

### 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

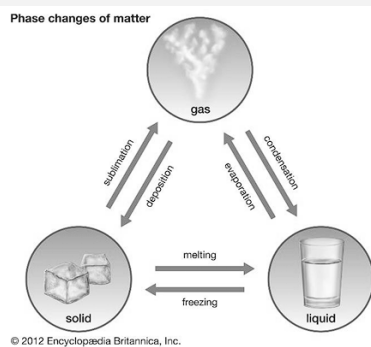


Figure 1.1. Phase Changes of Matter

### 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 <sub>2</sub> O), Salt (NaCl), Ammonia (NH <sub>3</sub> )

### 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 <sup>+</sup> (1A) and Cl <sup>-</sup> (7A) creates a stable bond (octet rule)
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By **Jerstellar**  
[cheatography.com/jerstellar/](https://cheatography.com/jerstellar/)

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### Types of bonds (cont)

Covalent atoms share electrons;  $O^{2-}(6A)$  and 2 atoms of nonmetals  $H^{+}(1A) = H_2O$

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  ${}^A_Z E$

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 carbons 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 -  $Ca_3(PO_4)_2$

### Some Polyatomic Ions

Monovalent ( $1^-$ )		Bivalent ( $2^-$ )		Trivalent ( $3^-$ )	
Ammonium	$NH_4^+$	Carbonate	$CO_3$	Phosphate	$PO_4$

### Some Polyatomic Ions (cont)

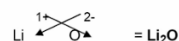
Acetate	$C_2H_3O_2$	Chromate	$CrO_4$	Borate	$BO_3$
Chlorate	$ClO_3$	Oxalate	$C_2O_4$		
Chlorite	$ClO_2$	Sulfate	$SO_3$		
Bicarbonate	$HCO_3$	Sulfite	$SO_2$		
Biculfate	$HSO_4$	Peroxide	$O_2$		
Hydroxide	$OH$				
Nitrate	$NO_3$				
Nitrite	$NO_2$				

### Diatomic Molecules

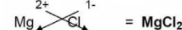
$H_2$	hydrogen
$N_2$	nitrogen
$F_2$	fluorine
$O_2$	oxygen
$I_2$	iodine
$Cl_2$	chlorine
$Br_2$	bromine

### Criss-Cross Method

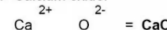
a. Lithium oxide



b. Magnesium chloride



c. Calcium oxide:



- Determine the charge or valence number of the elements
- Exchange their valence numbers
- Reducing by their gcd is possible

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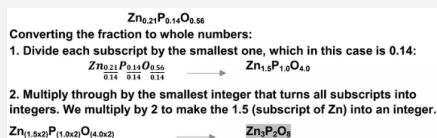
### Calculating Empirical Formula

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



### 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

$$\Sigma = 260$$

Suppose the molar mass is 520.8; divide it by the summation ( $520.8/260 \approx 2$ ). 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

$\Sigma = 18.02$

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



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[cheatography.com/jerstellar/](https://cheatography.com/jerstellar/)

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### 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

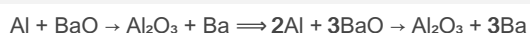
Product - substance formed in a chemical reaction

*Reactants* → *Products*

"to yield" or "to form" (→)

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

### Examples



### Types of Chemical Reactions

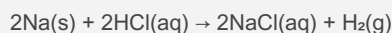
Type	Definition	Example
Combination/-Synthesis	two or more reactants combine to form a single product	$2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$
Decomposition	one reactant breaks down into two or more products	$\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
Single Displacement	one element 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

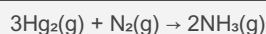
Stoichiometric coefficient - added before an element, ion, or molecule to balance chemical reactions

Mole method using mole-mole factor:



2 moles Na  $\equiv$  2 moles NaCl; hence,

### Calculating Amount of Product and Reactant



How many moles of  $\text{H}_2$  are needed to produce **26.5 moles of  $\text{NH}_3$** ?

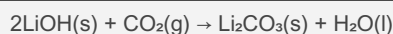
$$\cancel{26.5 \text{ moles NH}_3} \times (3 \text{ moles H}_2) / \cancel{(2 \text{ moles NH}_3)} = 39.8 \text{ moles of H}_2$$

How many moles of  $\text{NH}_3$  will be produced if **33.7 moles of  $\text{N}_2$**  reacts completely with  $\text{H}_2$

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

-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  $\text{CO}_2$  can be absorbed by **236.1 g of LiOH**?

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

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

$$(\text{Li}=6.941, \text{O}=15.999, \text{H}=1.008, \text{C}=12.011)$$

-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  $\text{H}_2$  was mixed with 4 moles of  $\text{N}_2$ . To determine which is the limiting reagent, the amount of  $\text{NH}_3$  must be computed given the moles of  $\text{H}_2$  and  $\text{N}_2$  and the mole-mole factor of the equation

### Solution

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

-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  $\text{NH}_3$  is the limiting reagent and vice versa; in this case,  $\text{H}_2$  is the limiting reagent and  $\text{N}_2$  is the excess reagent

### Calculate the excess

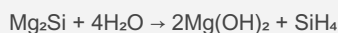
-To determine how much of 4 moles of  $\text{N}_2$  is in excess, use mole-mole factor of  $\text{N}_2$  and  $\text{H}_2$

$$6 \text{ moles of H}_2 \times (1 \text{ mole N}_2)/(3 \text{ moles H}_2) = 2 \text{ moles N}_2 \text{ in excess}$$

-The number of moles of  $\text{N}_2$  required to react with 6 moles of  $\text{H}_2$  is only 2, thus, **6 moles of  $\text{N}_2$  has an excess of 4 moles**

$$6 \text{ moles of N}_2 - 2 \text{ moles of N}_2 \text{ in 6 moles of H}_2 = 4 \text{ moles excess N}_2$$

### Limiting and Excess Reagent with Molar Mass



If we start with 50.0 g of each reactant, how much in grams  $\text{SiH}_4$  can be formed?

$$50 \text{ g Mg}_2\text{Si} \times (1 \text{ mol Mg}_2\text{Si})/(76.7 \text{ g Mg}_2\text{Si}) \times (1 \text{ mol SiH}_4)/(1 \text{ mol Mg}_2\text{Si}) \times (32.1 \text{ g SiH}_4)/(1 \text{ mol SiH}_4) = 20.9 \text{ SiH}_4$$

$$50 \text{ g H}_2\text{O} \times (1 \text{ mole H}_2\text{O})/(18.0 \text{ g H}_2\text{O}) \times (1 \text{ mol SiH}_4)/(4 \text{ mol H}_2\text{O}) \times (32.1 \text{ g SiH}_4)/(1 \text{ mol SiH}_4) = 22.3 \text{ g SiH}_4$$

$$50 \text{ g Mg}_2\text{Si} = 20.9 \text{ SiH}_4 \text{ (Limiting reactant)}$$

$$50 \text{ g H}_2\text{O} = 22.3 \text{ g SiH}_4 \text{ (Excess reactant)}$$

-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

$$\text{-Determine how much of 50 g of H}_2\text{O is in excess by } 50 \text{ g H}_2\text{O} \times (1 \text{ mol Mg}_2\text{Si})/(4 \text{ mol H}_2\text{O}) = 12.5 \text{ g Mg}_2\text{Si}$$

$$50 \text{ g Mg}_2\text{Si} - 12.5 \text{ g Mg}_2\text{Si} \text{ excess in } 50 \text{ g H}_2\text{O} = 37.5 \text{ g excess Mg}_2\text{Si}$$

### Theoretical and Percent Yield

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

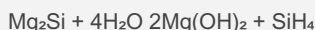
$$\text{Percent Yield} = (\text{Actual Yield})/(\text{Theoretical Yield}) \times 100\%$$

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 reagent. *from the previous example*



If 19.87 g  $\text{SiH}_4$  is formed, what is the percent yield of the reaction?

$$50 \text{ g Mg}_2\text{Si} = 20.9 \text{ SiH}_4 \text{ (Limiting reactant) = Theoretical yield}$$

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

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

$$\text{Percent error} = 100\% - 94.89\% = 5.11\%$$

### 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:

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

### Gas Laws

Boyle's Law	Volume is inversely proportional to its pressure at a constant temperature; $\uparrow V \leftrightarrow \downarrow P$	$P_1V_1 = P_2V_2$
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### Gas Laws (cont)

Charles' Law	Volume is directly proportional to its absolute temperature and <b>constant pressure</b> ; $\uparrow V \Leftrightarrow \uparrow T$	$V_1/T_1 = V_2/T_2$
Avogadro's Law	Volume is directly proportional to the number of moles contained in the volume at <b>constant temperature and pressure</b> ; $\uparrow V \Leftrightarrow \uparrow n$	$V_1/n_1 = V_2/n_2$
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 = <b>0.0821 atm.L/mol.K</b> )

### 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	$K - 273.15$
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



### Kinetic Molecular Theory (cont)

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 :)



### Calculating Empirical Formula with Molar mass

Given	Mass number	Quotient	Quotient (x/0.75)	Smallest Ratio (y*2)
28.03%	24.035	1.15	1.5	3
1.6%	28.086	0.75	1	2
1.16%	1.008	1.15	1.5	3
49.21	15.999	3.08	4	8

Answer =  $Mg_3Si_3H_2O_8$

- Divide the given percent composition to the mass number of each element
- Divide each quotient by the smallest number among them
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Given	Mass number	Quotient	Quotient (x/0.75)	Smallest Ratio (y*2)
28.03%	24.305	1.15	1.53≈1.5	3
Mg				
21.6% Si	21.16	1.15	1.53≈1.5	3
1.16% H	1.16	0.75	1	2
49.21%	49.21	3.08	4.1≈4	8
O				

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

