I. Stoichiometry

A) Laws of Chemical Combination

A *law* is a statement that describes the general behavior of nature.

Law of Conservation of Mass: The mass of the products is equal to the mass of the reactants.

Law of Definite Proportion: A given compound, however it is obtained, contains its component elements in a fixed proportion by weight.

Example: For any given weight (**wt**) of water (H_2O)

$$\frac{wt\ of\ H\ in\ the\ given\ wt\ of\ H_2O}{wt\ of\ O\ in\ the\ given\ wt\ of\ H_2O} = 0.126$$

Law of Multiple Proportions: When two elements combine to form more than one compound, the different weights of one element that combined with a fixed weight of the other element are in a ratio of a small whole number.

Example: 16 g of oxygen combine with 2 g of hydrogen to form water (H_2O) and 32 g of oxygen combine with 2 g of hydrogen to form hydrogen peroxide (H_2O_2). For the same fixed weight (2 g) of hydrogen

$$\frac{wt\ of\ O\ in\ H_2O_2}{wt\ of\ O\ in\ H_2O} = \frac{32g}{16g} = 2$$

B) Atomic Theory - John Dalton 1803

A *theory* is an explanation of the observed behavior (law) in terms of a model.

Postulates - underlying hypotheses or assumptions - **of the Atomic Theory**

1: A chemical element is composed of very small, discrete particles called *atoms* that remain unchanged during a chemical reaction.

2: All atoms of the same element are alike in all respects (thought to have the same mass) and atoms of different elements have different masses. Today we known that atoms of the same element may have different masses but these atoms will have the same number of protons. Thus, atoms of different elements will differ in the number of their protons.

3: Chemical compounds consist of *molecules* which are a stable grouping of an integral number of atoms of each component element.

Example: A water molecules contains two hydrogen atoms and one oxygen atom. The chemical formula H_2O for water indicates this composition, i.e. subscript states the number of atoms of that element in the molecule.

Explanation of the Laws of Chemical Combination

Law of Conservation of Mass: Since atoms are conserved in a chemical reaction (Postulate 1), the number of atoms of each element in the products will equal the number of atoms of each element in the reactants.

$$H_2 + \tfrac{1}{2} O_2 \rightarrow H_2O$$

Law of Definite Proportion: Consider a compound that contains hydrogen (H) and oxygen (O). A molecule of the compound contains x hydrogen atoms and y oxygen atoms, i.e. H_xO_y and x and y are whole numbers (Postulate 3). Each hydrogen atom has a mass h and each oxygen atom has a mass o (Postulate 2). If a given weight of the compound H_xO_y contains N molecules, then

$$\frac{wt\ of\ H\ in\ the\ given\ wt\ of\ H_xO_y}{wt\ of\ O\ in\ the\ given\ wt\ of\ H_xO_y} = \frac{Nxh}{Nyo} = \frac{xh}{yo} = constant$$

Law of Multiple Proportions: Consider the compounds H_xO_y and H_wO_z. x, y, w, and z are whole numbers (Postulate 3). Each hydrogen atom has a mass h and each oxygen atom has a mass o (Postulate 2). If a fixed weight of hydrogen contains N hydrogen atoms, then the number of oxygen atoms needed to form H_xO_y is given by the equation

$$Number\ of\ H_xO_y\ molecules\ that\ can\ be\ formed\ from\ N\ hydrogen\ atoms = \frac{N}{x}$$

$$Number\ of\ oxygen\ atoms\ in\ \tfrac{N}{x}\ molecules\ of\ H_xO_y = (y)\frac{N}{x} = \frac{yN}{x}$$

and the number of oxygen atoms needed to form H_wO_z is

$$Number\ of\ oxygen\ atoms\ in\ \tfrac{N}{w}\ molecules\ of\ H_wO_z = (z)\frac{N}{w} = \frac{zN}{w}$$

The weight of oxygen needed to combine with the given weight of hydrogen to form H_xO_y is equal to the mass of one oxygen atom times the number of oxygen atoms needed , i.e. (o)(yN/x) and the weight of oxygen needed to combine with the same given weight of hydrogen to form H_wO_z is (o)(zN\w). For the same fixed weight of hydrogen, the ratio of the weight of oxygen needed to form H_xO_y to the weight of oxygen needed to form H_wO_z is

$$\frac{wt\ of\ O\ in\ H_xO_y}{wt\ of\ O\ in\ H_wO_z} = \frac{o\left(\frac{Ny}{x}\right)}{o\left(\frac{Nz}{w}\right)} = \frac{yw}{xz}$$

C) Law of Combining Volumes

Gaseous volumes measured at the same temperature (temp) and pressure

1 volume of hydrogen + 1 volume of chlorine ➜ 2 volumes of hydrogen chloride

2 volumes of hydrogen + 1 volume of oxygen ➜ 2 volumes of water

Law of Combining Volumes: When gases react, the volumes produced or consumed , measured at the same temperature and pressure are in the ratio of small whole numbers.

Explanation proposed by Avogadro

Avogadro's Hypothesis: Equal volumes of gases, measured at the same temperature and pressure, contain the same number of **molecules**.

Implication: At the same temperature and pressure

1 volume of hydrogen + 1 volume of chlorine → 2 volumes of hydrogen chloride

Let N = the number of molecules in one volume at the temperature and pressure

N molecules of hydrogen + N molecules of chlorine → 2N molecules of hydrogen chloride

Divide by N.

1 molecule of hydrogen + 1 molecule of chlorine → 2 molecules of hydrogen chloride

The chemical formula for hydrogen chloride has been experimentally determined to be HCl. Inserting the chemical formulas for hydrogen, chlorine, and hydrogen chloride

$$H_x + Cl_y \rightarrow 2\ HCl$$

The Law of Conservation of Mass demands that x and y are 2.

$$H_2 + Cl_2 \rightarrow 2\ HCl$$

Some elements that exist as diatomic molecules

H_2, N_2, O_2, Cl_2, Br_2

D) The Mole Concept

According to the Dalton's Atomic Theory atoms of one element or molecules of one compound can be distinguish from atoms of other elements or molecules of other compounds by their *absolute* masses. Since the absolute mass of an atom or molecule could not be measured in the early part of the nineteenth century, a set of *relative* masses were established to characterize elements and compounds. The relative mass of element A (RM_A) is defined with respect to the relative mass of a reference element (RM_{ref}).

$$\frac{RM_A}{RM_{ref}} = \frac{Absolute\ mass\ of\ an\ atom\ of\ element\ A}{Absolute\ mass\ of\ an\ atom\ of\ reference\ element}$$

Example: When 0.600 g of carbon are burned in pure oxygen, 2.20 g of carbon dioxide are formed. The chemical formula for carbon dioxide is CO_2.

$$\frac{RM_O}{RM_C} = \frac{o}{c} = \frac{Absolute\ mass\ of\ an\ atom\ of\ oxygen}{Absolute\ mass\ of\ an\ atom\ of\ carbon} \qquad (1)$$

$$c = \frac{0.600\,g}{N}$$

$$where\ N\ is\ the\ number\ of\ carbon\ atoms\ in\ the\ 2.20\,g\ of\ CO_2$$

$$o = \frac{2.20\,g - 0.600\,g}{2N} = \frac{1.60\,g}{2N}$$

If the value of RM_C is arbitrarily set at 12.00, then

$$\frac{RM_O}{RM_C} = \frac{o}{c} = \frac{\frac{1.60\,g}{2N}}{\frac{0.600\,g}{N}} = \frac{0.800\,g}{0.600\,g}$$

$$RM_O = RM_C\,\frac{0.800\,g}{0.600\,g} = (12.00)\,\frac{0.800\,g}{0.600\,g} = 16.00$$

In a similar manner the relative masses of other elements and compounds were determined.

Substance	**Relative Mass**
H	1.00
B	10.81
H_2	2.00
CO_2	44.00
H_2O	18.00

Note: The relative mass of each substance is characteristic of the substance in the same way that the absolute mass is characteristic of the substance.

Rearranging eq 1

$$\frac{RM_O}{o} = \frac{RM_C}{c} = Total\ number\ of\ atoms\ in\ the\ relative\ mass = N_A$$

Thus

$$\frac{Relative\ mass\ of\ an\ element\ or\ compound}{Absolute\ mass\ of\ an\ atom\ of\ the\ element\ or\ molecule\ of\ the\ compound} = constant = N_A$$

N_A is defined as the number of atoms of carbon in 12.0000 g of carbon-12 isotope and that number, called *Avogadro's Number*, is $6.022{\cdot}10^{23}$.

The quantity of material that contains Avogadro's Number, $6.022{\cdot}10^{23}$, of particles is called a *mole*.

dozen ⇒ 12 particles

mole ⇒ $6.022{\cdot}10^{23}$ particles

The weight of one mole of atoms is equal to the *atomic weight* (AW) or *molar mass* of the element. Units: g/mole. The AWs of the elements are listed in the periodic table. These AWs are the relative masses of the elements.

The weight of one mole of molecules is equal to the *molecular weight* (MW) or molar mass of the compound. Units: g/mole. *The molecular weight is equal to the sum of the atomic weights of all the atoms in the chemical formula for the compound.*

Example: Calculate the MW of H_2O

$$MW_{H_2O} = 2\,AW_H + AW_O = 2\,(1.008\,\tfrac{g}{mole}) + 15.999\,\tfrac{g}{mole} = 18.015\,\tfrac{g}{mole}$$

The number of moles in a specific weight (wt) of a substance is given by the equation

$$moles\ of\ a\ substance = \frac{wt\ of\ substance\ (g)}{molar\ mass\ of\ substance\ (g/mole)}$$

Example: Find the number of moles in 1.802 g of H_2O.

$$mole\ H_2O = \frac{wt\ of\ H_2O}{MW_{H_2O}} = \frac{1.802\,g}{18.015\,\frac{g}{mole}} = 0.1000\,mole\,H_2O$$

E) Chemical Formulas

1) Covalent (Molecular) Compounds

The *valence* is defined as the combining capacity of the atom.

Compound		**Atom**	**Valence**
HCl	H-Cl	H	1
		Cl	1
BCl_3	Cl Cl-B-Cl	B	3
H_2O	O-H H	O	2

Group in Periodic Table	**Common Valence**
1	1
2	2
13	3
14	4
15	3
16	2
17	1

When writing the chemical formula for a covalent compound, the valence of each atom must be satisfied.

Example: Write the chemical formula for hydrogen sulfide.

Element	**Group**	**Valence**	
H		1	-H
S	16	2	| S -

H
|
S - H ⇒ SH_2

2) Ionic Compounds

Ionic compounds do NOT exist as molecules but rather as a collection of ions. Positive ions are called *cations* and negative ions are called *anions*. The chemical formula indicates the simplest ratio of the ions in the ionic compound. Example: Sodium chloride

$$NaCl \Rightarrow Na^{+} + Cl^{-}$$

has one sodium cation for every chloride anion. The molar mass of an ionic compound is called the *formula weight* (FW), units: g/mole. See handout of Common Ions.

When writing the chemical formula for an ionic compound, the total positive charge of the cations must equal in magnitude the total negative charge of the anions.

Example: Write the chemical formula for lead hydroxide.

See handout of Common Ions

Ions	**Charge**
Pb^{2+}	+2
OH^{-}	- 1

Pb^{2+}, 2 [OH^{-}] $\Rightarrow$ $Pb(OH)_2$

F) Oxidation Number

The oxidation number (oxid. no.) is the charge an atom appears to have in a compound.

Rules for Determining Oxidation Numbers

1) In free elements, the oxidation number of each atom is zero.
Example: In Cl_2, the oxidation number of each Cl atom = 0.

2) In simple ions, the oxidation number is equal to charge on the ion.
Example: For Ca^{2+} the oxidation number of Ca = +2.

3) In most compounds that contain oxygen, the oxidation number of each oxygen atom is -2.
Example: In H_2O, oxidation number of O = -2.

Exception: In peroxides, -O-O-, the oxidation number of each oxygen atom is -1.
Example: In HOOH, oxidation number of each oxygen atom = -1.

4) In most compounds that contain hydrogen, the oxidation number of each hydrogen is +1.
Example: In H_2O, oxidation of each hydrogen atom = +1.

Exception: In metal hydrides, MH, the oxidation number of each hydrogen atom is -1.
Example: In NaH, the oxidation number of H = -1.

5) All oxidation numbers must be consistent with the conservation of charge. In neutral molecules the sum of the oxidation numbers of all the atoms must equal zero. In a complex ion (an ion with two or more atoms), the sum of the oxidation numbers of all the atoms must equal the charge on the ion.

Example: H_2O

$$2(oxid.\ no.\ of\ H) + oxid.\ no.\ of\ O = 2(+1) + -2 = 0$$

Example: Determine the oxidation number of the nitrogen atom in the nitrate ion, NO_3^-.

$$(oxid.\ no.\ of\ N) + 3(oxid.\ no.\ of\ O) = (oxid.\ no.\ of\ N) + 3(-2) = -1$$

$$oxid.\ no.\ of\ N = +5$$

G) Chemical Equations

$$2\ H_2 + O_2 \rightarrow 2\ H_2O$$

The species on the left side of the arrow are called the *reactants* and the species on the right side are called *products*. The numbers to the left of the chemical formulas are called *stoichiometric coefficients*.

For a chemical equation to be valid, it must satisfy the

i) Law of Conservation of Mass,
ii) Law of Conservation of Charge,
iii) and experimental fact.

When criteria i and ii are satisfied, the chemical equation is said to be *balanced*.

Chemical equations for acid-base, precipitation, and complex-formation reactions may be balanced by inspection.

Example: Write a balanced chemical equation for the precipitation of lead sulfate from an aqueous solution of lead nitrate and sodium sulfate.

lead nitrate: Pb^{2+}, 2 $[NO_3^-]$ ⇒ $Pb(NO_3)_2$ lead sulfate: Pb^{2+}, SO_4^{2-} ⇒ $PbSO_4$

sodium sulfate: $2[Na^+]$, SO_4^{2-} ⇒ Na_2SO_4 sodium nitrate: Na^+,NO_3^- ⇒ $NaNO_3$

$$\_\ Pb(NO_3)_{2\ (aq)} + \_\ Na_2SO_{4\ (aq)} \rightarrow \_\ PbSO_{4\ (s)} + \_\ NaNO_{3\ (aq)}$$

where g = vapor state, l = liquid state, s = solid state, and aq = aqueous solution

$$Pb(NO_3)_{2\ (aq)} + Na_2SO_{4\ (aq)} \rightarrow PbSO_{4\ (s)} + 2\ NaNO_{3\ (aq)}$$

Ionic compounds in aqueous solutions exist as dissociated ions, i.e.

$$Pb(NO_3)_2\ (s) + \text{excess water} \rightarrow Pb^{2+}\ (aq) + 2\ NO_3^-\ (aq).$$

$$Pb(NO_3)_2\ (s) \xrightarrow{H_2O} Pb^{2+}\ (aq) + 2\ NO^{3-}\ (aq)$$

Thus, the above chemical equation can be written as

$$Pb^{2+}\,(aq) + 2\,NO_3^-\,(aq) + 2\,Na^+\,(aq) + SO_4^{2-}\,(aq) \rightarrow PbSO_4\,(s) + 2\,NO_3^-\,(aq) + 2\,Na^+\,(aq)$$

The *net ionic equation* (includes only those ions involved in the reaction) for the precipitation reaction is

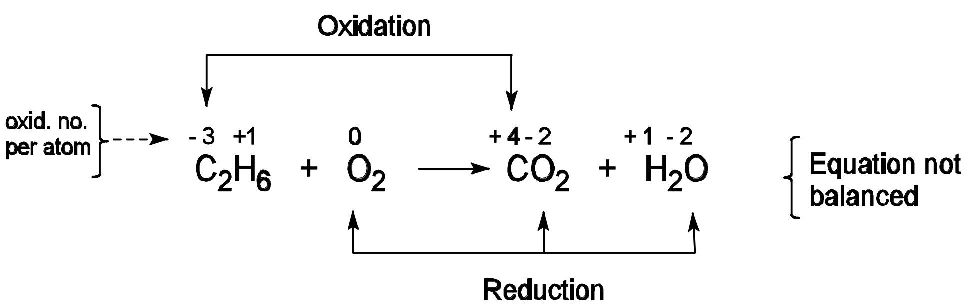
$$Pb^{2+}\,(aq) + SO_4^{2-}\,(aq) \rightarrow PbSO_4\,(s)$$

H) Oxidation-Reduction Reactions

Oxidation refers to a chemical change in which there is an increase in an oxidation number, i.e. electrons are lost.

Reduction refers to a chemical change in which there is a decrease in an oxidation number, i.e. electrons are gained.

An example of an oxidation-reduction reaction



To balance a chemical equation for a simple oxidation-reduction reaction, assign a stoichiometric coefficient of one to the substance with the largest number of atoms and then assign the remaining coefficients so that the equation satisfies the Laws of Conservation of Mass and Charge.

Example: Balance the above chemical equation for the combustion of C_2H_6.

$\underline{1}\ C_2H_6 + \_\ O_2 \rightarrow \_\ CO_2 + \_\ H_2O$

$1\ C_2H_6 + \_\ O_2 \rightarrow \underline{2}\ CO_2 + \underline{3}\ H_2O$

$1\ C_2H_6 + \underline{7/2}\ O_2 \rightarrow 2\ CO_2 + 3\ H_2O$

In some cases it will be advantageous to express the stoichiometric coefficients as integers. Multiply the coefficients in the above equation by 2.

$2\ C_2H_6 + 7\ O_2 \rightarrow 4CO_2 + 6\ H_2O$

I) Stoichiometric Calculations

Example: How many grams of N_2 and H_2 are needed to prepare 34.0 g of NH_3?

Unknown: wt of H_2 and wt of N_2 needed to prepare 34.0 g of NH_3
Knowns: wt of NH_3 = 34.0 g, H_2 and N_2 react to form NH_3
Concepts: stoichiometry, moles, MW, balanced chemical equations

Step #1: Write a balanced chemical equation.

$$\_\ N_2 + \_\ H_2 \rightarrow \_\ NH_3$$

An oxidation-reduction reaction

$$\_\ N_2 + \_\ H_2 \rightarrow \underline{1}\ NH_3$$

$$\underline{1/2}\ N_2 + \underline{3/2}\ H_2 \rightarrow 1\ NH_3$$

$$N_2 + 3\ H_2 \rightarrow 2\ NH_3$$

Step #2: Find moles of the given substances (knowns).

$$mole\ NH_3 = \frac{wt\ of\ NH_3}{MW_{NH_3}} = \frac{34.0g}{17.0\frac{g}{mole}} = 2.00\ mole\ NH_3$$

Step #3: Find moles of the sought substances with the aid of an appropriate mole ratio and the answers in Step #2. The mole ratio is determined from the stoichiometric coefficients or subscripts in a chemical formula.

Units (moles N_2) on left side of equal sign are the same as units on the right side.

Stoichiometric coefficient

$$\text{mole } N_2 = (2.00 \text{ mole } NH_3)\left[\frac{\text{mole } N_2}{\text{mole } NH_3}\right] = (2.00 \text{ mole } NH_3)\left[\frac{1 \text{ mole } N_2}{2 \text{ mole } NH_3}\right] = 1.00 \text{ mole } N_2$$

Stoichiometric coefficient

Mole Ratio

$$mole\ H_2 = (2.00\ mole\ NH_3)\left[\frac{3\ mole\ H_2}{2\ mole\ NH_3}\right] = 3.00\ mole\ H_2$$

Step #4: Express answer in desired units.

$$wt\ N_2 = (mole\ N_2)(MW_{N_2}) = (1.00\ mole\ N_2)\ (28.0\tfrac{g}{mole}) = 28.0 g\ N_2$$

$$wt\ H_2 = (mole\ H_2)(MW_{H_2}) = (3.00\ mole\ H_2)\ (2.0\tfrac{g}{mole}) = 6.00 g\ H_2$$

Example: *Limiting Reactant Problem:* Find the maximum weight of nitric oxide, NO, that can be prepared from 68.0 g of NH_3 and 128 g of O_2.

Unknown: maximum wt of NO that can be prepared
Knowns: wt of NH_3 = 68.0 g, wt of O_2 = 128 g, and NH_3 and O_2 react to form NO and H_2O
Concepts: mole ratio, moles, MW, balanced chemical equations

Step #1:

$$\_\ NH_3 + \_\ O_2 \rightarrow \_\ NO + \_\ H_2O$$

$$\underline{1}\ NH_3 + \_\ O_2 \rightarrow \underline{1}\ NO + \underline{3/2}\ H_2O$$

$$1\ NH_3 + \underline{5/4}\ O_2 \rightarrow 1\ NO + 3/2\ H_2O$$

$$4\ NH_3 + 5\ O_2 \rightarrow 4\ NO + 6\ H_2O$$

Step #2:

$$mole\ NH_3 = \frac{wt\ NH_3}{MW_{NH_3}} = \frac{68.0\ g}{17.0\frac{g}{mole}} = 4.00 mole\ NH_3$$

$$mole\ O_2 = \frac{wt\ O_2}{MW_{O_2}} = \frac{128.0\ g}{32.0\frac{g}{mole}} = 4.00 mole\ O_2$$

Step #3:

One of the reactants is limiting, i.e. it is the reactant that is completely consumed and thus, the moles of the limiting reactant determine the maximum number of moles of product that can be produced. To determine which reactant is the limiting reactant, consider all cases.

Case I: Assume that NH_3 is the limiting reactant, i.e. the 4.00 moles of NH_3 are consumed.

$$mole\ NO = (4.00 mole\ NH_3)\left[\frac{4\ mole\ NO}{4\ mole\ NH_3}\right] = 4.00 mole\ NO$$

Case II: Assume that O_2 is the limiting reactant, i.e. the 4.00 moles of O_2 are consumed.

$$mole\ NO = (4.00 mole\ O_2)\left[\frac{4\ mole\ NO}{5\ mole\ O_2}\right] = 3.20 mole\ NO$$

The case with the smaller (or smallest) number of moles of product is the correct case. **In this example Case II is the correct case**.

Since NH_3 is not the limiting reactant, how many moles of NH_3 will remain after the O_2 has been consumed?

Step #4:

$$wt\ NO = (mole\ NO)(MW_{NO}) = (3.20\,mole\ NO)(30.0\,\tfrac{g}{mole}) = 96.0\,g\ NO$$

The *empirical formula* is the simplest integral ratio in which the atoms combine.

$$molecular\ formula = (empirical\ formula)_i$$

where i is an integer.

Example:

Molecular Formula	**Empirical Formula**	**i**
C_2H_6	CH_3	2
$H_6B_3N_3$	H_2BN	3
C_6H_6	CH	6

$$i = \frac{MW}{FW}$$

where MW is the molecular weight of the compound and FW is the formula weight for the empirical formula.

Procedure for Determining Empirical Formula

Step #1: Find moles of substances formed from given sample of the compound.

Step #2: Find moles of each element in the given sample of the compound.

Step #3: Obtain ratios of the moles of each element to the element with the smallest number of moles. If these ratios are not roughly whole numbers, multiply them by an integer such that the products are whole numbers.

Example: The composition of a compound that contains phosphorus and sulfur is found to be 43.6% P and 56.4% S by weight. What is the empirical formula of the compound?

Unknown: empirical formula of the compound
Knowns: compound is 43.6 % phosphorus and 56.4% sulfur by weight. Assume a 100 g sample of the compound.
Concepts: stoichiometry, moles, molar mass, empirical formula

Step #2: Assume 100 g of the compound.

$$wt\ P = (100g\ sample)\left[\frac{43.6\%\ P}{100\%}\right] = 43.6\ g\ P$$

$$wt\ S = (100g\ sample)\left[\frac{56.4\%\ S}{100\%}\right] = 56.4\ g\ S$$

$$mole\ P = \frac{43.6g}{31.0\frac{g}{mole}} = 1.41 mole\ P \quad and \quad mole\ S = \frac{56.4g}{32.0\frac{g}{mole}} = 1.76 mole\ S$$

Step #3: The element with the smaller number of moles is phosphorus.

$$\frac{1.41 mole\ P}{1.41 mole\ P} = 1.00 \quad ; \quad 1.00 \times 4 = 4$$

$$\frac{1.76 mole\ S}{1.41 mole\ P} = 1.25 \quad ; \quad 1.25 \times 4 = 5$$

$$P_{1.00}S_{1.25} \quad ; \quad P_4S_5$$

Example: A 21.2 g sample of a compound that contained carbon, hydrogen, and oxygen was burned in excess oxygen. The H_2O that was produced weighed 10.80 g and the CO_2 weighed 26.40 g. The MW of the compound was determined to be 212 g/mole. What is the molecular formula of the compound?

Unknown: molecular formula, $C_xH_yO_z$ i.e. what are the values for x, y, and z?
Knowns: 21.2 g sample of $C_xH_yO_z$, 10.80 g of H_2O and 26.40 g of CO_2 produced from 21.2 g sample of $C_xH_yO_z$, MW of $C_xH_yO_z$ is 212 g/mole
Concepts: mole ratio, molar mass, moles, empirical formula, molecular formula

$$C_xH_yO_z + O_2 \rightarrow x\ CO_2 + \frac{y}{2}\ H_2O \qquad NOT\ BALANCED$$

Step #1:

$$mole\ CO_2 = \frac{26.40g\ CO_2}{44.0\frac{g}{mole}} = 0.600 mole\ CO_2$$

$$mole\ H_2O = \frac{10.80g\ H_2O}{18.0\frac{g}{mole}} = 0.600 mole\ H_2O$$

Step #2:

All of the *carbon* in the 21.2 g sample of $C_xH_yO_z$ ends up in the 26.40 g of CO_2.

$$moles\ of\ C\ in\ 26.40g\ of\ CO_2 = moles\ of\ C\ in\ 21.2g\ of\ C_xH_yO_z$$

$$moles\ of\ C\ in\ 26.40g\ of\ CO_2 = (0.600\,mole\ CO_2)\left[\frac{1\ mole\ C}{1\ mole\ CO_2}\right] = 0.600\,mole\ C$$

$$\boldsymbol{moles\ of\ C\ in\ 21.2g\ of\ C_xH_yO_z = 0.600\,mole\ C}$$

All of the *hydrogen* in the 21.2 g sample of $C_xH_yO_z$ ends up in the 10.80 g of H_2O.

$$moles\ of\ H\ in\ 10.80g\ of\ H_2O = moles\ of\ H\ in\ 21.2g\ of\ C_xH_yO_z$$

$$moles\ of\ H\ in\ 10.80g\ of\ H_2O = (0.600\,mole\ H_2O)\left[\frac{2\ mole\ H}{1\ mole\ H_2O}\right] = 1.20\,mole\ H$$

$$\boldsymbol{moles\ of\ H\ in\ 21.2g\ of\ C_xH_yO_z = 1.20\,mole\ H}$$

To find the moles of *oxygen* in the 21.2 g sample of $C_xH_yO_z$, utilize the concept that the total equals the sum of the parts.

$$wt\ of\ sample\ of\ C_xH_yO_z = wt\ C + wt\ H + wt\ O\ \ (in\ the\ sample)$$

$$21.2g\ of\ C_xH_yO_z = (0.600\,mole\ C)(12.01\tfrac{g}{mole}) + (1.20\,mole\ H)(1.008\tfrac{g}{mole}) + wt\ O$$

$$wt\ O = 12.8g\ O$$

$$\boldsymbol{moles\ of\ O\ in\ 21.2g\ of\ C_xH_yO_z = \frac{12.8g\ O}{16.00\frac{g}{mole}} = 0.800\,mole\ O}$$

Step #3:

$$\frac{0.600\,mole\ C}{0.600\,mole\ C} = 1.00\ ;\qquad 1.00 \times 3 = 3$$

$$\frac{1.20\,mole\ H}{0.600\,mole\ C} = 2.00\ ;\qquad 2.00 \times 3 = 6$$

$$\frac{0.800\,mole\ O}{0.600\,mole\ C} = 1.33\ ;\qquad 1.33 \times 3 = 4$$

$$C_{1.00}H_{2.00}O_{1.33}\ ;\qquad C_3H_6O_4$$

Empirical Formula

$$molecular\ formula = (empirical\ formula)_i$$

$$molecular\ formula = (C_3H_6O_4)_i$$

$$i = \frac{MW_{C_xH_yO_z}}{FW_{C_3H_6O_4}}$$

$$FW_{C_3H_6O_4} = 3\ AW_C + 6\ AW_H + 4\ AW_O = 106\frac{g}{mole}$$

$$i = \frac{MW_{C_xH_yO_z}}{FW_{C_3H_6O_4}} = \frac{212\frac{g}{mole}}{106\frac{g}{mole}} = 2$$

$$molecular\ formula = (C_3H_6O_4)_2 = C_6H_{12}O_8$$