Stoichiometry

– Studies the quantitative aspects of chemical reactions

3.1 The Mole

- Unit for amount of substance in terms of the number of entities (atoms, molecules, ...) in it.

1 mol of entities → # of atoms in 12 g of ^{12}C

1 atom of $^{12}C \rightarrow 1.99265\times10^{-23}$ g^{12}C (mass spectrometry)

12 g^{12}C×[1 atom/1.99265×10^{-23} g^{12}C] = 6.022×10^{23} atoms

⇒1 mol of entities → 6.022×10^{23} entities

- Avogadro's number (N_A) – number of entities per 1 mol → N_A **= 6.022×10^{23} /mol**

- Conversion between moles and entities

  [1 mol entities/6.022×10^{23} entities]

Example:

1) How many molecules of water are present in 2.7 mol of water?

2) How many atoms of hydrogen are present in 2.7 mol of water?

$$2.7\ \cancel{\text{mol } H_2O}\left(\frac{6.022\times10^{23}\text{ molec. } H_2O}{1\ \cancel{\text{mol } H_2O}}\right) = 1.6\times10^{24}\text{ molec. } H_2O$$

$$1.6\times10^{24}\ \cancel{\text{molec. } H_2O}\left(\frac{2\text{ atoms H}}{1\ \cancel{\text{molec. } H_2O}}\right) = 3.3\times10^{24}\text{ atoms H}$$

- The atomic mass (in *amu*) of an element is numerically equal to the mass (in *g*) of 1 mol of the element
  - $^{12}C \rightarrow$ 12 amu 1 mol $^{12}C \rightarrow$ 12 g (definitions)
  - C → 12.01 amu 1 mol C → 12.01 g
  - H → 1.008 amu 1 mol H → 1.008 g
  - O → 16.00 amu 1 mol O → 16.00 g
- The molecular (formula) mass (in *amu*) of a compound is numerically equal to the mass (in *g*) of 1 mol of the compound
  - $CO_2 \rightarrow$ 44.01 amu 1 mol $CO_2 \rightarrow$ 44.01 g

⇒1 mol of a substance has a fixed mass (can be used to measure moles of substances by weighing them)

Molar Mass (M)

- Mass of a substance per 1 mol of its entities
  - element → atoms (or molecules for H_2, O_2, P_4, ...)
  - molecular compound → molecules
  - ionic compound → formula units
- Units of M → **g/mol**
- $M = m_{particle} \times N_A$

Example:

What is the molar mass of ^1H, if the mass of 1 atom ^1H is 1.673×10^{-24} g?

M **= 1.673×10^{-24}g × 6.022×10^{23} /mol = 1.007 g/mol**

- M is numerically equal to the atomic, molecular, or formula mass of the substance
  - For elements, M = atomic mass (from per. table)
  - For molecular compounds and molecular elements, M = molecular mass
  - For ionic compounds, M = formula mass

⇒ For compounds and molecular elements, M equals the sum of the molar (atomic) masses of the elements in the formula

Example:

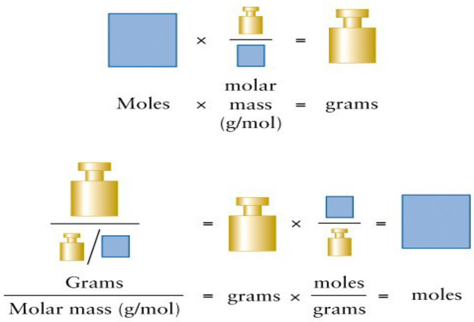
Calculate the molar masses of O_2 and Li_2O.

$M(O_2)$ = 2×16.00 = 32.00 g/mol

$M(Li_2O)$ = 2×6.941 +1×16.00 = 29.88 g/mol

- $\boldsymbol{M}$ can be used as a conversion factor
- Conversion between moles ($\boldsymbol{n}$) and mass ($\boldsymbol{m}$)

$$m = n \times M \quad \leftrightarrow \quad n = m/M$$



- Conversion between moles ($\boldsymbol{n}$) and masses ($\boldsymbol{m}$) of elements

Example:

What is the mass of 1.221 mol Kr?

$$m = \mathbf{1.221\ mol \times 83.80\ g/mol = 102.3\ g}$$

Example:

How many moles of atoms are present in 1.23 g of Kr?

$$1.23\ \cancel{\text{g Kr}} \times \left(\frac{1\ \text{mol Kr}}{83.80\ \cancel{\text{g Kr}}}\right) = 1.47 \times 10^{-2}\ \text{mol Kr}$$

- Conversion between moles ($\boldsymbol{n}$) and masses ($\boldsymbol{m}$) of compounds (same as for elements)

Example:

Calculate the number of moles of urea, $(NH_2)_2CO$, in 2.3×10^5 kg of this compound.

$$\boldsymbol{M} = \mathbf{2\times14.00 + 4\times1.008 + 1\times12.01 + 1\times16.00 = 60.04\ g/mol}$$

$$2.3\times10^5\ \cancel{\text{kg urea}} \times \left(\frac{10^3\ \cancel{\text{g urea}}}{1\ \cancel{\text{kg urea}}}\right) \times \left(\frac{1\ \text{mol urea}}{60.04\ \cancel{\text{g urea}}}\right)$$

$$= 3.8\times10^6\ \text{mol urea}$$

- Conversion between masses and number of entities of elements and compounds

Example:

Calculate the number of CO_2 molecules and oxygen atoms in **15.8 g** of CO_2.

$$\boldsymbol{M}(\mathbf{CO_2}) = \mathbf{12.01 + 2\times16.00 = 44.01\ g/mol}$$

$$\mathbf{15.8\ \cancel{g\ CO_2}\left(\frac{1\ \cancel{mol\ CO_2}}{44.01\ \cancel{g\ CO_2}}\right)\left(\frac{6.022\times10^{23}\ molec.CO_2}{1\ \cancel{mol\ CO_2}}\right)}$$

$$\mathbf{= 2.16\times10^{23}\ molec.CO_2}$$

$$\mathbf{2.16\times10^{23}\ \cancel{molec.CO_2}\left(\frac{2\ atoms\ O}{1\ \cancel{molec.CO_2}}\right) = 4.32\times10^{23}\ atoms\ O}$$

- Conversion between masses of compounds and masses of their elements using chemical formulas (*The subscripts in formulas refer to individual atoms as well as to moles of atoms*)

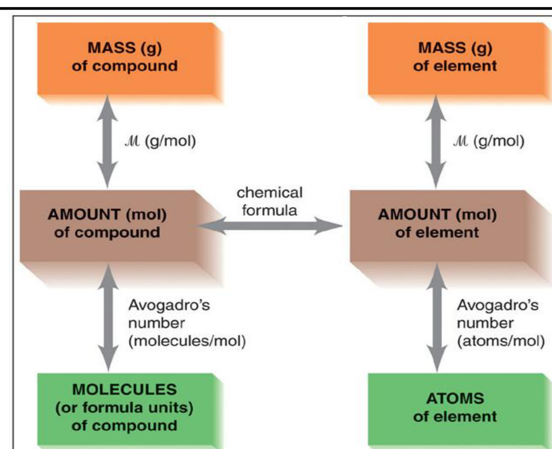
Example: What is the mass of H in **5.00 g** CH_4?

$$\mathbf{CH_4 \rightarrow \boldsymbol{M} = 1\times12.01 + 4\times1.008 = 16.04\ g/mol}$$

$$\mathbf{H \rightarrow \boldsymbol{M} = 1.008\ g/mol}$$

$$5.00\ \cancel{\text{g CH}_4} \times \left(\frac{1\ \cancel{\text{mol CH}_4}}{16.04\ \cancel{\text{g CH}_4}}\right) \times \left(\frac{4\ \cancel{\text{mol H}}}{1\ \cancel{\text{mol CH}_4}}\right) \times$$

$$\times \left(\frac{1.008\ \text{g H}}{1\ \cancel{\text{mol H}}}\right) = 1.26\ \text{g H}$$



Mass Percentage Composition

- Percentage by mass of each element in a compound

$$\textbf{Mass\%} = [m_{element}/m_{compound}]\times 100\%$$

- Calculation of Mass% from chemical formulas
  - Consider 1 mol of a compound

m_{comp} = M of comp

m_{elem} = (# moles of elem in 1 mol of comp)×(M of elem)

Note: The # of moles of the element in 1 mol of the compound equals the # of atoms of the element in the formula of the compound

$$\frac{\text{Mass\% of element}}{} = \left[\frac{(\#\text{ atoms of element in formula})(M\text{ of element})}{(M\text{ of compound})}\right]\times 100\%$$

Example:

Calculate the Mass% of O in CO_2.

$CO_2 \rightarrow M = 1\times 12.01 + 2\times 16.00 = 44.01$ g/mol

$O \rightarrow M = 16.00$ g/mol

$$\text{Mass\% O} = \left(\frac{2\times 16.00\text{ g/mol}}{44.01\text{ g/mol}}\right)\times 100\% = 72.71\%$$

- Calculation of Mass% from chemical analysis

Example:

Calculate the mass percentage of C in nicotine, if analysis shows that 5.00 g of nicotine contain 3.70 g C, 0.44 g H and 0.86 g N.

$$\text{Mass\% C} = \left(\frac{3.70\text{ g C}}{5.00\text{ g nicotine}}\right)\times 100\% = 74.0\%$$

3.2 Determination of Chemical Formulas

- **Molecular formulas** – numbers of atoms of each element in a molecule
- **Empirical formulas** – relative numbers of atoms of each element using the smallest whole numbers

Example:

acetic acid $\rightarrow C_2H_4O_2$ **(MF)** $\rightarrow CH_2O$ **(EF)**

formaldehyde $\rightarrow CH_2O$ **(MF)** $\rightarrow CH_2O$ **(EF)**

glucose $\rightarrow C_6H_{12}O_6$ **(MF)** $\rightarrow CH_2O$ **(EF)**

Determining Empirical Formulas

- **Elemental analysis** – gives the masses of the elements in a given mass of the compound or the Mass% composition
- EF from Mass% data
  1. Consider 100 g of the compound
  2. The masses of the elements equal their mass%
  3. Convert the masses of the elements to moles
  4. Determine the relative number of moles (mol ratio)
  5. Simplify the mole ratio to whole numbers
- EF from mass data
  - Omit steps 1 and 2 above

Example:

- Determine the EF of nicotine, if the mass% of C, H and N in it are 74.0, 8.7 and 17.3 %, respectively.

1. Consider 100 g nicotine
2. **74.0 g C, 8.7 g H, 17.3 g N**
3. Convert masses to moles:

74.0 g C×(1 mol C/12.01 g C) = 6.16 mol C

8.7 g H×(1 mol H/1.008 g H) = 8.6 mol H

17.3 g N×(1 mol N/14.01 g N) = 1.23 mol N

4. Mol ratio:

6.16 mol C : 8.6 mol H : 1.23 mol N

5. Simplify the mole ratio:

(divide by the smallest number, and if necessary, multiply by a factor to get whole numbers)

6.16/1.23 = 5.01 ≅ 5

8.6/1.23 = 7.0 ≅ 7

1.23/1.23 = 1.00 ≅ 1

→ simplest whole-number ratio:

5 mol C : 7 mol H : 1 mol N

→ EF:

$\mathbf{C_5H_7N}$

Determining Molecular Formulas

- The MF is a whole-number multiple of the EF

$$\Rightarrow \boldsymbol{M} = \boldsymbol{M_{EF}} \times \mathbf{n}$$

– $\boldsymbol{M}$ → molar mass

– $\boldsymbol{M_{EF}}$ → EF mass

– **n** → whole number (number of EFs per molecule)

$$\Rightarrow \mathbf{n} = \boldsymbol{M/M_{EF}}$$

- Determining MFs from EFs and molar masses

Example:

- The empirical formula of nicotine is $\mathbf{C_5H_7N}$ and its molar mass is 162.23 g/mol. MF = ?

$\boldsymbol{M_{EF}}$ **→ 5×12.01+7×1.008+1×14.01 = 81.12 g/mol**

$$\mathbf{n} = \frac{M}{M_{EF}} = \frac{162.23\text{ g/mol}}{81.12\text{ g/mol}} = 2.000 \cong 2$$

⇒MF = 2 × EF

MF →$\mathbf{C_{10}H_{14}N_2}$

Combustion Analysis

- A method for elemental analysis of combustible organic compounds through their combustion in excess O_2
  – The **C** in the sample is converted to $\mathbf{CO_2}$ which is absorbed in a NaOH absorber and weighed
  – The **H** in the sample is converted to $\mathbf{H_2O}$ which is absorbed in a P_4O_{10} absorber and weighed
  – If a third element (O, N, …), it passes through the absorbers

1 mol C from the sample → 1 mol $\mathbf{CO_2}$

2 mol H from the sample → 1 mol $\mathbf{H_2O}$

Example:

When 0.236 g aspirin is burned in excess $\mathbf{O_2}$, 0.519 g $\mathbf{CO_2}$ and 0.0945 g $\mathbf{H_2O}$ are formed. Determine the mass % of C, H and O in aspirin.

- **Calculate the masses of C and H in the sample based on the masses of CO_2 and H_2O:**
- **Calculate the mass of O by subtracting the masses of C and H from the total mass of the sample**

$$0.519\text{ g }CO_2 \times \left(\frac{1\text{ mol }CO_2}{44.01\text{ g }CO_2}\right) \times \left(\frac{1\text{ mol C}}{1\text{ mol }CO_2}\right) \times \left(\frac{12.01\text{ g C}}{1\text{ mol C}}\right) = 0.142\text{ g C}$$

$$0.0945\text{ g }H_2O \times \left(\frac{1\text{ mol }H_2O}{18.02\text{ g }H_2O}\right) \times \left(\frac{2\text{ mol H}}{1\text{ mol }H_2O}\right) \times \left(\frac{1.008\text{ g H}}{1\text{ mol H}}\right) = 0.0106\text{ g H}$$

0.236-0.142-0.0106=0.084 g O

$$\%C = \left(\frac{0.142\ g\ C}{0.236\ g\ sample}\right) \times 100\% = 60.0\%$$

$$\%H = \left(\frac{0.0106\ g\ H}{0.236\ g\ sample}\right) \times 100\% = 4.48\%$$

$$\%O = \left(\frac{0.084\ g\ O}{0.236\ g\ sample}\right) \times 100\% = 35.5\%$$

The empirical formula can be determined from the percentage composition in a subsequent step

3.3 Chemical Equations

- Represent chemical reactions

Reactants → Products

- Skeletal equations – show identities of reactants and products

$$H_2 + O_2 \rightarrow H_2O$$

- Law of conservation of mass
  - Atoms are neither created nor destroyed (they only change bonding partners)
  - Same atoms are present in the reactants as in the products

- Balanced chemical equations
  - same number of atoms of each element appear on each side of the equation
  - stoichiometric coefficients – needed to balance the equations

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

  - Microscopic view

$$(2\ molec.\ H_2 + 1\ molec.\ O_2 \rightarrow 2\ molec.\ H_2O)$$

  - Macroscopic view

$$(2\ mol\ H_2 + 1\ mol\ O_2 \rightarrow 2\ mol\ H_2O)$$

$$(4.032\ g\ H_2 + 32.00\ g\ O_2 \rightarrow 36.032\ g\ H_2O)$$

  - The stoichiometric coefficients can be treated as relative number of moles of reactants and products
  - Physical state symbols
    - (s) solid; (l) liquid; (g) gas; (aq) aqueous solution

$$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$$

Balancing Chemical Equations

- Balancing by inspection (only simple cases)
  - Change stoichiometric coefficients only
  - Never change subscripts in formulas
  - Never add other substances to the equation

- Systematic method
  1. Write the skeletal equation
  2. Balance one element at a time using coefficients
     - Start with the elements in the most complex substance and finish with those in the least complex one
     - Alternatively, start with the element present in the fewest number of formulas and finish with the element present in the greatest number of formulas
     - Use fractional coefficients if necessary
  3. If necessary multiply the whole equation by a factor to clear the fractional coefficients
  4. Verify that the equation is balanced and the coefficients are the smallest whole numbers
  5. Specify physical states

Example: Write the balanced equation for the combustion of ethane, C_2H_6, to carbon dioxide and liquid water.

$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$ **skeletal**

$C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$ **for C**

$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$ **for H**

$C_2H_6 + (7/2)O_2 \rightarrow 2CO_2 + 3H_2O$ **for O**

multiply eq. by 2

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

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$$

– Often polyatomic ions can be treated as single entities

Example: Balance the following skeletal eq. in aqueous solution:

$$Co(NO_3)_3 + (NH_4)_2S \rightarrow Co_2S_3 + NH_4NO_3$$

→**balance Co and S:**

$$2Co(NO_3)_3 + 3(NH_4)_2S \rightarrow Co_2S_3 + NH_4NO_3$$

→ **balance NH_4 and NO_3:**

$$2Co(NO_3)_3 + 3(NH_4)_2S \rightarrow Co_2S_3 + 6NH_4NO_3$$

→**add physical state symbols:**

$$2Co(NO_3)_{3(aq)} + 3(NH_4)_2S_{(aq)} \rightarrow Co_2S_{3(s)} + 6NH_4NO_{3(aq)}$$

3.4 Calculating Amounts of Reactants and Products

Stoichiometric Equivalences

- Balanced chemical equations contain definite stoichiometric relations between reactants and products → stoichiometric **mole ratios**

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

$$\left.\begin{array}{l} \text{2 mol } H_2 \Leftrightarrow \text{1 mol } O_2 \\ \text{2 mol } H_2 \Leftrightarrow \text{2 mol } H_2O \\ \text{1 mol } O_2 \Leftrightarrow \text{2 mol } H_2O \end{array}\right\} \text{stoichiometric equivalences}$$

$$\left.\begin{array}{l} \text{1 mol } O_2 / \text{2 mol } H_2 \\ \text{2 mol } H_2O / \text{2 mol } H_2 \\ \text{2 mol } H_2O / \text{1 mol } O_2 \end{array}\right\} \text{stoichiometric mole ratios}$$

Mole-to-Mole Conversions

- Conversion method
  – The mole ratios are used as conversion factors

(mol given)×(mole ratio) = (mol required)

Example: Determine the number of moles of water produced from **3.4 mol O_2**.

→ balanced equation: $2H_2 + O_2 \rightarrow 2H_2O$

→ mole ratio (conversion factor): **[2 mol H_2O/1 mol O_2]**

$$3.4\,\cancel{\text{mol } O_2} \times \left(\frac{2\text{ mol } H_2O}{1\,\cancel{\text{mol } O_2}}\right) = 6.8\text{ mol } H_2O$$

- Stoichiometric conversion factors are reaction specific

Example: Calculate the amount of O_2 needed to produce **3.5 mol H_2O** by combustion of methane (CH_4).

→ balanced equation:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

→ mole ratio (conversion factor):

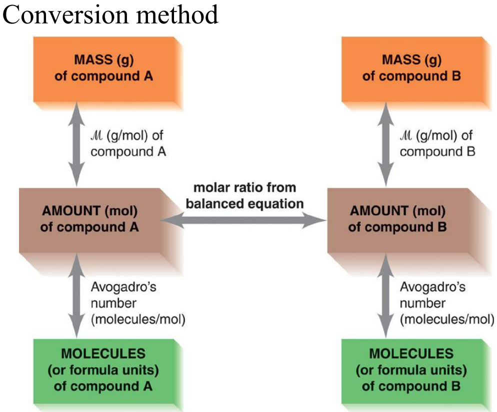
2 mol O_2 ⇔ 2 mol H_2O

[2 mol O_2/2 mol H_2O]

$$3.5\,\cancel{\text{mol } H_2O} \times \left(\frac{2\text{ mol } O_2}{2\,\cancel{\text{mol } H_2O}}\right) = 3.5\text{ mol } O_2$$

Mass-to-Mass Calculations

- Conversion method



Example: Calculate the mass of oxygen needed to completely burn **5.4 kg** of butane (C_4H_{10}).

→ balanced equation:

$$2C_4H_{10} + 13O_2 \rightarrow 8CO_2 + 10H_2O$$

→ mole ratio: **[13 mol O_2/2 mol C_4H_{10}]**

→ molar masses:

C_4H_{10} → 58.1 g/mol O_2 → 32.0 g/mol

$$5.4\,\cancel{\text{kg } C_4H_{10}} \times \left(\frac{10^3\,\cancel{\text{g } C_4H_{10}}}{1\,\cancel{\text{kg } C_4H_{10}}}\right) \times \left(\frac{1\,\cancel{\text{mol } C_4H_{10}}}{58.1\,\cancel{\text{g } C_4H_{10}}}\right) \times$$

$$\left(\frac{13\,\cancel{\text{mol } O_2}}{2\,\cancel{\text{mol } C_4H_{10}}}\right) \times \left(\frac{32.0\text{ g } O_2}{1\,\cancel{\text{mol } O_2}}\right) = 1.9 \times 10^4\text{ g } O_2 = 19\text{ kg } O_2$$

Reaction Yield

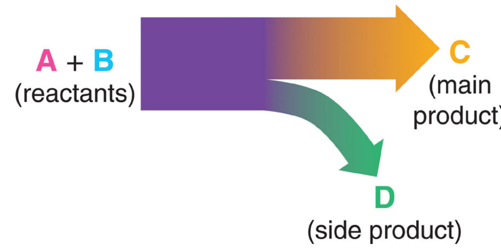
- **Theoretical yield** - the maximum amount of product that can be expected from a given amount of reactant
- **Actual yield** - the actual amount of product isolated in a reaction

Actual Yield ≤ Theoretical Yield

- **Percentage yield**:

$$\textbf{\% Yield} = \frac{\textbf{Actual Yield}}{\textbf{Theoretical Yield}} \times \textbf{100\%}$$

- Reasons for the difference between actual and theoretical yield
  - incomplete reaction
  - loss of product
  - side reactions



Example: Calculate the theoretical yield of carbon dioxide produced by the combustion of **25.0 g** propane (C_3H_8) in excess oxygen.

→balanced equation:

$$\mathbf{C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O}$$

→mass-to-mass conversion:

$$\mathbf{25.0\ \cancel{g\ C_3H_8} \times \left(\frac{1\ \cancel{mol\ C_3H_8}}{44.09\ \cancel{g\ C_3H_8}}\right) \times \left(\frac{3\ \cancel{mol\ CO_2}}{1\ \cancel{mol\ C_3H_8}}\right) \times}$$

$$\mathbf{\times \left(\frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}}\right) = 74.9\ g\ CO_2} \rightarrow \textbf{\textit{Theor. Yield}}$$

Example: Calculate the percentage yield of carbon dioxide, if the combustion of **25.0 g** propane in excess oxygen yields **48.5 g** carbon dioxide.

→ theoretical yield (from prev. problem): **74.9 g CO_2**

→ side reaction (consumes some of the propane):

$$\mathbf{2C_3H_8 + 7O_2 \rightarrow 6CO + 8H_2O}$$

→ actual yield: **48.5 g CO_2**

→ percentage yield:

$$\textbf{\% Yield} = \frac{\mathbf{48.5\ g\ CO_2}}{\mathbf{74.9\ g\ CO_2}} \times \textbf{100\% = 64.8\%}$$

Limiting Reactants

- Reactants present in equivalent amounts
  - All reactants are consumed at the same time
- Nonequivalent amounts of reactants
  - One reactant, called **limiting reactant**, is consumed before the others
  - The other reactants are **in excess**
- Limiting reactant
  - The reaction stops when the limiting reactant is consumed
  - Limits the maximum amount of product achievable (limits the theoretical yield)
  - Stoichiometric calculations based on the limiting reactant give the **lowest amount of product** compared to calculations based on the other reactants

Example: Identify the limiting reactant in the reaction of **5.0 mol H_2** with **3.0 mol N_2**, and determine the theoretical yield of NH_3 in this reaction.

→balanced equation: $\mathbf{3H_2 + N_2 \rightarrow 2NH_3}$

→calculate the theoretical yield based on each of the reactants and chose the **smaller result**:

$$\mathbf{3.0\ \cancel{mol\ N_2} \times \left(\frac{2\ mol\ NH_3}{1\ \cancel{mol\ N_2}}\right) = 6.0\ mol\ NH_3}$$

$$\mathbf{5.0\ \cancel{mol\ H_2} \times \left(\frac{2\ mol\ NH_3}{3\ \cancel{mol\ H_2}}\right) = 3.3\ mol\ NH_3} \rightarrow \textbf{\textit{Theor. Yield}}$$

smaller amount

⇒ **H_2 is the limiting reactant**

Example: Calculate the theoretical yield of HNO_3 in the reaction of **28 g NO_2** and **18 g H_2O** by the chemical equation:

$$3NO_2(g) + H_2O(l) \rightarrow 2HNO_3(l) + NO(g).$$

→Calculate the theoretical yield based on each of the reactants and chose the smaller result:

$$18\ \cancel{g\ H_2O} \times \left(\frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g\ H_2O}}\right) \times \left(\frac{2\ \cancel{mol\ HNO_3}}{1\ \cancel{mol\ H_2O}}\right) \times \left(\frac{63.0\ g\ HNO_3}{1\ \cancel{mol\ HNO_3}}\right) = 130\ g\ HNO_3$$

$$28\ \cancel{g\ NO_2} \times \left(\frac{1\ \cancel{mol\ NO_2}}{46.0\ \cancel{g\ NO_2}}\right) \times \left(\frac{2\ \cancel{mol\ HNO_3}}{3\ \cancel{mol\ NO_2}}\right) \times \left(\frac{63.0\ g\ HNO_3}{1\ \cancel{mol\ HNO_3}}\right) = 26\ g\ HNO_3 \rightarrow \textit{Theor. Yield}$$

smaller amount

⇒**The smaller amount of product results from the calculation based on NO_2**

⇒**NO_2 is the limiting reactant and 26 g HNO_3 is the theoretical yield**

3.5 Solution Stoichiometry

- Solutions – homogeneous mixtures
  - Solvent and solute(s)
  - Solution concentration

Molarity (M)

- Measure of the solute concentration

$$M = \left(\frac{\text{amount of solute (mol)}}{\text{volume of solution (L)}}\right) \quad \text{or} \quad M = \frac{n}{V}$$

- Units – molar (M) **1 M = 1 mol/L**

- Preparation of solutions with known molarity
  - Transfer a known mass of solute in a volumetric flask
  - Dissolve in small amount of water
  - Add water to the mark

Example:

Calculate the molarity of a solution prepared by dissolving **5.33 g NaOH** in water using a **100.0 mL** volumetric flask.

⇒ **convert the mass to moles:**

$$5.33\ \cancel{g\ NaOH} \times \left(\frac{1\ mol\ NaOH}{40.00\ \cancel{g\ NaOH}}\right) = 0.133\ mol\ NaOH$$

⇒ **convert volume to liters:** **100.0 mL = 0.1000 L**

⇒ **divide moles by solution volume:**

$$\frac{0.133\ mol\ NaOH}{0.1000\ L\ solution} = 1.33\ mol\ NaOH/L \rightarrow 1.33\ M\ NaOH$$

- Molarity as a conversion factor

Example:

Calculate the mass of NaOH in **2.50 L** of **1.33 M NaOH** solution.

$$2.50\ \cancel{L}\left(\frac{1.33\ \cancel{mol\ NaOH}}{1\ \cancel{L}}\right)\left(\frac{40.00\ g\ NaOH}{1\ \cancel{mol\ NaOH}}\right) = 133\ g\ NaOH$$

Example:

Calculate the volume of **1.33 M NaOH** solution that contains **5.00 mol NaOH**.

$$5.00\ \cancel{mol\ NaOH} \times \left(\frac{1\ L}{1.33\ \cancel{mol\ NaOH}}\right) = 3.76\ L$$

Dilution

- Reducing the concentration of the solute by adding more solvent
- Stock solutions – concentrated solutions used to store reagents
- Dilution Procedure
  - Use a pipette to measure a small volume of the concentrated solution and transfer it to a volumetric flask
  - Add solvent to fill the volumetric flask to the mark

- Dilution calculations
  - dilution doesn't change the total # of moles of solute in the solution

$n = M \times V \qquad n_d = n_c \qquad M_d \times V_d = M_c \times V_c$

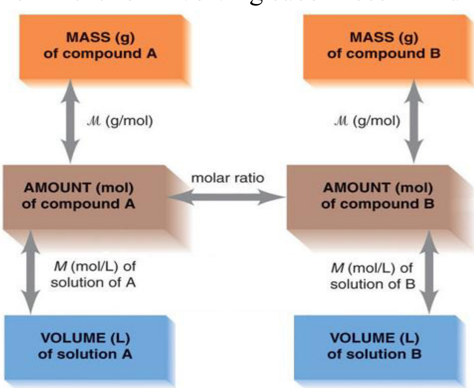
Example:

Calculate the molarity of a solution prepared by dilution of **5.00 mL 2.0 M HCl** stock solution to **100.0 mL**.

$$M_d = \frac{M_c \times V_c}{V_d} = \frac{2.0\ \text{M} \times 5.00\ \text{mL}}{100.0\ \text{mL}} = 0.10\ \text{M}$$

Stoichiometric calculations in solution

- For a reaction involving substances A and B



Example:

What volume of **0.0836 M H_2SO_4** solution will react completely with **16.4 mL 0.255 M KOH**.

⇒ **balanced equation:**

$2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

⇒ **mole ratio:** $[1\ \text{mol}\ H_2SO_4/2\ \text{mol KOH}]$

$$16.4 \times 10^{-3}\,\cancel{\text{L}} \times \left(\frac{0.255\ \cancel{\text{mol KOH}}}{1\ \cancel{\text{L}}}\right) \times \left(\frac{1\ \cancel{\text{mol}\ H_2SO_4}}{2\ \cancel{\text{mol KOH}}}\right) \times \left(\frac{1\ \text{L}}{0.0836\ \cancel{\text{mol}\ H_2SO_4}}\right) = 25.0 \times 10^{-3}\,\text{L} = 25.0\ \text{mL}$$

Limiting reactant problems in solution

Example:

What mass of **H_2** gas can be produced by the reaction of **2.5 g Zn** with **2.0 L 0.15 M HCl** solution. The other product is $ZnCl_2(aq)$.

⇒ **balanced equation:**

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

⇒ **mole ratios:** $[1\ \text{mol}\ H_2/2\ \text{mol HCl}]$

$[1\ \text{mol}\ H_2/1\ \text{mol Zn}]$

⇒ **Calculate the mass of H_2 produced based on both reactants and choose the smaller amount**

⇒ **calculation based on HCl:**

$$2.0\,\cancel{\text{L}} \times \left(\frac{0.15\ \cancel{\text{mol HCl}}}{1\ \cancel{\text{L}}}\right) \times \left(\frac{1\ \cancel{\text{mol}\ H_2}}{2\ \cancel{\text{mol HCl}}}\right) \times \left(\frac{2.02\ \text{g}\ H_2}{1\ \cancel{\text{mol}\ H_2}}\right)$$

$= 0.30\ \text{g}\ H_2$

⇒ **calculation based on Zn:**

$$2.5\ \cancel{\text{g Zn}} \times \left(\frac{1\ \cancel{\text{mol Zn}}}{65.4\ \cancel{\text{g Zn}}}\right) \times \left(\frac{1\ \cancel{\text{mol}\ H_2}}{1\ \cancel{\text{mol Zn}}}\right) \times \left(\frac{2.02\ \text{g}\ H_2}{1\ \cancel{\text{mol}\ H_2}}\right)$$

$= \boxed{0.077\ \text{g}\ H_2}$ ← **smaller amount**

⇒ **The calculation based on Zn yields less product so Zn is the limiting reactant**