Chapter 2: Mass Relations in Formulas, Chemical Reactions, and Stoichiometry

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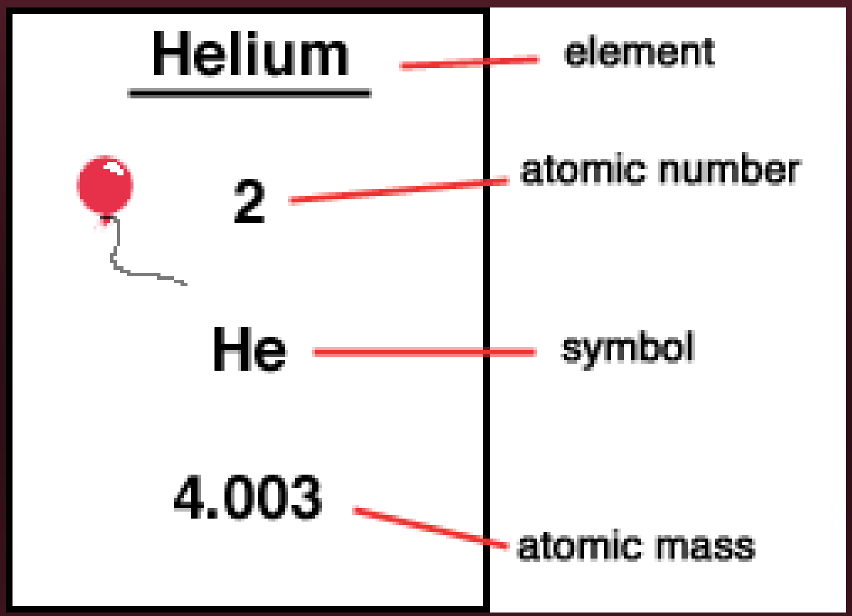
Section 2.1: The Atomic Mass

- The atomic mass is the mass of 1 atom of an element.
- Atoms are very small and their mass is a very small number, with need for a special unit…. the **atomic mass unit** (expressed as **a.m.u**. or just **u**).
- The value of 1 atomic mass unit is chosen as 1/12 of the mass of one carbon-12 isotope.
  - Remember that for carbon-12 the mass number A is equal to 12 (that is carbon-12 has 12 nucleons.
  - The mass of a carbon-12 atom is equal to 1.9926×10^{-23} g.
  - $1 \text{ u} = 1.6605 \times 10^{-24}$ g.
  - The mass of carbon-12 atom is measured with an instrument called the mass spectrometer.

Section 2.1: The Atomic Mass (cont.)

- Atomic masses are given in the Periodic Table and are located below the element symbol.
  - For helium, we note that the atomic mass is not 4 but 4.003 u.
  - This is because the helium has several isotopes and the number 4.003 u is the average atomic mass of all the isotopes of helium present in a typical sample on earth.
- Note: Atomic masses are also called atomic weights.



Section 2.3: Avagadro's Number and the Mole

- Typical samples of matter contain huge numbers of atoms, often numbers as large as 10^{24} or more.
- The **mole** was established as a unit that is very useful when counting the numbers of atoms, ions or molecules.
- One mole is equal to the number of carbon atoms in 12 g of carbon-12.
- 1 atom of carbon-12 has a mass of 1.9927×10^{-23} g. Hence, in 12 g of carbon-12, there are: 6.022×10^{23} atoms.

Section 2.3: Avagadro's Number and the Mole (cont.)

- The number of 6.022×10^{23} is called **Avogadro's number**.
  - Avogadro's number is expressed by the symbol $\mathbf{N_A}$.
  - One mole of atoms of carbon-12 (i.e. 12 g of carbon-12) contains Avogadro's number or 6.022×10^{23} atoms of carbon-12.

- Note: The term "mole" is analogous to the term "dozen". While a dozen eggs refer to twelve eggs, a mole of particles (atoms, ions or molecules) refers to 6.022×10^{23} particles. It follows that while 2 dozen eggs consists of 24 eggs, 2 moles of particles consists of $2 \times (6.022 \times 10^{23})$ particles (i.e. 1.2044×10^{24} particles).

$$\frac{12\ \text{g}}{1.9927 \times 10^{-23}\ \text{g / C atom}} = 6.022 \times 10^{23}\ \text{C atoms}$$

Section 2.4-2.5: The Concept of Formula Mass or Molar Mass

- The formula mass, or molar mass, is the sum of atomic masses in a chemical formula.
- **Examples:**

Chemical Formula	Formula Mass (a.m.u.)	Molar Mass (g/mol)
H	1.0	1.0
H_2	2.0	2.0
O	16.0	16.0
O_2	32.0	32.0

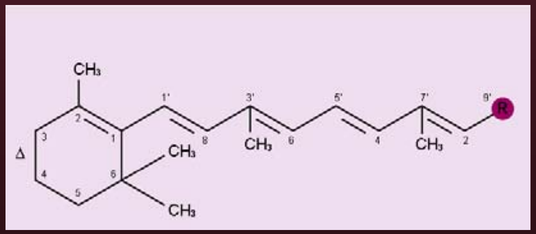
Section 2.4-2.5: The Concept of Formula Mass or Molar Mass (cont.)

- **Formula mass:** the sum of atomic masses of all atoms in a formula of any molecular or ionic compound.
  - The formula mass is expressed in a.m.u.

- **Molar mass:** the sum of atomic masses of all atoms in a mole of pure substance.
  - The molar mass is expressed in g/mol.
- It is important that you know how to write chemical formulas from chemical names and vice versa so that your math will include the correct number of atoms to multiply masses by to find the mass of whole molecules.

Section 2.4-2.5: Concept of Formula Mass or Molar Mass (cont.)

- **Example: Calculate the molar mass of vitamin A, $C_{20}H_{30}O$.**



- The vitamin A molecule consists of:

Element	Number of Atoms	Atomic Mass
C	20 x	12.0 = 240.0
H	30 x	1.0 = 30.0
O	1 x	16.0 = 16.0
		286.0

- The molar mass of vitamin A is 286.0 g/mol.

Section 2.4-2.5: Concept of Formula Mass or Molar Mass (cont.)

- **Example: Calculate the molar mass of sodium chloride.**
  - Note: Here the chemical name is given but not the chemical formula. It is important to know the names and formulas of chemicals.
  - Sodium chloride has the chemical formula NaCl.

Element	Number of Atoms	Atomic Mass
Na	1 X	23.0 = 23.0
Cl	1 X	35.5 = 35.5
		58.5 g/mol

- The molar mass of sodium chloride is 58.5 g/mol.

Section 2.6-2.8: Conversion between Moles and Mass

- When discussing the amount of a substance, it is common practice to use the word “moles” instead of the more rigorous wording “number of moles”.
- The symbol “n” is used to describe the number of moles or “moles” of a substance.
- How many moles of a substance are present in a given sample?
  - We calculate this from the mass and the molar mass of that substance, according to the equation:

$$\text{Number of moles of a substance is equal to : } n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

- mass = n x molar mass
- Thus, if the moles and the chemical formula are given, one can calculate the mass in grams of that chemical.

Section 2.6-2.8: Conversion between Moles and Mass (cont.)

- **Example: Calculate the number of moles of NH_3 in 1.0 g of NH_3.**

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

- The mass of NH_3 is 1.0 g. However, we need to calculate the molar mass of NH_3.

Element	Number of Atoms	Atomic Mass
N	1 x	14.0 = 14.0
H	3 x	1.0 = 3.0
		17.0 g/mol

$$n = \frac{1.0\text{ g}}{17.0\text{ g/mol}} = 0.0588235\text{ mol}$$

- In the correct number of significant figures, the answer is 0.059 mol.

Section 2.9-2.10: Problems on Avagadro's Number

- In some instances, we want to know how many atoms, ions or molecules are involved in a chemical or physical process.
- We can calculate this number from:
  - The mass of the substance,
  - The molar mass of the substance, and
  - Avogadro's number.
- Avogadro's number is given by the symbol $N_A = 6.022 \times 10^{23}$ or 6.022E23

$$\text{Remember mole, } n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \text{ and one mole} = 6.022 \times 10^{23}$$

Section 2.9-2.10: Problems on Avagadro's Number (cont.)

- **A sample of the compound, C_3H_6O, contains 14.0 x 10^{14} carbon atoms.**

- **(a) Calculate the number of C_3H_6O molecules**
  - Each molecule of C_3H_6O contains 3 carbon atoms. Therefore, the number of C_3H_6O molecules containing 14.0 x 10^{14} C atoms is:

$$\frac{14.0 \times 10^{14}}{3} = 4.67 \times 10^{14} \text{ molecules of } C_3H_6O$$

- **A sample of the compound, C_3H_6O, contains 14.0 x 10^{14} carbon atoms.**

- **(b) Calculate the number of moles of C_3H_6O**.

$$\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times 4.67 \times 10^{14} \text{ molecules} = 7.75 \times 10^{-10} \text{ or } 7.75E-10 \text{ mol}$$

Section 2.9-2.10: Problems on Avagadro's Number (cont.)

- **A sample of the compound, C_3H_6O, contains 14.0 x 10^{14} carbon atoms.**
- **(c) Calculate the number of grams of C_3H_6O.**

Element	Number of Atoms	Atomic Mass
C	3 x	12.0 = 36.0
H	6 x	1.0 = 6.0
O	1 x	16.0 = 16.0
		58.0 g/mol

$$\mathbf{n} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

- Therefore, mass (g) = n x molar mass (g/mol)

Thus, mass = 7.75 x 10^{-10} mol x 58.0 g/mol

mass = 4.50 x 10^{-8} or 4.50E-8 g

Section 2.11-2.12: Percent Composition

- The percent composition of a compound is the mass percent of the elements present.

$$\text{Mass percent of element} = \frac{\text{mass of element in sample}}{\text{total mass of sample}} \times 100$$

- Knowing the chemical formula of a compound, the mass percent of its constituent elements can be calculated.
- Note: the subscripts in a chemical formula allow us to define the atom ratio as well as the mole ratio in which the different elements are combined.

Section 2.13-2.14: Empirical Formula

- When a new compound is formed or discovered, it is important to determine the chemical formula. This is done by taking a known amount of sample and decomposing it into its component elements.

Or

- Taking a known amount of sample and reacting it with oxygen to produce CO_2 and H_2O. The component elements or CO_2 and H_2O, are then collected and weighed. The results of such analyses give the mass of each element in the compound.
- This is used to determine the mass percent of each element in the compound. Knowing the mass percent of each element in the compound makes it possible to determine its chemical formula.
- Empirical formula is the simplest chemical formula.
- The simplest formula gives only the ratios of atoms in a compound.

Section 2.13-2.14: Empirical Formula (cont.)

- **Example: A 50.00 g sample contains 13.28 g of potassium, 17.68 g of chromium, and 19.04 g of oxygen. Find the simplest formula.**
  - *Analyze the problem. The sample mass is given, the masses of elements are also given.*
  - *Make sure the sum of masses of all elements in the sample is equal to the mass of the sample.*

Section 2.13-2.14: Empirical Formula (cont.)

- **Step1:** Calculate the number of moles of K, Cr and O in the given masses.

$$\text{Remember that moles} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$13.28\ \text{g (K)} \Rightarrow \frac{13.28\ \text{g}}{39.1\ \text{g/mol}} = 0.340\ \text{mol (K)}$$

$$17.68\ \text{g (Cr)} \Rightarrow \frac{17.68\ \text{g}}{52.0\ \text{g/mol}} = 0.340\ \text{mol (Cr)}$$

$$19.04\ \text{g (O)} \Rightarrow \frac{19.04\ \text{g}}{16.0\ \text{g/mol}} = 1.19\ \text{mol (O)}$$

Section 2.13-2.14: Empirical Formula (cont.)

- **Step 2:** Divide each of the numbers of moles by the smallest number of moles to obtain the relative amounts in moles of each element in the substance.

$$K: \frac{0.340\ \text{mol}}{0.340\ \text{mol}} = 1.00$$

$$Cr: \frac{0.340\ \text{mol}}{0.340\ \text{mol}} = 1.00$$

$$O: \frac{1.19\ \text{mol}}{0.340\text{mol}} = 3.50$$

Section 2.13-2.14: Empirical Formula (cont.)

- **Step 3:** Write the formula using these relative numbers of moles of each element.
  - Remember that the subscripts in a formula give the relative numbers of atoms or moles of atoms in that substance.
  - The results in Step 2 suggest that the simplest formula is: $K_1Cr_1O_{3.5}$

Section 2.13-2.14: Empirical Formula (cont.)

- **Step 4:** Write the final formula, ensuring all subscripts are whole numbers.
  - We multiply each subscript by 2 to get the empirical formula: $\mathbf{K_2Cr_2O_7}$. This is potassium dichromate.
  - This formula makes sense because the dichromate ion has a -2 charge. The potassium ion has a +1 charge. Hence, this substance has a neutral formula, as it should.

Section 2.15-2.16: More Problems on Empirical Formula

- To find the composition of a substance, it is often useful to react that substance with oxygen gas.
- Combustion reactions are useful for the chemical analysis of substances containing carbon and hydrogen as they produce CO_2 and H_2O.
- Measuring the amount of CO_2 and H_2O produced allows the determination of how much carbon and hydrogen are present in that amount of substance.

Section 2.15-2.16: More Problems on Empirical Formula (cont.)

- **Example: When 5.000 g of ibuprofen is burnt with oxygen gas ($O_{2(g)}$), 13.86 g of $CO_{2(g)}$ and 3.926 g of $H_2O_{(l)}$ are formed. Use the following information to determine the empirical formula of ibuprofen. Ibuprofen is known to contain only carbon, oxygen and hydrogen elements.**

- **Step 1:** How much carbon is there in 5.000 g of ibuprofen?
  - All the carbon in ibuprofen ends up in the 13.86 g of $CO_{2(g)}$. So, the question is how many moles of carbon are present in 13.86 g of $CO_{2(g)}$?

$$13.86\,g\,(CO_2) \Rightarrow \frac{13.86\,g\,(CO_2)}{44.01\ g/mol\ (CO_2)} = 0.3149\ mol\,(CO_2) \Rightarrow 0.3149\,mol\,(C)$$

  - ⇨ mass (C) in 5 g ibuprofen = 0.3149 mol (C) x 12.01 g/mol (C) = 3.782 g (C)

Section 2.15-2.16: More Problems on Empirical Formula (cont.)

- **Step 2:** How much hydrogen is there in 5.000 g of ibuprofen?

  - All the hydrogen in ibuprofen ends up in the 3.926 g of $H_2O_{(l)}$. So, the question is how many moles of hydrogen are present in 3.926 g of $H_2O_{(l)}$.

$$3.926\ \mathrm{g\ (H_2O)} \Rightarrow \frac{3.926\ \mathrm{g\ (H_2O)}}{18.00\ \mathrm{g/mol\ (H_2O)}} = 0.2181\ \mathrm{mol\ (H_2O)} \Rightarrow 0.4362\ \mathrm{mol\ (H)}$$

  - ⇨ mass (H) in 5 g ibuprofen = 0.4362 mol (H) x 1.00 g/mol (H) = 0.4362 g (H)

Section 2.15-2.16: More Problems on Empirical Formula (cont.)

- **Step 3:** How much oxygen is there in 5.000 g of ibuprofen?

  - Since ibuprofen only contains oxygen, carbon and hydrogen and 5.000 g of ibuprofen contain 3.782 g (C) and 0.4362 g (H), then, the mass of oxygen is:

    Mass (O) = 5.000 g – 3.782 g – 0.4362 g = 0.7818 g (O).

Section 2.15-2.16: More Problems on Empirical Formula (cont.)

- **Step 4:** Now, we can use the strategy shown in Section 3.13 to determine the formula of ibuprofen.

$$C: 3.782\,g \Rightarrow \frac{3.782\,g\,(C)}{12.0\ g/mol\ (C)} = 0.3149\ mol\,(C)$$

$$H: 0.4362\,g \Rightarrow \frac{0.4362\,g\,(H)}{1.0\ g/mol\ (H)} = 0.4362\ mol\,(H)$$

$$O: 0.7818\,g \Rightarrow \frac{0.7818\,g\,(O)}{16.0\ g/mol\ (O)} = 0.04886\ mol\,(O)$$

$$\text{For Carbon} \Rightarrow \frac{0.3149\,mol\,(C)}{0.04886\,mol\,(O)} = 6.4\ (C)\ \text{per}\,(O)\ \text{approximately}\ 6.5$$

$$\text{For Hydrogen} \Rightarrow \frac{0.4362\,mol\,(H)}{0.04886\,mol\,(O)} = 8.9\ (H)\ \text{per}\,(O)\ \text{approximately}\ 9$$

- Hence, the formula for ibuprofen is: $C_{6.5}H_9O_1$ or more appropriately, **$C_{13}H_{18}O_2$.**

Section 2.17-2.18: Molecular Formula

- The empirical formula of a substance is always written using the smallest possible whole number subscripts to give the relative number of atoms of each element in the substance.

  - The empirical formula for sodium chloride is written as NaCl and not Na_2Cl_2.
  - NaCl is an ionic compound, not a molecule.
  - The entity NaCl is called a **formula unit**.

Remember that for ionic compounds, the chemical formula and the empirical formula are always one and the same formula.

For molecular (covalent) compounds, however, molecular and empirical formula may be different.

Section 2.17-2.18: Molecular Formula (cont.)

- A molecular formula is a whole number multiple of the simplest chemical formula.

Or

- A molecular formula is a whole multiple of the empirical formula.

$$\text{Multiple} = \frac{\text{molar mass}}{\text{empirical formula mass}}$$

- To find the multiple, the molar mass is needed. The empirical formula mass can be calculated from the empirical formula.

Section 2.17-2.18: Molecular Formula (cont.)

- **Example: The mass composition of lindane is 24.78% C, 2.08% H and 73.14% Cl. The molar mass of lindane is 290.85 g/mol. Determine the molecular formula.**

Element	Mass (g)	Molar Mass (g/mol)	Moles	Mole Ratio
C	24.78	12.0	2.06	1
H	2.08	1.0	2.08	1
Cl	73.14	35.5	2.06	1

Section 2.17-2.18: Molecular Formula (cont.)

- The simplest formula or the empirical formula is CHCl.
- The empirical formula mass is 12 + 1 + 35.5 = 48.5 g/mol

$$\text{Multiple} = \frac{290.85\ \text{g/mol}}{48.5\ \text{g/mol}} = 6$$

- The molecular formula is: $C_6H_6Cl_6$

Section 2.19: Balancing Chemical Reactions

- Chemical reaction: an actual transformation of substances called reactants into substances called products.
- To represent a chemical reaction we use a chemical equation, a sort of recipe which shows in a symbolic form
  - 1) who the participating substances are (reactants and products),
  - 2) the state or phase these substances are in (solid, liquid, gas, aqueous solution) and
  - 3) the amount in which they must be present (number of atoms, molecules (for covalent compounds) or formula units (for ionic compounds)).

Section 2.19: Balancing Chemical Reactions (cont.)

- **Example:** Consider the reaction of aluminum metal ($Al_{(s)}$) with solid iron oxide ($Fe_2O_{3\,(s)}$) forming solid aluminum oxide ($Al_2O_{3\,(s)}$) and solid iron ($Fe_{(s)}$). This reaction is represented by the following chemical equation:

$$\underbrace{2\ Al_{(s)} + Fe_2O_{3\,(s)}}_{\textbf{REACTANTS}} \rightarrow \underbrace{Al_2O_{3\,(s)} + 2\ Fe_{(s)}}_{\textbf{PRODUCTS}}$$

- The arrow (→) shows the direction in which the chemical transformation takes place.
- The **reactants** (shown on the left-hand side of the arrow) are the substances with which the reaction is started.
- The **products** (shown on the right-hand side of the arrow) are the substances resulting from the reaction.
- In the above reaction all substances are in the solid state, as indicated by the subscript “s” in parentheses.

Section 2.19: Balancing Chemical Reactions (cont.)

- The **state** or **phase** of a substance is always indicated by a **subscript in parentheses** after the chemical formula of that substance.

  The following notations are used for the various phases encountered in chemical reactions:
  - (s) solids
  - (l) liquids
  - (g) gases
  - (aq) aqueous solutions

Section 2.19: Balancing Chemical Reactions (cont.)

- One of the most important pieces of information conveyed by a chemical equation is the **number of atoms, ions, formula units or molecules** associated with each substance.
- **Stoichiometric coefficients:** The number <u>in front of</u> each substance

$$\underbrace{2\,Al_{(s)} + Fe_2O_{3\,(s)}}_{\textbf{REACTANTS}} \rightarrow \underbrace{Al_2O_{3\,(s)} + 2\,Fe_{(s)}}_{\textbf{PRODUCTS}}$$

- For the above reaction, the stoichiometric coefficients are 2, 1, 1 and 2, respectively. Note that when a stoichiometric coefficient is 1, it is not shown.
- Stoichiometric coefficients insure that the number of atoms of each type is the same on the reactants and products sides.
- **Dalton's hypothesis**: in a chemical reaction, atoms are neither destroyed nor created.

Section 2.19.4: Balancing the Chemical Reaction: $CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$

- Always start balancing the elements that are present in the least number of compounds. In the above equation, we can start with either carbon or hydrogen, because oxygen is in everything.

- **Balancing the element carbon.**
  - There is one carbon atom on the left-hand side (in $CO_{2(g)}$) and six carbon atoms on the right-hand side (in $C_6H_{12}O_{6(s)}$).
  - Place a coefficient of 6 in front of $CO_{2(g)}$ and 1 in front of $C_6H_{12}O_{6(s)}$.

- **Balancing the element hydrogen.**
  $6\ CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$
  - There are two hydrogen atoms on the left-hand side (in $H_2O_{(l)}$) and 12 hydrogen atoms on the right-hand side (in $C_6H_{12}O_{6(s)}$).
  - Place a coefficient of 6 in front of $H_2O_{(l)}$.

Section 2.19.4: Balancing the Chemical Reaction: $CO_{2(g)} + H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$

- **Balancing the element oxygen.**
  $6\ CO_{2(g)} + 6\ H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + O_{2(g)}$
  - There are 12 + 6 = 18 oxygen atoms on the left-hand side.
  - On the right-hand side here are 6 oxygen atoms in $C_6H_{12}O_{6(s)}$ and 2 oxygen atoms in $O_{2(g)}$.
  - To obtain 18 oxygen atoms on the right-hand side, we must assign a stoichiometric coefficient of 6 to $O_{2(g)}$.

- The balanced chemical equation is:
  $6\ CO_{2(g)} + 6\ H_2O_{(l)} \rightarrow C_6H_{12}O_{6(s)} + 6\ O_{2(g)}$

- The sum of the stoichiometric coefficients for reactants and products is **19**.

Section 2.19.5: Balancing the Equation of a Combustion Reaction

- A combustion reaction is a reaction in which a substance (element or compound) is burnt with oxygen gas (O_2) and leads to the formation of carbon dioxide (CO_2) gas and water (H_2O).

- Consider the combustion reaction of hexane, $C_6H_{14(l)}$.

  $C_6H_{14(l)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$

- **balance carbon:**
  - There are 6 carbon atoms in C_6H_{14} and 1 carbon atom in CO_2. Use stoichiometric coefficients of 1 (not shown) for C_6H_{14} and 6 for CO_2.

- **balance hydrogen:**
  - There are 14 hydrogen atoms in one molecule of C_6H_{14} and 2 hydrogen atoms in one molecule of H_2O. Use a coefficient of 7 for H_2O.

Section 2.19.5: Balancing the Equation of a Combustion Reaction (cont.)

- **balance oxygen:**
  - On the left-hand side there are 2 oxygen atoms. On the right-hand side there are 6x2 + 7x1 = 19 oxygen atoms. However, oxygen is present as O_2 on the reactant side. Use 19/2 O_2 molecules on the right-hand side.
  - The balanced equation is: $C_6H_{14(l)} + 19/2\ O_{2(g)} \rightarrow 6\ CO_{2(g)} + 7\ H_2O_{(l)}$

- to use only whole numbers as stoichiometric coefficients, multiply all coefficients by 2.

  $\mathbf{2}\ C_6H_{14(l)} + \mathbf{19}\ O_{2(g)} \rightarrow \mathbf{12}\ CO_{2(g)} + \mathbf{14}\ H_2O_{(l)}$

- The sum of the stoichiometric coefficients for reactants and products is 47.

Section 2.19.6: Balancing the Chemical Reaction: $Fe_2(SO_4)_{3(aq)} + 6\ NH_{3(g)} + 6\ H_2O_{(l)} \rightarrow 2\ Fe(OH)_{3(s)} + 3\ (NH_4)_2SO_{4(s)}$

- **balance iron:**
  - There are 2 iron atoms in $Fe_2(SO_4)_3$ and 1 iron atom in $Fe(OH)_3$. Use 1 (not shown) for $Fe_2(SO_4)_3$ and 2 for $Fe(OH)_3$.
- **balance sulfur:**
  - There are 3 sulfur atoms in one formula unit of $Fe_2(SO_4)_3$ and 1 sulfur atom in one formula unit of $(NH_4)_2SO_4$. Use 3 for $(NH_4)_2SO_4$.
- **balance nitrogen:**
  - On the left, there is 1 nitrogen atom in NH_3.
  - On the right, there are 3 x 2 = 6 nitrogen atoms in 3 formula units of $(NH_4)_2SO_4$.
  - Use 6 for NH_3.

Section 2.19.6: Balancing the Chemical Reaction:
$Fe_2(SO_4)_{3(aq)} + 6\ NH_{3(g)} + 6\ H_2O_{(l)} \rightarrow 2\ Fe(OH)_{3(s)} + 3\ (NH_4)_2SO_{4(s)}$

- **balance oxygen:**
  - On the left, there are 3 x 4 = 12 oxygen atoms in one formula unit of $Fe_2(SO_4)_3$ and 1 oxygen atom in one molecule of $H_2O_{(l)}$.
  - On the right, there are 18 oxygen atoms [2 x 3 = 6 oxygen atoms in 2 formula units of $Fe(OH)_3$ and 3 x 4 = 12 oxygen atoms in 3 formula units of $(NH_4)_2SO_4$].
  - Use 6 for H_2O.

- **balance hydrogen:** we expect hydrogen to be balanced as all stoichiometric coefficients have already been assigned, and there are 30 hydrogen atoms on each side. Done!

- **fully balanced chemical equation:** $Fe_2(SO_4)_{3(aq)} + \mathbf{6}\ NH_{3(g)} + \mathbf{6}\ H_2O_{(l)} \rightarrow \mathbf{2}\ Fe(OH)_{3(s)} + \mathbf{3}\ (NH_4)_2SO_{4(s)}$

- The sum of the stoichiometric coefficients for products and reactants is **18.**

Section 2.20-2.21: Stoichiometry

- We are interested in **reaction stoichiometry** whenever we ask questions such as:

  - What is the amount of each reactant required to produce a known amount of product?

  Or

  - What is the amount of product formed from a known amount of reactants?

- Practically speaking, “amounts” of reactants or products are the masses of these compounds measured in the laboratory in grams.

Section 2.20-2.21: Stoichiometry (cont.)

- To relate masses of products to masses of reactants requires relating:

  - Masses to moles using the molar masses, and, moles or reactants to moles of products, using the stoichiometry of the balanced chemical reaction (that is, using the values of the stoichiometric coefficients).

Section 2.20-2.21: Stoichiometry (cont.)

- **Example: In the Haber process, nitrogen reacts with hydrogen to produce ammonia gas.**

$$N_{2(g)} + 3\ H_{2(g)} \longrightarrow 2\ NH_{3(g)}$$

  - The coefficients (1, 3 and 2) represent the number of moles.
  - This equation can be represented as: one mole of N_2 reacts with three moles of H_2 to produce two moles of NH_3.

- Whenever we carry out stoichiometric calculations (relating masses of products to masses of reactants), we will always follow the 4 steps given below.
  - Write the balanced chemical equation.
  - Convert masses given for reactants and products to moles, using the molar masses.
  - Write down the **mole ratio** using the stoichiometric coefficients of the balanced chemical equation.

$$\text{Mole Ratio} = \frac{\text{Moles Desired}}{\text{Moles Given}}$$

Section 2.20-2.21: Stoichiometry (cont.)

- **Moles Desired**: Stoichiometric coefficient of the substance for which we want to calculate the amount reacted or produced.
- **Moles Given:** Stoichiometric coefficient of the substance for which we know the amount reacted or produced.
- Multiply the mole ratio by the number of moles given in the problem or calculated in step 2. Then, convert the calculated number of moles to the mass if necessary.

Section 2.22-2.23: Theoretical Yield, Limiting Reactant and Percent Yield

- Consider the chemical reaction: $2\ Sb_{(s)} + 3\ Cl_{2(g)} \rightarrow 2\ SbCl_{3(s)}$
- The reaction equation provides a recipe for the preparation of antimony chloride.
  - We need 3 moles of chlorine gas for every two moles of antimony solid to form 2 moles of antimony chloride.
  - Whenever reactants are present in the relative amount of 3 moles of chlorine for 2 moles of antimony, we say that the reactants are present in the **stoichiometric amount**.

- <u>Whenever we have reactants present in the stoichiometric amount, we can use any of the reactants to calculate how many moles of product are formed from a given amount of reactants.</u>

- **Theoretical Yield:** the maximum mass of products that can be formed from a given amount of reactants, assuming the reaction is complete.

Section 2.22-2.23: Theoretical Yield, Limiting Reactant and Percent Yield (cont.)

- **For example**: If we react 2 moles of Sb with 3 moles of Cl_2, we obtain 2 moles of $SbCl_3$.
  - The mass of 2 moles of $SbCl_3$ is equal to 2 x (121.8 + 3 x 35.5) = 456.6 g.
  - We say that when the reaction is carried out with 2 moles of Sb and 3 moles of Cl_2, the theoretical yield is equal to 456.6 g.
- When the reactants are present in the stoichiometric amount and the reaction is complete, there is **no reactant left at the end of the reaction**.
- <u>What happens if reactants are not present in the stoichiometric amount</u>?
  - Some of the reactants are completely consumed in the reaction (**limiting reactant**)
  - And some reactants remain at the end of the reaction (**excess reactant**).

Section 2.22-2.23: Theoretical Yield, Limiting Reactant and Percent Yield (cont.)

- **Experimental Yield, Theoretical Yield and Percent Yield**
  - For a variety of reasons, the vast majority of chemical reactions carried out in laboratories or in chemical plants do not produce the maximum amount of product possible (theoretical yield).
  - Remember the **Theoretical Yield** is calculated from the amounts of reactants present and from the balanced chemical reaction.
  - **Experimental Yield:** the amount of products <u>actually</u> obtained from a given amount of reactants.
  - The **Percent Yield** is defined as the ratio of experimental yield to theoretical yield multiplied by 100.

$$\%\ \text{Yield} = 100 \times \frac{\text{Experimental Yield}}{\text{Theoretical Yield}}$$