

Module 1 - Matter and its Properties

Matter - has mass and occupies space.

3 States of Matter

| State | Definition | Examples |
|--------|--|-----------------------------|
| Solid | rigid; has a fixed shape and volume | ice cube, diamond, iron bar |
| Liquid | has a definite volume but takes the shape of its container | gasoline, water, blood |
| Gas | has no fixed volume or shape; takes the shape of its container | air, helium, oxygen |

Phase Changes of Matter

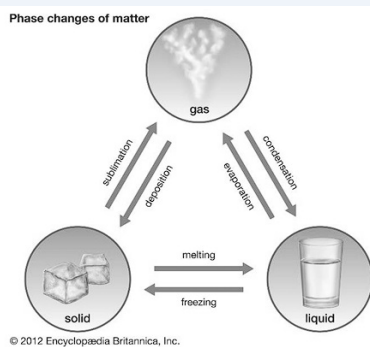


Figure 1.1. Phase Changes of Matter

Elements and Compounds

| | | |
|----------|--|---|
| Elements | cannot be broken down into other substances by chemical means | iron, aluminum, oxygen, and hydrogen |
| Compound | substances that have the same composition no matter where we find them; can be broken down into elements | Water (H ₂ O), Salt (NaCl), Ammonia (NH ₃) |

Physical and Chemical Properties and Changes

| | |
|---------------------|---|
| Physical Properties | odor, color, volume, state (gas, liquid, or solid), density, melting point, boiling point |
| Chemical Properties | burning, digestion, fermentation, rusting, electrolysis |

Other Properties

| | | |
|-----------|---|---|
| Extensive | changes when the amount of material changes | mass, length, volume, shape |
| Intrinsic | does not depend on the size of the material | temperature, odor, color, hardness, density |

Mixture and Pure Substances

| | |
|----------------|---|
| Mixture | has variable composition |
| Homogenous | also called a solution; does not vary in composition from one region to another |
| Heterogenous | contains regions that have different properties from those of other regions |
| Pure Substance | always have the same composition; either elements or compounds |

Types of bonds

| | | |
|----------|---|--|
| Ionic | when one atom shifts or transfers an electron to another atom; metals + nonmetals | Na ⁺ (1A) and Cl ⁻ (7A) creates a stable bond (octet rule) |
| Covalent | atoms share electrons; nonmetals | O ²⁻ (6A) and 2 atoms of H ⁺ (1A) = H ₂ O |

Types of bonds (cont)

Metallic a metal shares an electron with another metal; positively charged ions in electrons

Module 2 - Isotopes, Compounds, Empirical Formula

Atoms have a constant or fixed number of protons

Atomic Number - gives the protons in the nucleus of an atom; represented as **Z**

Neutral Atom - number of protons is equal to the number of electrons

$Z = \text{nuclear charge} = \text{number of protons} = \text{number of electrons in neutral form}$

Mass Number - sum of the number of protons and neutrons; represented by **A**

An atom can be represented by the nuclear symbol ${}^A_Z\text{E}$

Nucleons - protons + neutrons

John Dalton's Atomic Theory

All atoms of an element have the same mass, although isotopes are atoms of the same element but has different numbers of protons

Ex: All carbon atoms ($Z=6$) have 6 protons and electrons, but only 98.89% of naturally occurring carbon atoms have 6 neutrons ($A=12$)

Chemical Compounds

Radicals/Polyatomic Ions - stable groups which form chemical bonds as an intact unit.

The valence number is taken as one.

If a molecule contains **more than one radical** (At least two unpaired electrons), the formula uses **parentheses**. Calcium Phosphate -



Some Polyatomic Ions

| Monovalent (1^-) | | Bivalent (2^-) | | Trivalent (3^-) | |
|----------------------|----------------------------------|--------------------|------------------------|---------------------|---------------|
| Ammonium | NH_4^+ | Carbonate | CO_3 | Phosphate | PO_4 |
| Acetate | $\text{C}_2\text{H}_3\text{O}_2$ | Chromate | CrO_4 | Borate | BO_3 |
| Chlorate | ClO_3 | Oxalate | C_2O_4 | | |
| Chlorite | ClO_2 | Sulfate | SO_3 | | |

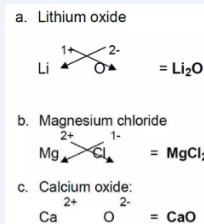
Some Polyatomic Ions (cont)

| | | | |
|-------------|----------------|----------|---------------|
| Bicarbonate | HCO_3 | Sulfite | SO_2 |
| Biculfate | HSO_4 | Peroxide | O_2 |
| Hydroxide | OH | | |
| Nitrate | NO_3 | | |
| Nitrite | NO_2 | | |

Diatomic Molecules

| | |
|---------------|----------|
| H_2 | hydrogen |
| N_2 | nitrogen |
| F_2 | fluorine |
| O_2 | oxygen |
| I_2 | iodine |
| Cl_2 | chlorine |
| Br_2 | bromine |

Criss-Cross Method



-Determine the charge or valence number of the elements

-Exchange their valence numbers

-Reducing by their gcd is possible

Calculating Empirical Formula

Percentage Composition - amounts of the elements for a given amount of compound

Empirical Formula - simplest formula of any compound (smallest ratio of moles); derived from mass analysis

- Determine the given number of moles in each element

- Divide each by the smallest number of moles given

- Multiply each by the smallest number that will turn them into whole numbers.

Example

$Zn_{0.21}P_{0.14}O_{0.56}$
 Converting the fraction to whole numbers:
 1. Divide each subscript by the smallest one, which in this case is 0.14:
 $Zn_{0.21}P_{0.14}O_{0.56}$ → $Zn_{1.5}P_{1.0}O_{4.0}$
 2. Multiply through by the smallest integer that turns all subscripts into integers. We multiply by 2 to make the 1.5 (subscript of Zn) into an integer.
 $Zn_{1.5 \times 2}P_{1.0 \times 2}O_{4.0 \times 2}$ → $Zn_3P_2O_8$

Calculating Empirical Formula with Molar mass

| Given | Mass number | Quotient | Quotient (x/0.75) | Smallest Ratio (y*2) |
|-----------|-------------|----------|-------------------|----------------------|
| 28.03% Mg | 24.035 | 1.15 | 1.53≈1.5 | 3 |
| 21.6% Si | 28.086 | 0.75 | 1 | 2 |
| 1.16% H | 1.16 | 1.15 | 1.53≈1.5 | 3 |
| 49.21% O | 49.21 | 3.08 | 4.1=4 | 8 |

Answer = $Mg_3Si_2H_3O_8$

- Divide the given percent composition to the mass number of each element
- Divide each quotient by the smallest number among them
- Multiply the quotients by the smallest number that will make them whole

For more examples:

[Chem Calculation Worksheet](#)

Calculating Molecular Formula By Empirical Formula

| Empirical Composition | Mass number | Product (rounded off) | Product (emp* (mass/x)) |
|-----------------------|-------------|-----------------------|-------------------------|
| Mg ₃ | 24.305 | 73 | 6 |
| Si ₂ | 28.086 | 56 | 4 |
| H ₃ | 1.008 | 3 | 6 |
| O ₈ | 15.999 | 128 | 16 |
| | | Σ = 260 | |

Calculating Molecular Formula By Empirical Formula (cont)

Suppose the molar mass is 520.8; divide it by the summation (520.8/260 ≈ 2). Multiply 2 by the empirical composition of each element. Answer = $Mg_6Si_4H_6O_{16}$

- Get the summation summation of the product of each empirical composition to their mass number
- Divide the summation from the molar mass
- Multiply the quotient to the empirical composition of each element

Module 3 - Molar Mass, Chem Reactions, Eq

Mole(mol) - SI unit for determining molar mass; amount of substance that contains the same number of atoms in 12g of Carbon-12

Avogadro's number - $6.02214076 \times 10^{23}$

Elements - mass in amu of 1 atom of an element is the same as the mass in grams of 1 mole of atoms of the element

Mass of S (32.07 amu) is equal to the mass of 1 mol ($6.02214076 \times 10^{23}$) of S (32.07 amu)

Calculating Molecular Mass/Weight

| Composition | Number of Atoms | Mass Number (amu) | Product (amu) |
|----------------|-----------------|-------------------|---------------|
| H ₂ | 2 | 1.008 | 2.02 |
| O | 1 | 16.00 | 16.00 |
| | | | Σ = 18.02 |

- Determine the number of atoms of each element then multiply to their corresponding mass number
- Get the summation of the products

Writing and Balancing Chem Eq

Law of Conservation of Mass - mass is neither created nor destroyed in a chemical reaction

Antoine Lavoisier - French chemist; proponent

Reactants - starting material in a chemical reaction

Product - substance formed in a chemical reaction

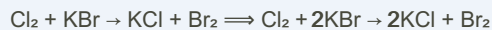
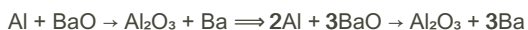
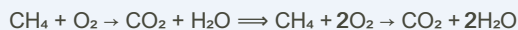
Reactants → Products

"to yield" or "to form" (→)

"to react with" or "to combine with" (+)



Examples



Types of Chemical Reactions

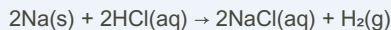
| Type | Definition | Example |
|--|--|---|
| Combination/- Synthesis | two or more reactants combine to form a single product | $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$ |
| Decomposition | one reactant breaks down into two or more products | $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$ |
| Single Displacement | one element is substituted for another element in a compound | $\text{K} + \text{NaCl} \rightarrow \text{KCl} + \text{Na}$ |
| Double Displacement/salt metathesis | two substances react by exchanging ions to produce two new molecules | $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$ |

Module 4 - Mass Relationships in Chem Reactions

Stoichiometry - quantitative relationship between reactants and products in a chemical reaction

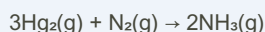
Stoichiometric coefficient - added before an element, ion, or molecule to balance chemical reactions

Mole method using mole-mole factor:



2 moles Na \equiv 2 moles NaCl; hence,

Calculating Amount of Product and Reactant



How many moles of H_2 are needed to produce **26.5 moles of NH_3** ?

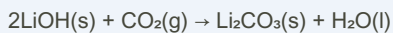
$$26.5 \text{ moles } \text{NH}_3 \times (3 \text{ moles } \text{H}_2) / (2 \text{ moles } \text{NH}_3) = 39.8 \text{ moles of } \text{H}_2$$

How many moles of NH_3 will be produced if **33.7 moles of N_2** reacts completely with H_2

$$33.7 \text{ moles of } \text{N}_2 \times (2 \text{ moles of } \text{NH}_3) / (1 \text{ mole of } \text{N}_2) = 67.4 \text{ moles of } \text{NH}_3$$

-In using mole-mole factor, the arrangement of fractions is done in a way that there is cancellation of similar units

Calculating " " with Molar Mass



How many grams of CO_2 can be absorbed by **236.1 g of LiOH**?

$$236.1\text{g LiOH} \times (1 \text{ mole of LiOH}) / (23.95\text{g LiOH}) \times (1 \text{ mole } \text{CO}_2) / (2 \text{ moles LiOH}) \times (44.01\text{g of } \text{CO}_2) / (1 \text{ mole } \text{CO}_2) = 221.1\text{g } \text{CO}_2$$

-Determine the mass of each element and add each to the given compounds

(Li=6.941, O=15.999, H=1.008, C=12.011)

-Since the given number has 4 significant figures, the numbers also have 4 significant figures

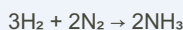
Limiting and Excess Reagent

In chemical reactions, the amount of reactants isn't always stoichiometrically exact, so scientists use cheaper reactants (excess)

Limiting Reagent - Reagent that is completely reacted or used up

Excess Reagent - Reactant present with higher quantity than what is required to react in a limiting reagent

Example



Suppose 6 moles of H_2 was mixed with 4 moles of N_2 . To determine which is the limiting reagent, the amount of NH_3 must be computed given the moles of H_2 and N_2 and the mole-mole factor of the equation

Solution

$$\begin{aligned} \text{moles } \text{NH}_3 &= \# \text{ moles of } \text{H}_2 \times \frac{2 \text{ moles } \text{NH}_3}{3 \text{ moles } \text{H}_2} \\ \text{moles } \text{NH}_3 &= 6 \text{ moles of } \text{H}_2 \times \frac{2 \text{ moles } \text{NH}_3}{3 \text{ moles } \text{H}_2} \\ \text{moles } \text{NH}_3 &= 4 \text{ moles} \\ \text{moles } \text{NH}_3 &= \text{moles of } \text{N}_2 \times \frac{2 \text{ moles } \text{NH}_3}{1 \text{ mole } \text{N}_2} \\ \text{moles } \text{NH}_3 &= 4 \text{ moles of } \text{N}_2 \times \frac{2 \text{ moles } \text{NH}_3}{1 \text{ mole } \text{N}_2} \\ \text{moles } \text{NH}_3 &= 8 \text{ moles} \end{aligned}$$

-Simplify the number of moles by multiplying each final no. of moles of reagent by the proportion of the initial number of moles and given reagent

-The reagent with lesser number of moles of NH_3 is the limiting reagent and vice versa; in this case, H_2 is the limiting reagent and N_2 is the excess reagent



Calculate the excess

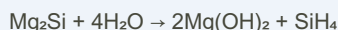
-To determine how much of 4 moles of N₂ is in excess, use mole-mole factor of N₂ and H₂

~~6 moles of H₂~~ x (1 mole N₂)/(3 moles H₂) = 2 moles N₂ in excess

-The number of moles of N₂ required to react with 6 moles of H₂ is only 2, thus, **6 moles of N₂ has an excess of 4 moles**

6 moles of N₂ - 2 moles of N₂ in 6 moles of H₂ = 4 moles excess N₂

Limiting and Excess Reagent with Molar Mass



If we start with 50.0 g of each reactant, how much in grams SiH₄ can be formed?

~~50g Mg₂Si~~ x (1 mol Mg₂Si)/(76.7g Mg₂Si) x (1 mol SiH₄)/(1 mol Mg₂Si) x (32.1 g SiH₄)/(1 mol SiH₄) = **20.9 SiH₄**

~~50g H₂O~~ x (1 mole H₂O)/(18.0g H₂O) x (1 mol SiH₄)/(4 mol H₂O) x (32.1 g SiH₄)/(1 mol SiH₄) = **22.3 g SiH₄**

50g Mg₂Si = 20.9 SiH₄ (*Limiting reactant*)

50 g H₂O = 22.3 g SiH₄ (*Excess reactant*)

-Divide the number of initial molar mass of compound by the final/given molar mass and multiply with the molar mass of required compound to convert

-Determine how much of 50 g of H₂O is in excess by 50 g H₂O x (1 mol Mg₂Si)/(4 mol H₂O) = 12.5 g Mg₂Si

50 g Mg₂Si - 12.5 g Mg₂Si *excess in 50 g H₂O* = **37.5 g excess Mg₂Si**

Theoretical and Percent Yield

Percent Yield - ratio of actual yield to the theoretical yield expressed as a percentage

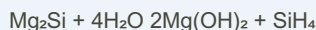
$$\text{Percent Yield} = (\text{Actual Yield})/(\text{Theoretical Yield}) \times 100\%$$

Theoretical Yield - maximum/expected amount of product produced from the given amount of reactant

Actual Yield - actual amount of product produced from the given amount of reactant (determined experimentally)

Calculating Theoretical and Percent Yield

In calculating theoretical yield, always use the limiting reactant. *from the previous example*



If 19.87 g SiH₄ is formed, what is the percent yield of the reaction?

50g Mg₂Si = 20.9 SiH₄ (Limiting reactant) = Theoretical yield

$$\text{Percent Yield} = (\text{Actual Yield})/(\text{Theoretical Yield}) \times 100\%$$

$$= 19.87 \text{ g SiH}_4/20.9 \text{ SiH}_4 \times 100\% = \mathbf{94.89\%}$$

Calculating Theoretical and Percent Yield (cont)

$$\text{Percent error} = 100\% - 98.89\% = \mathbf{5.11\%}$$

Module 5 - Gases I

Pressure - amount of force exerted per unit area

Standard atmosphere (atm) - widely used unit for pressure; **1 atm = 760mmHg**

Torr (or mmHg) - milliliter of mercury equal to 1 atmosphere; named after Italian scientist Evangelista Torricelli (invented barometer)

Pounds per square inch (psi) - amount of pressure in pounds that gas exerts in a container per square inch of unit area

kilopascal (kPa) - equal to 1000 Pa, modern unit for pressure/default

Conversion Factor:

$$\mathbf{1 \text{ atm} = 760\text{mmHg} = 760 \text{ Torr} = 101.3 \text{ kPa} = 14.7 \text{ psi} = 1 \text{ k Pa} = 1000 \text{ Pa}}$$

Gas Laws

| | | |
|-------------|--|-------------------|
| Boyle's Law | Volume is inversely proportional to its pressure at a constant temperature ; $\uparrow V \iff \downarrow P$ | $P_1V_1 = P_2V_2$ |
|-------------|--|-------------------|

| | | |
|--------------|---|---------------------|
| Charles' Law | Volume is directly proportional to its absolute temperature and constant pressure ; $\uparrow V \iff \uparrow T$ | $V_1/T_1 = V_2/T_2$ |
|--------------|---|---------------------|

| | | |
|----------------|---|---------------------|
| Avogadro's Law | Volume is directly proportional to the number of moles contained in the volume at constant temperature and pressure ; $\uparrow V \iff \uparrow n$ | $V_1/n_1 = V_2/n_2$ |
|----------------|---|---------------------|



Gas Laws (cont)

| | | |
|--------------------------------|---|---|
| Gay-Lussac's Law/Ideal Gas Law | Sums up and combines Boyle's, Charles', and Avogadro's Laws | $PV = nRT$ (where R or universal gas constant = 0.0821 atm.L/mol.K) |
|--------------------------------|---|---|

Gas Mixtures

Different gases can be present in a container and can be represented as n_1 (gas 1), n_2 (gas 2), or n_3 (gas 3), etc.; the total number of moles as n_{total}

The pressure exerted by the mixture can be interpreted as $P_{mixture} = (n_{total}RT)/V$

It can be simplified as $P_1 = (n_1RT)/V$; $P_2 = (n_2RT)/V$; $P_3 = (n_3RT)/V$

Pressures P_1 , P_2 , and P_3 are partial pressure of each gas

Dalton's Law of Partial Pressure - pressure exerted by the mixture is the sum of the pressures exerted by each component

Get the partial pressure of gas 1 by $P_1 = P_{mixture} X_1$ (wherein X_1 is the mole fraction of gas 1)

Module 6 - Gases II

Stoichiometric ratio - dictates the ratio of components to start the reaction

Standard Temperature and Pressure (STP) = 0°C (273 K) and pressure of 1 atm

Amount of gaseous products are determined using $n = (PV_{STP})/(RT)$ wherein V_{STP} is the volume of gases involved measured in STP in liters (L)

$$n = (PV_{STP})/(0.0821)(273 \text{ K}) = V_{STP}/22.4$$

$$n = V_{STP}/22.4$$

Gases are also measured in *Standard Ambient Temperature and Pressure (SATP)* which is more accurate than STP which is at 25°C (298 K) and 1 atm.

$$n = (PV_{SATP})/(0.0821)(298 \text{ K}) = V_{SATP}/24.5$$

$$n = V_{SATP}/24.5$$

Temperature Conversion

| | | | |
|------------|---|------------|--------------------------|
| Celsius | → | Kelvin | $C + 273.15$ |
| Celsius | → | Fahrenheit | $C(9/5) + 32$ |
| Fahrenheit | → | Kelvin | $(F - 32)(5/9) + 273.15$ |
| Fahrenheit | → | Celsius | $(F - 32)(5/9)$ |
| Kelvin | → | Celsius | $K - 273.15$ |

Temperature Conversion (cont)

| | | | |
|--------|---|------------|--------------------------|
| Kelvin | → | Fahrenheit | $(K - 273.15)(9/5) + 32$ |
|--------|---|------------|--------------------------|

Kinetic Molecular Theory

1. Gases are very small molecules separated by expansive space between them
2. Force of attraction between particles is **negligible**
3. The molecules are in **constant motion** and move randomly in all directions
4. Sometimes particles **collide** with each other or with the walls of container
5. The collisions are **perfectly elastic**, hence, there is no change in momentum
6. The average kinetic energy is determined only by the absolute temperature of the gas

To determine the kinetic energy of gas particles, the root-mean-square velocity is used:

$$v_{rms} = \sqrt{(3RT/M)}$$

where R = ideal gas constant, T = absolute temperature in K, M = molar mass in g/mol

To compare the velocities of gases with different molar masses at the same absolute temperature:

$$v_{rms1}/v_{rms2} = \sqrt{(M_2)/\sqrt{(M_1)}}$$

where M_1 or M_2 = molar mass of gas 1 or 2

This expression is also known as **Graham's Law of Diffusion** which states that the diffusion rate (rate at which the gas moves), is **inversely proportional** to the square root of its molar mass

YEY! you finished q1, I am so proud of you :)

