GenChem q1 module Cheat Sheet

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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

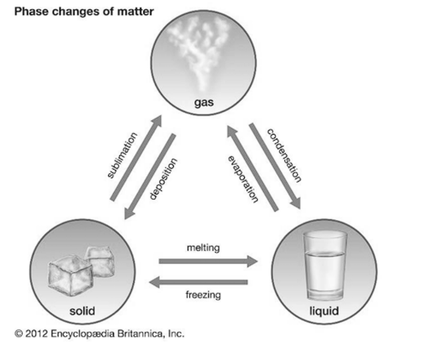


Figure 1.1. Phase Changes of Matter

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 (H20), Salt (NaCl), Ammonia (NH3)

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
Intrinsive	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
	Hetero-genous	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_2O

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 = nuclear charge = number of protons = 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}zE
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 occuring carbon atoms have 6 neutrons (A=12)

Chemical Compounds

Radicals/Polyatomic Ions - stable groups which form chemical bonds as an intact unit.
The valence numbe 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
Acetate	$C_2H_3O_2$	Chromate	CrO_4	Borate	BO_3
Chlorate	ClO_3	Oxalate	C_2O_4		
Chlorite	ClO_2	Sulfate	SO_3		

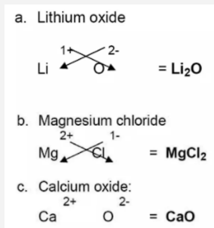
Some Polyatomic Ions (cont)

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 gcf is possible

Calculating Empirical Formula

Percentage Composition - amounts of the elements for a given amount of compound
Empirical Formula - simpest 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:

$Zn_{\frac{0.21}{0.14}}P_{\frac{0.14}{0.14}}O_{\frac{0.56}{0.14}} \rightarrow Zn_{1.5}P_{1.0}O_{4.0}$

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.

$Zn_{(1.5x2)}P_{(1.0x2)}O_{(4.0x2)} \rightarrow Zn_3P_2O_8$

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 compostion 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

$CH_4 + O_2 \rightarrow CO_2 + H_2O \Longrightarrow CH_4 + \mathbf{2}O_2 \rightarrow CO_2 + \mathbf{2}H_2O$

$Al + BaO \rightarrow Al_2O_3 + Ba \Longrightarrow \mathbf{2}Al + \mathbf{3}BaO \rightarrow Al_2O_3 + \mathbf{3}Ba$

$Cl_2 + KBr \rightarrow KCl + Br_2 \Longrightarrow Cl_2 + \mathbf{2}KBr \rightarrow \mathbf{2}KCl + Br_2$

Types of Chemical Reactions

Type	Definition	Example
Combination/-Synthesis	two or more reactants combine to form a single product	$2Mg + O_2 \rightarrow 2MgO$
Decomposition	one reactant breaks down into two or more products	$CaCO_3 \rightarrow CaO + CO_2$
Single Displa-cement	one element is substituted for another element in a compound	$K + NaCl \rightarrow KCl + Na$
Double Displa-cement/salt metathesis	two substances react by exchanging ions to produce two new molecules	$AgNo_3 + NaCl \rightarrow AgCl + NaNo_3$

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:
$2Na(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2(g)$
2 moles Na ≅ 2 moles NaCl; hence,

Calculating Amount of Product and Reactant

$3Hg_2(g) + N_2(g) \rightarrow 2NH_3(g)$
How many moles of **H_2** are needed to produce **26.5 moles of NH_3**?
~~26.5 moles NH_3~~ x (3 moles H_2)/~~(2 moles NH_3)~~ = 39.8 moles of 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 moles of N_2~~ x (2 moles of NH_3)/~~(1 mole of N_2)~~ = 67.4 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

$2LiOH(s) + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O(l)$
How many grams of **CO_2** can be absorbed by **236.1 g of LiOH**?
~~236.1g LiOH~~ x (1 mole of LiOH)/~~(23.95g LiOH)~~ x ~~(1 mole CO_2)~~ /(2 moles LiOH) x (44.01g of CO_2)/ ~~(1 mole CO_2)~~ = 221.1g 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

$3H_2 + 2N_2 \rightarrow 2NH_3$
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

$$moles\ NH_3 = \#\ moles\ of\ H_2\ X\ \frac{2\ moles\ NH_3}{3\ moles\ H_2}$$

$$moles\ NH_3 = 6\ \cancel{moles}\ of\ H_2\ X\ \frac{2\ moles\ NH_3}{3\ \cancel{moles}\ H_2}$$

$$moles\ NH_3 = 4\ moles$$

$$moles\ NH_3 = \cancel{moles}\ of\ N_2\ X\ \frac{2\ moles\ NH_3}{1\ \cancel{mole}\ N_2}$$

$$moles\ NH_3 = 4\ \cancel{moles}\ of\ N_2\ X\ \frac{2\ moles\ NH_3}{1\ \cancel{mole}\ N_2}$$

$$moles\ NH_3 = 8\ moles$$

-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 vise 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 N_2 is in excess, use mole-mole factor of N_2 and H_2
~~6 moles of H_2~~ x (1 mole N_2)/(~~3 moles H_2~~) = 2 moles N_2 in excess
-The number of moles of N_2 required to react with 6 moles of H_2 is only 2, thus, **6 moles of N_2 has an excess of 4 moles**
6 moles of N_2 - 2 moles of N_2 in 6 moles of H_2 = 4 moles excess N_2

Limiting and Excess Reagent with Molar Mass

$Mg_2Si + 4H_2O \rightarrow 2Mg(OH)_2 + SiH_4$
If we start with 50.0 g of each reactant, how much in grams SiH_4 can be formed?
~~50g Mg_2Si~~ x (1 mol Mg_2Si)/(~~76.7 g Mg_2Si~~) x (~~1 mol SiH_4~~)/(1 mol Mg_2Si) x (32.1 g SiH_4)/(~~1 mol SiH_4~~) = **20.9 SiH_4**
~~50 g H_2O~~ x (1 mole H_2O)/(~~18.0 g H_2O~~) x (~~1 mol SiH_4~~)/(4 mol H_2O) x (32.1 g SiH_4)/(~~1 mol SiH_4~~) = **22.3 g SiH_4**
50g Mg_2Si = 20.9 SiH_4 *(Limiting reactant)*
50 g H_2O = 22.3 g SiH_4 *(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 H_2O is in excess by 50 g H_2O x (1 mol Mg_2Si)/(4 mol H_2O) = 12.5 g Mg_2Si
50 g Mg_2Si - 12.5 g Mg_2Si *excess in 50 g H_2O* = **37.5 g excess Mg_2Si**

Theoretical and Percent Yield

Percent Yield - ratio of actual yield to the theoretical yield expressed as a percentage
Percent Yield = (Actual Yield)/(Theoretical Yield) x 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*
$Mg_2Si + 4H_2O\ 2Mg(OH)_2 + SiH_4$
If 19.87 g SiH_4 is formed, what is the percent yield of the reaction?
50g Mg_2Si = 20.9 SiH_4 (Limiting reactant) = Theoretical yield
Percent Yield = (Actual Yield)/(Theoretical Yield) x 100%
= 19.87 g SiH_4/20.9 SiH_4 x 100% =**94.89%**

Calculating Theoretical and Percent Yield (cont)

Percent error = 100% - 98.89% = **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 atm = 760mmHg = 760 Torr = 101.3 kPa = 14.7 psi = 1k Pa = 1000 Pa

Gas Laws

Boyle's Law	Volume is inversely proportional to its pressure at a **constant temparature**; ↑V ⟺ ↓P	$P_1V_1 = P_2V_2$
Charles' Law	Volume is directly proportional to its absolute temperature and **constant pressure**; ↑V ⟺ ↑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**; ↑V ⟺ ↑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 $\mathbf{n_{total}}$

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

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

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 $\mathbf{P_1 = P_{mixture} X_1}$ (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

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

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

$\mathbf{n = (PV_{stp})/(0.0821)(273\ K) = V_{stp}/22.4}$

$\mathbf{n = V_{stp}/22.4}$

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.

$\mathbf{n = (PV_{satp})/(0.0821)(298\ K) = V_{satp}/24.5}$

$\mathbf{n = V_{satp}/24.5}$

Temperature Conversion

Celsius	→	Kelvin	C + 273.15
Celsius	→	Farenheight	C (9/5) + 32
Farenheight	→	Kelvin	(F - 32)(5/9) + 273.15
Farenheight	→	Celsius	(F - 32)(5/9)
Kelvin	→	Celsius	K - 273.15

Temperature Conversion (cont)

Kelvin	→	Farenheight	(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 molcules 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 temparature of the gas

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

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

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:

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

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 :)