ O}_2 \text{ Consumed}$
Example:
A) How many moles of Hydrogen are produced when 1.0 g of Lithium is placed into exactly 100mL water. B) What is the mass of the system after the reaction is over if the hydrogen gas is allowed to escape.

\[
\begin{align*}
\text{Initial} & \quad 2 \text{ Li} + 2 \text{ H}_2\text{O} = 2 \text{ LiOH} + \text{ H}_2 \\
\text{mol react} & \quad 0.14 \text{ mol} \quad 0.14 \text{ mol} \\
\text{mol prod.} & \quad 0.14 \text{ mol} \quad 0.072 \text{ mol} \\
\text{end} & \quad 0.00 \text{ g} \quad 97.40 \text{ g} \\
\end{align*}
\]

A) \(7.2 \times 10^{-2}\) mol of Hydrogen gas produced.
B) \(101.00 \text{ g} - 0.15\text{g} = 100.85\ \text{g} \) remaining
Theoretical yield of the reaction is the amount of product that would result if all the limiting reagent reacted. The maximum obtainable yield

In practice, the actual yield, or the amount of product actually obtained from a reaction, is almost always less than the theoretical yield.

- Many reactions are reversible. Do not proceed 100 percent from left to right.
- Even if 100 percent complete, difficult to recover entire product
- The products formed may react further among themselves or with the reactants to form still other products.

The percent yield describes the proportion of actual yield to theoretical yield.

\[
\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%
\]

Therefore, % yield range from a fraction of 1 % to 100%
Example:

Hydrogen fluoride is used in the manufacture of Freons (which destroy ozone in the stratosphere) and in the production of aluminum metal. It is prepared by the reaction:

\[
\text{CaF}_2 + \text{H}_2\text{SO}_4 \rightarrow \text{CaSO}_4 + 2\text{HF}
\]

In one process 6.00 kg of CaF\(_2\) are treated with an excess of H\(_2\)SO\(_4\) and yield 2.86 kg of HF. Calculate the percent yield of HF.

\[
\begin{align*}
\text{CaF}_2 & \quad + \quad \text{H}_2\text{SO}_4 \quad \rightarrow \quad \text{CaSO}_4 \quad + \quad 2\text{HF} \\
78.08 \text{ g/mol} & \quad + \quad 98.08 \text{ g/mol} \quad \rightarrow \quad 136.14 \text{ g/mol} \quad + \quad 20.01 \text{ g/mol}
\end{align*}
\]

\[
6.00 \times 10^3 \text{ g} \div 78.08 \text{ g/mol} = 76.8 \text{ mol CaF}_2
\]

\[
\frac{1}{76.8} = \frac{2}{x}
\]

\[
\therefore 153.7 \text{ mol HF} \times 20.01 \text{ g/mol} = 3075.3 \text{ g} = 3.08 \text{ kg HF produced theoretically}
\]

\[
\% \text{ yield} = \frac{2.86 \text{ kg}}{3.08 \text{ kg}} \times 100\% = 92.9\%
\]
Determining Empirical Formula Experimentally

We can determine the empirical formula of a compound if we know the percent composition, therefore, we can identify compound formula experimentally.

- First, chemical analysis tells us the number of grams of each element present in a given amount of a compound.
- Convert the quantities in grams to number of moles of each element
- Normalize number of moles to the smallest value or, in other words, find the relative ratio’s of the moles.

Example:
11.5 grams of Ethanol is burned in an apparatus such as that shown in Figure 3.5. Carbon dioxide ($\text{CO}_2$) and water ($\text{H}_2\text{O}$) are given off. The masses of $\text{CO}_2$ and $\text{H}_2\text{O}$ determined by measuring the increase in mass of the $\text{CO}_2$ and $\text{H}_2\text{O}$ absorbers, was found to be 22.0 g and 13.5 g respectively. Determine the empirical formula of ethanol.

All carbon in $\text{CO}_2$ must come from ethanol and all hydrogen in $\text{H}_2\text{O}$ also.
Determining Empirical Formula Experimentally

Example (Solution):

22.0 g CO$_2$ $\div$ 44.01 g/mol = 0.4998 mol CO$_2$ $\therefore$ 0.4998 mol C
0.4998 mol C x 12.00 g/mol = 6.00 g Carbon in sample

13.5 g H$_2$O $\div$ 18.02 g/mol = 0.7491 mol H$_2$O $\therefore$ 1.498 mol H
1.498 mol H x 1.008 g/mol = 1.51 g Hydrogen in sample

11.5 g ethanol
-6.0 g Carbon
-1.51 g Hydrogen
4.0 g Oxygen $\div$ 16.00 g/mol = 0.25 mol O

$\therefore$ C$_{0.50}$H$_{1.5}$O$_{0.25}$ but need whole numbers

For C $\frac{0.50}{0.25} = 2$
For O $\frac{1.5}{0.25} = 6$

$\therefore$ C$_2$H$_6$O = Empirical Formula
Obtaining Mass % from a Molecular Formula

What is the mass % of C, H and O in $\text{C}_2\text{H}_6\text{O}$

1 mole of $\text{C}_2\text{H}_6\text{O}$ has a molecular mass of 46.05 g/mol

Contains 2 mole of Carbon = 24.00g

$$\frac{24.00}{46.05} = 0.521 = 52.1\%$$

Contains 6 mole of Hydrogen = 6.048 g

$$\frac{6.05}{46.05} = 0.131 = 13.1\%$$

Contains 1 mole of Oxygen = 16.00 g

$$\frac{16.00}{46.05} = 0.347 = 34.7\%$$

Be able to go from % values back to empirical formula
Solution Concentration:
Molarity is the most common and is defined as the number of moles per liter or:

\[
\text{Molarity (M)} = \frac{\text{Moles Solute}}{\text{Liters of Solution}}
\]

Example:
What is the concentration of a solution containing 2.0 grams of glucose (C\textsubscript{6}H\textsubscript{12}O\textsubscript{6}) in 250 mL of water?

\[
\text{C}_6\text{H}_{12}\text{O}_6 \rightarrow \text{molecular weight} = 6(12.00) + 12(1.008) + 6(16.00) = 180.10 \text{ g/mol}
\]

\[
2.0 \text{ g} \div 180.10 \text{ g/mol} = 0.011 \text{ mol}
\]

\[
\text{Molarity (M)} = \frac{0.011 \text{ mol}}{0.250 \text{ L}} = 0.044 \text{ mol/L} = 0.044 \text{ M} = [\text{C}_6\text{H}_{12}\text{O}_6]
\]

How many grams of glucose are in 50 ml of this solution?

\[
0.044 \text{ mol/L} \times 0.050 \text{ L} = 0.0022 \text{ mol} \times 180.10 \text{ g/mol} = 0.40 \text{ grams glucose}
\]
Solution Concentration:
When salts are dissolved the solid crystal structure is lost and the cations and anions are completely separated. i.e.,

\[ \text{NaCl}_\text{(s)} + \text{H}_2\text{O}_\text{(l)} \rightarrow \text{Na}^+\text{(aq)} + \text{Cl}^-\text{(aq)} \]

similarly: \[ \text{MgCl}_2\text{(s)} + \text{H}_2\text{O}_\text{(l)} \rightarrow \text{Mg}^{+2}\text{(aq)} + 2\text{Cl}^-\text{(aq)} \]

What is the \([\text{NO}_3^-]\) that results when 4.00 grams of \(\text{Ba(NO}_3)_2\) are dissolved in 500mL of deionized water?

\(\text{Ba(NO}_3)_2 \rightarrow \text{mol weight} = 137.327 + 2(14.01) + 6(16.00) = 261.35 \text{ g/mol}\)

4.00 g ÷ 261.35 g/mol = 0.0153 mol \(\text{Ba(NO}_3)_2\) = 0.0306 mol \(\text{NO}_3^-\)

\(\text{Ba(NO}_3)_2 + \text{H}_2\text{O}_\text{(l)} \rightarrow \text{Ba}^{+2}\text{(aq)} + 2\text{NO}_3^-\text{(aq)}\)

Molarity (M) = \[
\frac{0.0306 \text{ mol}}{0.500 \text{ L}} = 0.061 \text{ mol/L} = 0.061 \text{ M} = [\text{NO}_3^-]
\]

(see Text, Figure 4.6 Re: Making solutions of Known Concentration)
Dilution of Solutions:

**Dilution** is the procedure for preparing a less concentrated solution from a more concentrated one. A **Stock solution** is a more concentrated solution used to make “working” solutions.

Outline how to make a 500 mL solution of 0.400 M KMnO$_4$ from a stock solution of 1.50 M KMnO$_4$.

1) Determine how many moles are required,

\[ 0.400 \text{ mol/L} \times 0.500 \text{ L} = 0.200 \text{ mol} \]

2) What volume of the stock solution contains 0.200 mol?

\[
\frac{1.50 \text{ mol}}{1.00 \text{ L}} = \frac{0.200 \text{ mol}}{X \text{ L}}
\]

\[ 1.50 \times 0.200 = X \times 0.133 \text{ L} = 133 \text{ mL} \]

3) Put 133 mL of Stock solution into a 500 mL Volumetric Flask and then fill to mark with deionized water.
Gravimetric Analysis:
An analytical technique based on the measurement of mass. In one type, a specific insoluble compound is formed, isolated and weighed.

For example:

\[ \text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{AgCl}(s) \]

Net Reaction:

\[ \text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(s) \]

Could experimentally determine percent of Cl\(^-\) in NaCl by weighing an amount of NaCl, dissolving it and adding an access of AgNO\(_3\). The AgCl(s) precipitate could then be filtered, dried and weighed. From the mass of AgCl obtained we could determine the mass of Cl\(^-\) in solution and thus the mass of Cl in the original NaCl sample.
Example: 0.5662 of an unknown ionic compound containing chloride ions and an unknown metal is dissolved in water. This is treated with an excess of AgNO₃ and 1.0882 g of AgCl precipitates. What is the percentage by mass of Chloride in the original sample?

\[
\text{1.0882 g} \div 143.4 \text{ g/mol} = 7.589 \times 10^{-3} \text{ mol} \times 35.45 \text{ g/mol} = 0.2690 \text{ grams Cl}
\]

\[
\% \text{ Cl} = \frac{0.2690 \text{ g}}{0.5662 \text{ g}} \times 100\% = 47.51\%
\]
Acid-Base Titration:

In titration, a solution of accurately known concentration, called a standard solution, is added gradually to another solution of unknown concentration, until the chemical reaction between the two solutions is complete.

- **The equivalence point:** The point at which the acid (or base) has completely reacted with or been neutralized by the base (or acid). The equivalence point is signaled by a sharp change in the color of an indicator in the acid solution.

- **Indicators** are substances that have distinctly different colors in acidic and basic media.

**Example:** If 16.1 mL of 0.610 M NaOH is needed to neutralize 20.0 ml of H₂SO₄ solution. What is the concentration of the H₂SO₄ solution?

\[
2\text{NaOH}_{(aq)} + \text{H}_2\text{SO}_4\ (aq) \rightarrow \text{Na}_2\text{SO}_4\ (aq) + 2\text{H}_2\text{O}\ (l)
\]

Moles of NaOH = 0.610 mol/L x .0161 L = 0.009821 mol NaOH

\[
\frac{2}{9.82 \times 10^{-3}} = \frac{1}{x}
\]

\[
x = 4.91 \times 10^{-3} \text{ mol H}_2\text{SO}_4 \text{ in 0.020 L}
\]

\[
\therefore 4.91 \times 10^{-3} \text{ mol} \div 0.020 \text{ L} = 0.246 \text{ M H}_2\text{SO}_4
\]
Redox Titrations:
Redox reactions involve the transfer of electrons. An oxidizing agent can be titrated against a reducing agent in a similar manner to an acid/base titration.

➤ The **equivalence point**: The point at which the reducing agent has been completely oxidized by the oxidizing agent.

**Example**: If 16.42 mL volume of 0.1327 M KMnO₄ solution is needed to oxidize 20.0 mL of FeSO₄ solution in an acidic medium. What is the concentration of the FeSO₄ solution? The net ionic equation is;

\[
5\text{Fe}^{2+} \text{(aq)} + \text{MnO}_4^- \text{(aq)} + 8 \text{H}^+ \rightarrow \text{Mn}^{2+} \text{(aq)} + 5\text{Fe}^{3+} \text{(aq)} + 4\text{H}_2\text{O (l)}
\]

Moles of KMnO₄ = 0.1327 mol/L x 0.01642 L = 0.002179 mol NaOH

\[
\frac{1}{2.179 \times 10^{-3}} = \frac{5}{x}
\]

∴ \( x = 1.089 \times 10^{-2} \) mol FeSO₄ in 0.020 L

\[
1.089 \times 10^{-2} \text{ mol} \div 0.020 \text{ L} = 0.5445 \text{ M FeSO₄}
\]
Problem 3.121:
A mixture of CuSO₄•5H₂O and MgSO₄•7H₂O is heated until all the water is lost. If 5.020 g of the mixture gives 2.988 g of the anhydrous salts, what is the percent by mass of CuSO₄•5H₂O in the mixture?

Problem 3.123:
A mixture of methane (CH₄) and ethane (C₂H₆) of mass 13.43 g is completely burned in oxygen. If the total mass of CO₂ and H₂O produced is 64.84 g, calculate the fraction of CH₄ in the mixture.
Problem 3.121:
A mixture of CuSO₄•5H₂O and MgSO₄•7H₂O is heated until all the water is lost. If 5.020 g of the mixture gives 2.988 g of the anhydrous salts, what is the percent by mass of CuSO₄•5H₂O in the mixture?

Mol. Wt. CuSO₄•5H₂O = 249.70 g /mol; Mol. Wt. MgSO₄•7H₂O = 246.49 g /mol

Let x = grams of CuSO₄•5H₂O
∴ (5.020 - x) grams of MgSO₄•7H₂O

\[
x \text{ g} \div 249.70 \text{ g/mol} = \frac{x}{249.70} \text{ mol CuSO₄•5H₂O}
\]

\[
\text{CuSO₄•5H₂O} \rightarrow \text{CuSO₄} + 5\text{H₂O}
\]

∴ \[
\frac{5x}{249.70} \text{ mol H₂O produced from } \frac{x}{249.70} \text{ CuSO₄•5H₂O}
\]

\[
\frac{5x}{249.70} \text{ mol} = 0.0200X \text{ mol H}_2\text{O} \times 18.02 \text{ g/mol} = 0.360X \text{ g H}_2\text{O}
\]
(5.020 - x) grams of MgSO₄•7H₂O

\[
\frac{(5.020-x) \text{ g}}{246.49 \text{ g/mol}} = \text{ mol MgSO₄•7H₂O}
\]

(0.02037 - 0.004057X) mol MgSO₄•7H₂O

\[
\text{MgSO₄•7H₂O} \rightarrow \text{MgSO₄} + 7\text{H₂O}
\]

∴ 7 (0.02037 - 0.004057X) mol H₂O produced = (0.14259 - 0.02840X) mol

(0.14259 - 0.02840X) mol H₂O x 18.02 g/mol = (2.569 - 0.5118X) g H₂O

Total H₂O = 0.360X g + (2.569 - 0.5118X) g = 5.020 g - 2.988 g = 2.032 g

2.569 - 0.1518X = 2.032; X = -0.537 ÷ -0.1518 = 3.5375 g CuSO₄•5H₂O

3.5375 g ÷ 5.020 X 100% = 70.47 %
Problem 3.123:
A mixture of methane (CH₄) and ethane (C₂H₆) of mass 13.43 g is completely burned in oxygen. If the total mass of CO₂ and H₂O produced is 64.84 g, calculate the fraction of CH₄ in the mixture.

A) \[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]  
B) \[ 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \]

Let \( X \) = Mass of CH₄ \quad \therefore \text{Mass of C}_2\text{H}_6 = (13.43 - X) \)

\[
x \text{ g} \div 16.03 \text{ g/mol} = \frac{X}{16.03} \text{ mol CH}_4 = 0.06238 \times X \text{ mol CH}_4
\]

\( \therefore 0.06238 \times X \text{ mol CO}_2 \text{ from CH}_4 \text{ and } 0.1248 \times X \text{ mol H}_2\text{O from CH}_4\)

\[
(0.06238 \times X \text{ mol CO}_2) \times 44.00 \text{ g/mol} = 2.745 \times X \text{ g CO}_2 \text{ from CH}_4
\]

\[
(0.1248 \times X \text{ mol H}_2\text{O}) \times 18.02 \text{ g/mol} = 2.249 \times X \text{ g H}_2\text{O from CH}_4
\]
Mass of C₂H₆ = (13.43 - X)

\[(13.43 - X) \text{ g} / \text{mol} = \frac{(13.43 - X)}{30.04} \text{ mol} \]

\(\text{mol C₂H₆} = (0.4471 - 0.03329X) \text{ mol}\)

\[2C₂H₆ + 7O₂ \rightarrow 4CO₂ + 6H₂O\]

∴ (0.8942 - 0.06658X) mol CO₂ from C₂H₆

\[((0.8942 - 0.06658X) \text{ mol CO₂}) \times 44.00 \text{ g/mol} = (39.34 - 2.930X) \text{ g CO}_2 \text{ from C}_2\text{H}_6\]

\[((1.341 - 0.09987X) \text{ mol H}_2\text{O}) \times 18.02 \text{ g/mol} = (24.16 - 1.800X) \text{ g H}_2\text{O from C}_2\text{H}_6\]

\[
\begin{align*}
(39.34 - 2.930X) & \quad \text{g CO}_2 \\
2.745X & \quad \text{g CO}_2 \\
(24.16 - 1.800X) & \quad \text{g H}_2\text{O} \\
2.249X & \quad \text{g H}_2\text{O} \\
64.84 & \quad \text{g H}_2\text{O and CO}_2 \text{ total}
\end{align*}
\]

\[
(39.34 - 2.930X) + 2.745X + (24.16 - 1.800X) + 2.249X = 64.84
\]

\[63.50 + 0.264X = 64.84 \quad \therefore X = 5.08 \text{ g CH}_4 \quad \therefore 5.08 \div 13.43 = 0.38\]
Problem 1:

A) \( \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \)  

B) \( 2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O} \)

Let \( X = \text{Mass of CH}_4 \)  
\[ \therefore \text{Mass of C}_2\text{H}_6 = (13.43 - X) \]

\[ x \text{ g} \div 16.03 \text{ g/mol} = \frac{X}{16.03} \text{ mol CH}_4 = 0.06238X \text{ mol CH}_4 \]

\[ \therefore 0.06238X \text{ mol CO}_2 \text{ from CH}_4 \text{ and } 0.1248X \text{ mol H}_2\text{O} \text{ from CH}_4 \]

\( (0.06238X \text{ mol CO}_2) \times 44.00 \text{ g/mol} = 2.745X \text{ g CO}_2 \text{ from CH}_4 \)

\( (0.1248X \text{ mol H}_2\text{O}) \times 18.02 \text{ g/mol} = 2.249X \text{ g H}_2\text{O} \text{ from CH}_4 \)
Mass of C$_2$H$_6$ = (13.43 - X) g

\[
(13.43 - X) \text{ g} \div 30.04 \text{ g/mol} = \frac{(13.43 - X)}{30.04} \text{ mol C}_2\text{H}_6 = (0.4471 - 0.03329X) \text{ mol}
\]

\[2\text{C}_2\text{H}_6 + 7\text{O}_2 \rightarrow 4\text{CO}_2 + 6\text{H}_2\text{O}\]

\[
\therefore (0.8942 - 0.06658X) \text{ mol CO}_2 \text{ from C}_2\text{H}_6
\]

\[
((0.8942 - 0.06658X) \text{ mol CO}_2) \times 44.00 \text{ g/mol} = (39.34 - 2.930X) \text{ g CO}_2 \text{ from C}_2\text{H}_6
\]

\[
((1.341 - 0.09987X) \text{ mol H}_2\text{O}) \times 18.02 \text{ g/mol} = (24.16 - 1.800X) \text{ g H}_2\text{O} \text{ from C}_2\text{H}_6
\]

\[
\begin{align*}
(39.34 - 2.930X) & \text{ g CO}_2 \\
2.745X & \text{ g CO}_2 \\
(24.16 - 1.800X) & \text{ g H}_2\text{O} \\
+ 2.249X & \text{ g H}_2\text{O} \\
64.84 & \text{ g H}_2\text{O and CO}_2 \text{ total}
\end{align*}
\]

\[
(39.34 - 2.930X) + 2.745X + (24.16 - 1.800X) + 2.249X = 64.84
\]

\[
63.50 + 0.264X = 64.84 \quad \therefore X = 5.08 \text{ g CH}_4 \quad \therefore \frac{5.08}{13.43} = 0.38
\]