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CHEMISTRY POINTERS
BASIC CONCEPTS AND LAWS

Chemistry

Chemistry is the physical science that deals with the composition, structure, and properties of substances and also the transformations that these substances undergo. Because the study of chemistry encompasses the entire material universe, it is central to the understanding of other sciences. Several branches of Chemistry are the following:

a. ***Organic chemistry*** – mostly concerned with the study of chemicals containing the element carbon
b. ***Inorganic chemistry*** – the study of all elements of compounds other than organic compounds
c. ***Analytical Chemistry*** – the study of qualitative and quantitative analysis of elements and compounds.
d. ***Physical Chemistry*** – the study of reaction rates, mechanisms, bonding and structure

Matter

Matter in science, is a general term applied to anything that has the property of occupying space and the attributes of gravity and inertia.

STATES OF MATTER

Plasma **is** the collection of charged gaseous particles containing nearly equal numbers of negative and positive ions, is sometimes called the fourth state of matter.

Solid **is** characterized by resistance to any change in shape, caused by a strong attraction between the molecules of which it is composed.

Liquids have sufficient molecular attraction to resist forces tending to change their volume.

Gas molecules are widely dispersed and move freely, offers no resistance to change of shape and little resistance to change of volume.

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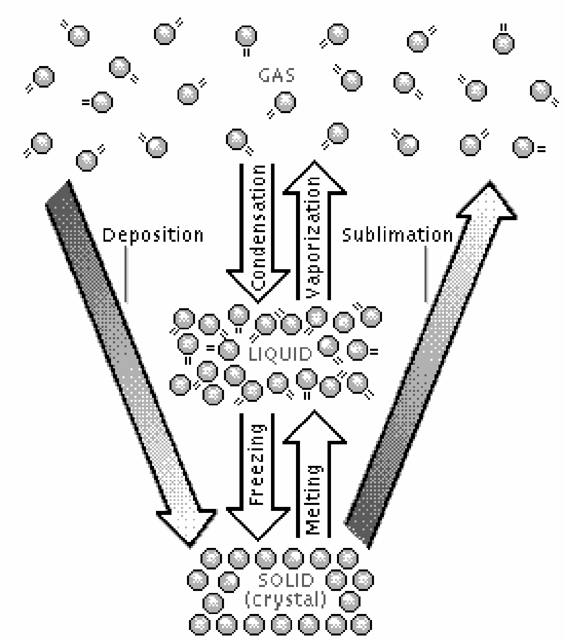
CHANGES IN MATTER

***Physical Change* is** characterized by a change in the phase or state of a substance. Some physical properties of the substance are altered, but its chemical composition remains unchanged. Ex. phase change

***Chemical change* is** characterized by a change in composition of its molecules changes. The properties of the original substance are lost, and new substances with new properties are produced. An example of a chemical change is the production of rust (iron oxide) when oxygen in the air reacts with iron.

Phase Changes



PHASE TRANSITION	NAME	EXAMPLES
Solid → liquid	Melting, fusion	Melting of snow and ice
Solid → gas	Sublimation	Sublimation of dry ice, freeze-drying of coffee
Liquid → solid	Freezing	Freezing of water or a liquid metal
Liquid → gas	Vaporization	Evaporation of water or refrigerant
Gas → liquid	Condensation, liquefaction	Formation of dew, liquefaction of carbon dioxide
Gas → solid	Condensation, deposition	Formation of frost and snow

Phase Transitions

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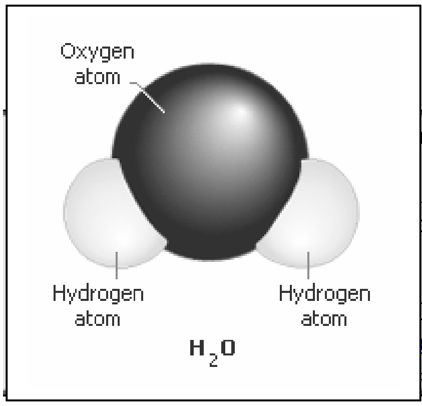
Matter exists in various forms, or phases. If the temperature and/or pressure of a sample of matter is adjusted, the matter may undergo a phase transition. During a phase transition, matter shifts between its three states: solid, liquid, and gas.

Elements and Compounds

An *Element* is a substance that cannot be broken down into simpler substances by ordinary means. Ninety elements are known to occur in nature, and 22 more have been made artificially. Out of this limited number of elements, all the millions of known substances are made.

Compounds are substance in which two or more elements joined by chemical bonds. A compound can be created or broken down by means of a reaction but not by mechanical or physical techniques



Water Molecule

Water is an example of a compound. A water molecule consists of an oxygen atom and two hydrogen atoms.

Salt, water, iron rust, and rubber are examples of compounds

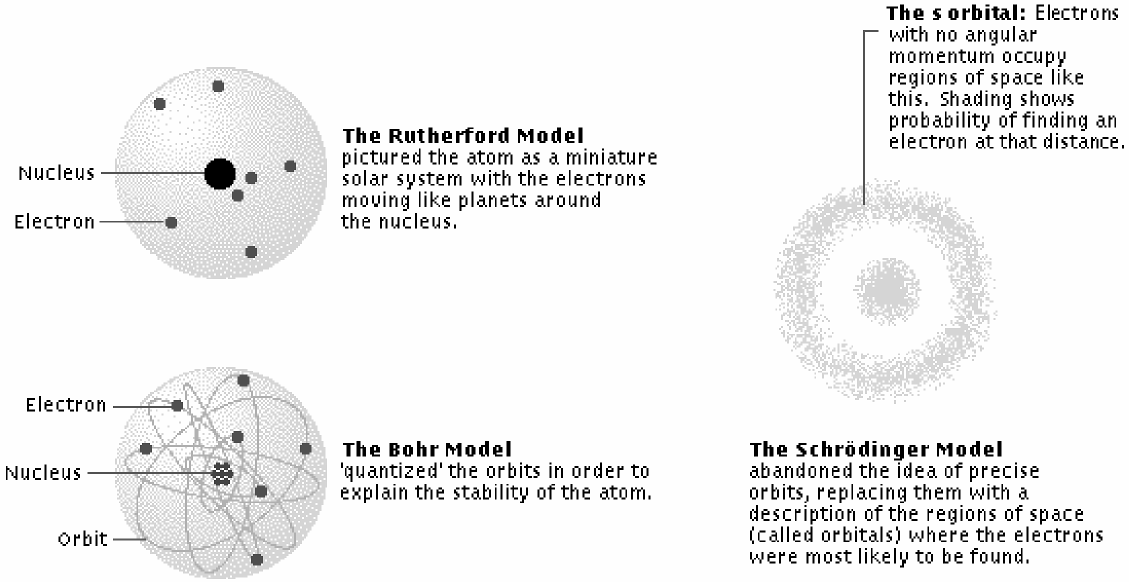
Atoms and Molecules

An ***atom*** is the smallest unit of an element that has the properties of the element; a ***molecule*** is the smallest unit of a compound or the form of an element in which atoms bind together that has the properties of the compound or element.

SUB-ATOMIC PARTICLES

Nucleus - is very small compared with the rest of the atom and contains most of the atomic mass (or weight). The nucleus is about 10^{-12} cm (3.94 x 10^{-13} in) in diameter. The size ratio of the atom to the nucleus is 10,000 to 1.
Electrons - is about 10^{-8} cm (3.94 x 10^{-9} in). It carries a negative electric charge with an assigned value of -1. The atom is determined by the size of this electron cloud.
***Proton* -** carries a positive electric charge with an assigned value of +1. The mass of a proton is 1836 times the mass of an electron.
Neutron - has nearly the same mass as the proton, but the neutron has no electric charge.
***Mass Number* -** The total number of protons and neutrons in a nucleus.
Atomic number- equals to the number of proton in an electron
***Isotopes* -** atoms of the same element having the same number of protons but different numbers of neutrons. The term *isotope* (from the Greek word meaning "same place") defines atoms that have the same number of protons but a different number of neutrons. That is, they are atoms of the same element that have different masses. (Ex. $7H_1$, $7H_1$, $H7_1$, H_1)
***Atomic Weight* -** the average weight (more correctly, the mass) of an atom of an element, taking into account the masses of all its isotopes and the percentage of their occurrence in nature.

MODELS OF THE ATOM



An atom of an element is denoted by $^{a}X_{b}$ where X corresponds to the nucleus of the atom (name of the atom, e.g.C), is the mass number and b, the atomic number. Ex. $^{12}C_{6}$

Element	Number of Protons	Number of electrons	Number of neutrons	Atomic number	Mass number
Carbon	6	6	6	6	12

APPLICATION:

Complete the table below.

Element	Number of Protons	Number of electrons	Number of neutrons	Atomic number	Mass number
Na					
Na^+					
Cl^-					

ELECTRON CONFIGURATION

The electron configuration of an atom is the arrangement of the atom's electrons with respect to its nucleus.

An electron may occupy a certain energy level (n). An orbital is generally visualized as a cloud with a specified size and shape determined, in general, by the energy level of the electron. Valence electrons are electrons found in the highest energy level of the electron cloud.

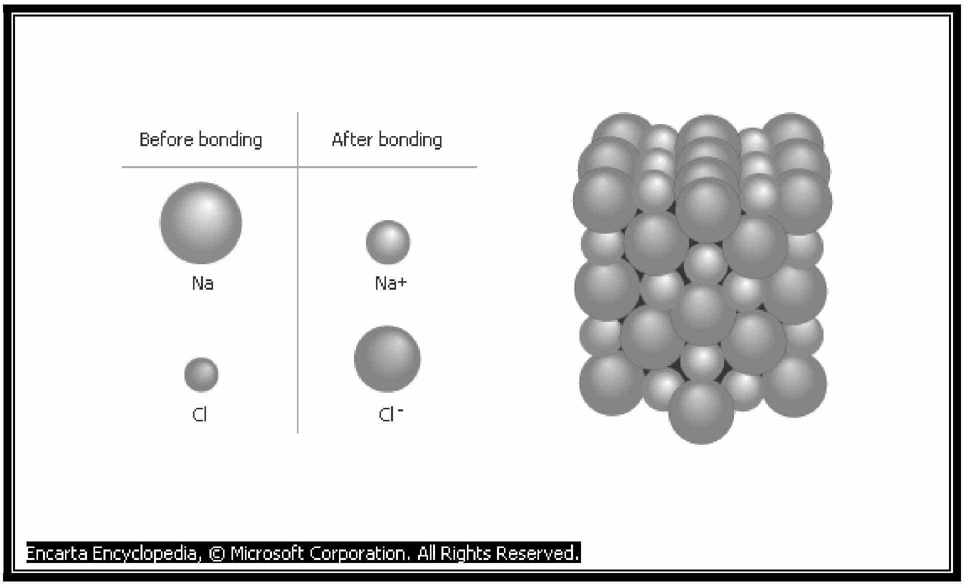
Given the electronic configuration, one can determine its position in the periodic table. The energy level corresponds to the period of the periodic table (1-7) while the number of valence electrons corresponds to the group number of that element. Example, the electron configuration of Na is 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p , the orbital number is while the valence number is 1. Therefore, Sodium is in group I period .

Chemical Bonds, Formulas, and Equation

Elements that do not have a noble-gas configuration (a stable configuration) try to attain such a configuration by entering into chemical reactions. Stable molecules are formed when atoms combine so as to have outer shells holding eight electrons.

Ionic Bonds

In the ionic model, electrons are transferred from one atom to another to achieve noble gas configuration. An ionic bond is formed. The atom giving up the electrons become positively charge (cation), while the atom accepting the electrons becomes negatively charged (Anion). Ionic bonds are formed when elements in Group IA to IIA (except hydrogen) combines with elements in group VIA to VIIA of the periodic table. Ionic solids form crystals. Cations and anions in crystals are arranged in a repeated fixed manner (crystal structure).



Covalent bonds

When sharing of available outer elements of the atoms occur, a covalent bond is formed. There is no electron transfer in covalent bonding. There can be multiple covalent bonds between two atoms. There can be a double bond or a triple bond.

Chemical Reactions

The reaction is the heart of the study of chemistry. All chemical reactions involve the breakage and reformation of chemical bonds of molecules to form different substances. Chemical reactions can be expressed through equations that resemble mathematical equations. The reactants (the substances that are combined to react with one another) appear on the left side of the equation, and the products (substances produced by the reaction) are written on the right side of the equation. The reactants and products are typically connected by an arrow or various types of double arrows. The single arrow shows that a reaction only proceeds in the direction indicated, while the double arrow indicates that a reaction can proceed in either direction (that products are also reacting with each other to reform reactants).

Ex. $Ca(HCO_3)_2(aq) \xrightleftharpoons{\Delta} CaCO_3(s) + CO_2(g) + H_2O$

TYPES OF CHEMICAL REACTIONS

A. ***Composition Reaction*** (synthesis or combination reaction) is a type of chemical reaction where a more complex substance is broken down into two or more simpler substances.
General Form: $A + X \rightarrow AX$
e.g. $2H_2 + O_2 \rightarrow 2H_2O$
$Fe + S \rightarrow FeS$

B. ***Decomposition Reaction*** is a type of chemical reaction where a more complex substance is broken down into two or more simpler substances.
General Form: $AX \rightarrow A + X$
e.g. $CaCO_3 \rightarrow CaO + CO_2$

There are five classes if decomposition, namely:
1. Decomposition of a metallic carbonate
2. Decomposition of a metallic hydroxide
3. Decomposition of a metallic chlorate
4. Decomposition of some acids
5. Decomposition of metallic oxides

C. ***Single Replacement Reaction*** is a type of chemical reaction in which a less reactive element is displaced from a compound by a more reactive element.

General Form: $A + BX \rightarrow AX + B$
$Y + BX \rightarrow BY + X$

There are four classes of replacement reactions. They are as given below:
1. Replacement of a less active metal from a compound by a more active metal.
2. Replacement of a less active nonmetal from a compound by more active non-metal.
3. Replacement of hydrogen from water by metals

4. Replacement of hydrogen from an acid by a more reactive metal.

D. ***Double Displacement reaction*** is a type of reaction wherein cations of two compounds switch anions to form new products.

General Form : AX + BY → AY + BX

e.g. NaOH + HCl → NaCl + H_2O

EXOTHERMIC AND ENDOTHERMIC REACTIONS

A chemical reaction either absorbs or releases energy. The energy released is in the form of heat energy. An **exothermic** reaction is defined as a chemical reaction that releases energy. Most often, the energy released is in the form of heat or light. When a bomb explodes, a tremendous amount of light and heat energy is released. This is an example of an exothermic reaction. On the other hand, there are chemical reactions that absorb energy as they take place. This reaction is **endothermic**. Photosynthesis requires light from the sun to proceed the reaction. Thus, it is an example of endothermic reaction.

FACTORS AFFECTING THE RATE OF A CHEMICAL REACTION

The rate of chemical reaction is affectedly several factors:

a. ***the nature of the reactants*** – the more reactive the reactant is to another reactant, the faster the reaction will proceed
b. ***temperature***-generally, the higher the temperature, the faster the reaction will take place
c. ***concentration of the reactants***-the higher the concentration to react with another reactant, the faster the reaction will take place
d. ***surface area***- the smaller the surface area, the faster the reaction
e. ***effect of catalyst***-with the presence of a catalyst, the reaction will become faster

Factor	**Example**
Nature of reactants	Gold + water =no reaction Sodium +water = fast
Temperature	Evaporation is faster when it is hot
Concentration of reactants	Wood+dilute acid= slow Wood +concentrated acid=fast
Surface area	Crushed eggshells will dissolve fast in acid Whole eggshells will dissolve slow in acid

Presence of catalyst	Decomposition of formic acid = slow Decomposition of formic acid in the presence of sulfuric acid = fast

LAWS OF CHEMICAL COMBINATION

The laws of chemical combination were formulated in the early part of the 19th century. They are a result of the first use of quantitative measurement in chemistry.

In 1799, Joseph Proust proposed the **law of definite proportions** (also called the law of constant composition). The law states that compounds contain elements in certain fixed proportions and in no other combinations, regardless of the method of preparation. Thus, chalk, or calcium carbonate, CaCOO, is always 40% calcium, 12% carbon, and 48% oxygen, by weight.

In 1803, John Dalton articulated the **law of multiple proportions**. This law states that if two elements combine to form more than one compound, then the ratio of the weights of the second element (which combines with a fixed weight of the first element) will be small whole numbers. For example, carbon and oxygen can form two compounds, carbon monoxide and carbon dioxide. In carbon monoxide 12 g of carbon combine with 16 g of oxygen, and in carbon dioxide, the same weight of carbon combines with 32 g of oxygen. Thus, the oxygen weight ratio that combines with 12 g of carbon is (32/16), or 2. The **law of combining weights,** also proposed by Dalton, states that in every compound, the proportion by weight of each element in the compound may be expressed by the atomic weight or a multiple of the atomic weight of each element. (This law was discovered before the atomic theory was postulated and was thus worded more generally.) The law of combining weights can be seen to follow directly from the atomic theory. In the case of water, HMO, each molecule of water is composed of two atoms of hydrogen (atomic weight 1) and one atom of oxygen (atomic weight 16). Thus, all molecules of water consist of 2 parts of hydrogen and 16 parts of oxygen by weight. All other compounds can be analyzed similarly.

These three laws, proposed from the first use of quantitative experimental techniques, resulted (1803) in Dalton's atomic theory.

Balancing Equations

With the Laws governing chemical reactions, the numbers of atoms for each element should be equal in the reactant side and the product side. We should therefore balance a chemical reaction. A chemical equation is considered balance if there are exactly the same number of atoms for each element on both sides of the equation.

Stoichiometry

Stoichiometry is the branch of chemistry that deals with quantitative relationships between the reactants and products of a chemical reaction.

In the reaction

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

two moles of hydrogen is needed to react with one mole of oxygen to yield two moles of water.

$2H_2$	+	O_2	$\rightarrow$	$2H_2O$
2moles		1mole		2moles
2g/mol H_2		32g/mol O_2		18 g/mol H_2O
4grams		32 grams		36 grams

This means that 4 grams of hydrogen is needed to react completely with 32 grams of oxygen to produce 36 grams of water.

APPLICATION

a. Mole to Mole

Combustion of butane produces carbon dioxide and water. How many moles of water will be produced if 5 moles of butane is used in the reaction?

Step1: Write the balanced chemical equation.

$$2CH_4H_{10} \quad + \quad 13O_2 \quad \rightarrow \quad 8\ CO_2 \quad + \quad 10H_20$$

Step2: The balanced chemical reaction shows that for every 2 moles of butane, 10 moles of water will be produced. Using the relationship, multiply the number of moles of butane to the stoichiometric ratio.

$$5\text{moles } CH_4H_{10} \quad \times \quad \frac{10 \text{ moles } H_20}{2 \text{ moles } CH_4H_{10}} \quad = \quad 25\text{moles } \underline{H_20}$$

b. Mole to Mass

If 0.25 moles of sodium carbonate is heated, what mass of sodium oxide will be produced?

Step 1: Write the balanced equation.

$Na_2CO_3 \rightarrow Na_2O + CO_2$

Step2: Determine the molecular weight of Na_2O and analyze the problem.

$Na_2O = 22.99(2) + 16 = 61.98$ g/mol

Step 3: The balanced chemical equation shows that for every 1mole of sodium carbonate, 1mole of sodium oxide will be produced. Using this relationship, multiply 0.25 moles of sodium carbonate to the stoichiometric ratio.

$$0.25 \text{ moles } Na_2CO_3 \times \frac{\underline{1\text{mole } Na_2O}}{1\text{mol } Na_2CO_3} = 0.25 \text{ moles } \underline{Na_2O}$$

Step 4: Convert the answer in step 3 into required unit (mass) using the relationship m=n

0.25 moles $\underline{Na_2O}$ x 61.98g/mol $\underline{Na_2O}$ = 15.50 g $\underline{Na_2O}$

c. Mass to Mass

How many grams of oxygen will be produced if 85.16 g of sodium chlorate is heated?

Step1: Balance the chemical equation.

2 NaClO3 → 2 NaCl + 3O2

Step2: Determine the molecular weights of the compounds involved and analyze the problem.

2 NaClO3 → 2 NaCl + 3O2

MW = 22.00+ 35.45+ 16.00 (3) = 106.44g/mol; =2(16) =32 g/mole

Step 3: Convert given mass of sodium chlorate to moles.

Moles of sodium chlorate = $\frac{\text{85.16g sodium chlorate}}{\text{10g/mol sodium chlorate}}$ =

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= 0.8001 mol sodium chlorate

Step4: From the balanced chemical equation, 2moles of sodium chlorate heated will produce3molesof oxygen. Using this relationship, multiply the result in Step3 with this ratio.

$$0.8001\text{molsodium chlorate} \times \frac{3\text{molsoxygen}}{2\text{molssodiumchlorate}}$$

= 1.200 mol oxygen

Step 5: Convert moles of oxygen produced to mass.

$$1.200\text{molsoxygen} \times \frac{32.00\text{g oxygen}}{1\text{mol oxygen}} = 38.40 \text{ g oxygen}$$

d. Mass to mole

Lighting a candle produces 0.13 g of water. What is the mass of butane used? (Butane is the fuel used in most lighters)

Step 1: Balance the chemical reaction. Remember that the complete combustion reaction of a hydrocarbon will yield carbon dioxide and water.

$2CH_4H_{10} \quad + \quad 13O_2 \quad \rightarrow \quad 8\ CO_2 \quad + \quad 10H_20$

Step2: Analyze the problem.

$2CH_4H_{10} \quad + \quad 13O_2 \quad \rightarrow \quad 8\ CO_2 \quad + \quad 10H_20$

Given: 0.012g

MW = 1.008(2)+16
=18.02g/mol

Step 3: Convert given mass of water to moles

$$\text{Moles water} = \frac{0.012\text{g water}}{18.02\text{g/mol water}} = 0.000666 \text{ mol water}$$

Step4: From the balanced chemical equation, we have the following relationships:

2molesofbutane will produce 10 moles of water
2moles of butane burned will produce 8 moles of carbon dioxide

$$0.000666 \text{ mol of water} \quad \times \quad \frac{2\text{molsof butane}}{10\text{mols of water}} = 0.0001333 \text{ mol of butane}$$

Step 5: Convert mole to mass:

$$0.0001333\text{molbutane} \quad \times \quad \frac{58.12 \text{ g butane}}{1\text{mole butane}} = 0.0077 \text{ g butane}$$

Gases

Kinetic molecular Theory

Postulate 1: gases are made up of very tiny particles, called molecules. There are big empty spaces between the molecules of gases. The sizes of the molecules are very small compared to the distance between them, thus making them compressible. This also explains why gases have low density.

Postulate 2: gaseous molecules are in constant random motion. These particles are moving in straight lines at different speeds and direction. Since they are moving constantly, gases can easily occupy a large container. Thus, gases have no definite shape and volume and exhibits expandability. The random motion of the gaseous molecules explains the diffusibility and effusibility of gases. Diffusibility is the ability of a gas to scatter in space, while effusibility is the ability of a gas to escape through a small opening.

Postulate 3: The intermolecular forces of attraction between gaseous particles (molecules or atoms) are very weak. This attractive force between molecules was discovered by Johannes Diderik van der Waals and is called the Van der Waals force.

Postulate 4: The collision of gas particles with each other or with the walls of its container is perfectly elastic. Thus, no energy is lost upon collision. This means that the gas particles continue to move even if they collide with each other or with the container walls. The collision of molecules with the walls exerts pressure on the container.

Postulate 5: The average kinetic energy of the gas particles is directly proportional to the absolute temperature. As temperature increases, the average kinetic energy of the gas particles also increases and vice versa.

The Gas Laws

The laws governing the behavior of gases are termed collectively as the gas laws.

1. **BOYLE'S LAW**

Robert Boyle was the first scientist to measure the relationship between pressure and volume of gases (with temperature held constant).Boyle's Law states that the volume of a certain amount of dry gas held at constant temperature is inversely proportional to the pressure exerted by the gas. This statement is expressed mathematically as

APPLICATION:

What is the volume of a gas at 750 mmHg if it exerted pressure of 650mmHg at 700 mL? Assumethat temperature is held constant.)

Given:

V1 = ? V2= 700 mL

P_1 = 750 mmHg P_2= 650mmHg

$P_1V_1 = P_2V_2$

$$V_1 = \frac{P_2V_2}{P_1}$$

$$= \frac{(650 \text{ m Hg})\ (700 \text{ mL})}{750 \text{ mm Hg}} \quad = 685.7\text{mL}$$

2. **CHARLE'S LAW**

Alexandre Charles performed experiments on the relationship between the volume and temperature of gases. Charles' Law states that the volume of a certain amount of dry gas at constant pressure is directly proportional to its absolute temperature. This statement is represented as

$$V_1/ T_1 = V_2/ T_2$$

APPLICATION: The volume of a gas at 27 C is 400 mL. What will be the volume of that gas at 47 C, if the pressure is held constant?

Given:

V1	=	400 mL	V_2	=	?
T_1	=	27 C + 273	T_2	=	47 C + 273
	=	300 K		=	320

(Note that temperature in Celsius should be converted into Kelvin scale)

Solution:

$V_1 / T_1 = V_2 / T_2$

$$V_2 = \frac{V_1T_2}{T_1} = \frac{(380\text{mL})\ (320\text{K})}{300\text{ K}} = 405.33\text{ mL}$$

3. COMBINED GAS LAWS

Given a fixed amount of gas at two different conditions of pressure, temperature and volume, we derive the following relationship:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

APPLICATION:

The volume of a Gas at 37 C and 700 mm Hg is 500mL. What is the volume of the gas at 15 C and 600 mm Hg?

Given:

V1	=	500mL	V2	=	?
T1	=	37 C+ 273	T2	=	17 C+273
	=	310K		=	290 K
P_1	=	700 mmHg	P_2	=	600 mmHg

Solution:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1V_2}{T_1P_2}$$

$$= \frac{(700\text{mmHg})\ (500\text{mL})\ (290\text{ K})}{(310\text{ K})\ (600\text{ mm Hg})}$$

$$= 545.7\text{ mL}$$

4. AVOGADRO'S LAW

Amadeo Avogadro interpreted Gay-Lussac's findings on gas reactions at constant temperature and pressure. Avogadro's Law states that at a given temperature and pressure, the same volume of any two gases contain equal number of molecules. The molar volume of any gas at STP (Standard temperature and Pressure) is 22.4 L.

5. IDEAL GAS LAW

Boyle's Law, Charle's Law and Avogadro's Law can be combined such that the variables V,P,T and n are all equated to a single constant, R (the universal gas constant or ideal gas constant). The combined equation becomes:

$$\frac{PV}{nRT} = 1 \qquad \text{or } PV = nRT$$

where R = 0.0821 atm-L/molK

APPLICATION:

A sample of oxygen occupies 8.0 L of space at STP. How many moles of oxygen are present in the sample?

Given:

At STP, P= 1 atm T= 273 K

V= 8.0 L n= ?

Solution:

Use the formula, PV = nRT

$$n = PV / RT$$

$$= \frac{(1\text{atm})(8.0\text{L})}{(273\text{ K})\ (0.0821\text{ at-L/mol K})}$$

$$= 0.36 \text{ mol oxygen}$$

6. GRAHAM'S LAW

Thomas Graham discovered the relationship between the ability of a gas to pass through a small opening and its molecular weight. He found out that the rate of effusion of any gas is inversely proportional to the square root of its molecular weight. Thus, the heavier the gas molecule is, the slower the effusion of that gas.

7. DALTON'S LAW OF PARTIAL PRESSURE

According to John Dalton, the pressure exerted by a mixture of non-reacting gases is the sum of the partial pressures that each gas in the mixture exerts individually. Mathematically, this law is expressed as:

$$P_t = P_1 + P_2 + P_3 \ldots. + P_n$$

SOLUTION

A *solution* is a homogeneous mixture of two or more substances whose components are uniformly distributed all throughout. Solutions have two components, a **solute**, the substance to be dissolved and the **solvent**, the dissolving medium. Usually, the solvent is greater in quantity than the solute. In a solution containing sugar and water, the solute is the sugar and the solvent is the water.

Properties of Solutions

Concentration is the measure of the quantity of a solute in a given amount of solution or solvent. It can be expressed qualitatively and quantitatively. A **concentrated solution** *contains a large amount of solute per volume of solvent*. A **dilute solution** *contains a small amount of solute per volume of solvent*. **Saturated solution** contains as much solute as it can dissolve. In a saturated solution, dissolving and crystallizing occur at equal rates. An equilibrium exists between solute and solvent. When a solution can still hold more solute, the solution is still **unsaturated.** When a solution has more solute than it should normally hold, it becomes **supersaturated.** Excess solute will crystallize, making it an unstable solution.

For liquid in liquid solutions, solubility is described in terms of *miscibility*. When a solute and a solvent readily dissolve in any amount in each other, they are referred to as **miscible.** When the components of a solution only have limited solubility, then it is only **partially miscible**. Substances which are **immiscible** do not dissolve in each other. They form two phases or layers.

Solubility

Solubility is the measure of the amount of solute that can be dissolved in a given quantity of solvent at a specific temperature.

TEMPERATURE

The solubility of gases in water is inversely proportional to temperature. This means that an increase in temperature will decrease the solubility of gases in water.

PRESSURE

Henry's Law states that *the solubility of a gas in a liquid is directly proportional to the pressure exerted by the gas on the surface of the liquid.*

THE NATURE OF THE SOLUTES

"Like dissolves like. Polar solvents will dissolve polar solutes and non polar solutes dissolve in non-polar solvents. Water is the universal solvent because of its ability to dissolve a great number of polar and ionic compounds.

FACTORS THAT AFFECT THE RATE OF DISSOLUTION

1. ***Size of particles***
   The smaller the solute particles, the faster it dissolves. Smaller particles have greater surface areas exposed to the solvent.

2. ***Rate of stirring***
   The rate of dissolution is increased by stirring constantly. Stirring allows faster contact between the solute and the solvent particles. Stirring also increases the kinetic energy of the system.
3. ***Heating***
   Heating also increases the kinetic energy of the solute and the solvent. This means that dissolution increases as the temperature increases.

MODES OF EXPRESSING CONCENTRATION

Concentration is the ratio of a specified amount of solute to a specified volume of solvent or solution.

$$\text{Concentration} = \frac{\text{amount of solute}}{\text{Volume of solute /solvent}}$$

The following are various means of expressing concentration.

1. ***Percent Concentration*** is the percent of the solute in the solution.
   a. **Percent by mass (P_m)** is the mass of solute divided by the total mass of solution multiplied by 100. The mass of solution is equal to the mass of the solute plus the mass of the solvent.

APPLICATION

a.1. A chemist needs to prepare ceramic tile cleanser which contains 300 grams of hydrogen chloride and 600 grams of water. What is the percent by mass of the hydrogen chloride in the cleanser?

Given:

Mass of solute = 300 g
Mass of solvent = 600 g
Mass of solution = 300+600
= 900 g
P_m = ?

Solution:

$$P_m = \frac{\text{mass of solute}}{\text{Mass of solution}} \times 100\%$$

$$= \frac{300}{900} \times 100\%$$

$$= 33\%$$

a.2. How much water must be added to 50 grams of salt o prepare 50% solution?

Given:

Mass of solute = 50 grams
P_m = 50%

$$50\% = \frac{50 \text{ grams}}{50 + X} \times 100\%$$

$$0.50 = \frac{50g}{50+x}$$

$$0.50\,(50+x) = 50$$

$$25 + 0.50\,x = 50$$

$$0.50x = 50-25$$

$$x = 25 / 0.50 \qquad x = 50 \text{ g water}$$

b. ***Percent by volume***

To determine percent by volume, simply divide the volume of solute by the total volume of the solution and multiply the result by 100%. The solute and the solution volumes have to be expressed in the same units.

$$P_v = V_{solute} / V_{solution} \times 100\%$$

APPLICATION:

A solution is prepared with 15 cc and 35 cc hydrocarbon to make to make a solution. Determine the concentration of benzene in percent by volume.

Given:
Volume of solute = 15 cc
Volume of solvent = 35 cc
Volume of solution = 15 cc + 35 cc = 50 cc solution
Pv = ?

Solution:
Pv = 15 cc / 50 ccx 100% = 30%

c. **Percent by Mass – Volume** is used when dealing with a solid solute and a liquid solvent. Percentage by mass-volume is obtained by dividing the mass of solute, in grams by the total volume of solution in mL and the result is multiplied by 100%.

$$\text{Percent by mass-volume} = \frac{\text{mass of solute (g)}}{\text{Volume of solution (mL)}} \times 100\%$$

APPLICATION

A 0.84% (m/v) sodium chloride solution has to be administered a patient intravenously. Determine the volume of the solution if 208 grams of sodium chloride is used.

Given:
Percent by mass-volume = 0.84%
Mass of solute = 2.8 grams
Volume of solution = ?

Solution:
Pmv = 2.8 g / x * 100%
0.84% x = 2.8g * 100%
x = 280 g / 0.84
x = 322 mL

2. ***Mole Fraction*** is the ratio of the number of moles of solute in a given mole of solution.

$$\text{Mole fraction} = \frac{\text{mole of solute}}{\text{Mole solution}}$$

APPLICATION: What is the mole fraction of 35 g of calcium hydroxide dissolved in 42 grams of water?

Analysis of the Problem:

Mass of solute = 35 g
Mass of solvent = 42 g
Mole fraction = ?

Solution:
Determine the moles of solute and solvent from their molecular weight.

$Ca(OH)2$ = 40 + 2(16) + 1(2) = 74 g/mol
$H2O$ = 1(2) + 16 = 18 g/mol
mole $Ca(OH)2$ = 35 g / (74 g/mol) = 0.473 moles
moles of $H2O$ = 42 g/ (18 g/mol) = 2.333 mole
moles of solution = 0.473 + 2.333 = 2.803 moles

$$\text{mole fraction of } Ca(OH)\,2 = \frac{0.473 \text{ moles}}{2.806 \text{ moles}} = 0.17$$

3. Molarity

Molarity expresses the amount of solute in moles per liter of solution.

M = moles of solute / volume soln. liter

APPLICATION :Calculate the molarity if 45 g of sodium hydroxide is present in 200 cc of solution

Given:
Mass of solute = 45 grams
Volume of solution = 200 cc or 0.200L
M= ?

Solution:
Determine the molecular weight of sodium hydroxide:
NaOH = 23 + 16 + 1 = 40 g/mol

Convert the given mass to moles:
45 g NaOH ? (40 g/mol) = 1.125 moles NaOH
M= 1.125 moles / (0.200 L) = 5.6 molar

4. Molality

Molality is based on a fixed volume of a solution while molarity is based on a fixed mass of solvent.
Molality = moles of solute / mass solvent (kg)

APPLICATION:
Calculate the molality of a solution that contains 35 grams ammonia in 500 grams of water.
Given:

Mass of solute = 35 g
Mass of solvent = 500 grams or 0.510 kg
Molality =?

Solution:

Determine the molecular weight of ammonia and change the mass of ammonia into moles.
NH3 = 14 + 1(3) = 17 g/mole
Moles of ammonia = 35 grams / 17 g/mole = 2.06 moles
Molality = 2.06 moles / 0.5 kg = 4.1 molal

5. Normality

Normality is the number of equivalents of solute in a liter of solution
Normality (N) = number of equivalence / liter of solution
An **equivalent** is defined as the number of moles of an acid or base multiplied by the number of replaceable hydrogen or hydroxide ions it has. **Equivalent weight** is the molecular weight of an acid multiplied by the number of equivalent of hydrogen or hydroxide ion it has per molecule.

APPLICATION
Determine the normality of 1.5L solution that contains 18.3 g of sulfuric acid

Given:
Mass of solute = 18.3 grams sulfuric
Volume of solution = 1.5 liters
N = ?

Solution:
Determine the equivalent weight of sulfuric acid. Since there are two replaceable hydrogen ions in H_2SO_4, simply divide the molecular weight of sulfuric acid (98.0734 g/mol) by two.
Convert the mass of sulfuric acid to equivalence. Then calculate normality.
18.3 g H2SO4 x (1 equivalent/ 49.0367) = 0.373

N= 0.373 equivalent / 1.5 L = 0.25 normal

Acids and Bases

Definition of Acids and Bases

	Acid	**Base**
Arrhenius	A substance that yields H+ ions in aqueous solution	A substance that yields OH^- ions in aqueous solution
Bronsted Lowry	Proton donor	Proton acceptor
Lewis	Electron pair acceptor	Electron pair donor

Properties of Acids and Bases

Acid	**Base**
Sour taste	Bitter taste
Irritating smell (for most acids)	Slippery or soapy touch
Turns blue litmus paper to red	Turns red litmus paper to blue
pH<7, pOH > 7	pH >7 , pH < 7
Neutralizes a base	Neutralizes an acid
Good conductor of electricity	Good conductor of electricity

Classification of Acids and bases

Acids can be classified according to the number of hydrogen and hydroxyl group. A **monoprotic** contains 1 hydrogen, **diprotic**, 2 hydrogen and so on. A **monobasic** contains only one hydroxide group, **dibasic**, 2 and so on. Water is **amphoteric** because it can act as an acid or base.

pH and pOH

The pH of the solution is equal to the negative logarithm of its hydrogen ion concentration ([h+].
pH= -log [H+].
pOH = -log [OH]
pH + pOH = 14

* water has a pH of 7. It is neutral.

Organic Chemistry

Hydrocarbons- contain only hydrogen and carbon in their molecules. It can be classified as ***alkanes, alkenes, and alkynes.***

ALKANES- hydrocarbons that contain only single bonds in their molecules.
ALKENES- there is at least one carbon-to carbon double bond.
ALKYNES- hydrocarbon where there is at least one carbon-carbon triple bond.

Aromatics

Aromatic compounds are organic compounds having cyclical hydrocarbon rings where all the atoms are sp^2 hybridized.
Substituted hydrocarbons

a. ***Alcohol*** – with OH functional group (R – OH)
b. ***Ethers***- hydrocarbon chains attached to an oxygen atom (R-O-O-R)
c. ***aldehydes and ketones*** – have carbonyl group (C=O)
d. ***Halogenated hydrocarbons*** – hydrocarbons where one or more hydrogen is replaced by a halogen.
e. ***Amines***- if you replace a hydrogen atom from an ammonia molecule with a hydrocarbon
f. ***Amides*** – carboxyl group (C=O) comes between the hydrocarbon chain and the nitrogen of an amine

Biological Chemistry

Biological chemistry deals with the chemical substances that make up living things such as their tissues, body fluids, and others.

Biological Substances

1. ***Carbohydrates*** – include the different types of sugar, starch (like those from rice, bread, sweet potato) and the like. They are primary sources of energy in plant bodies. They are classified according to the number of simple sugar units or "saccharides" into the following:

a. **Monosaccharides** – made of one simple sugar unit (ex. Fructose, ribose, glucose)
b. **Disaccharides** – made up of two simple sugar units (ex. Maltose, sucrose)
c. **Polysaccharides**- made of more than two simple sugar units. (ex. Glycogen, cellulose)

2. ***Proteins*** – Substances classified as proteins vary in shapes, sizes and functions, but there are similarities in their chemical structures. They are all poly peptides (polymers of amino acids)

   Amino acids are chemically carboxylic acids with an amino group. There are about 20 common amino acids found in the body, each has a different side chain. The most common secondary structures are the alpha-helix, the beta-pleated sheet and the triple helix.

3. ***Nucleic Acids*** – a polymeric chain of nucleotides (polynucleotides). Each nucleotide component is made up of a phosphate group, a sugar, and a nitrogen base. The base can be a purine (adenine or guanine) or a pyrimidine (cytosine, thymine or uracil) derivative. The sugar found is either ribose or deoxyribose. It s the sugar that determines whether the nucleic acid you have is a DNA or RNA. If the sugar used is ribose, you get an RNA. If the sugar used is deoxyribose, you get a DNA.

   a. Deoxyribonucleic acid (DNA)

   DNA structure is a double – helix. This is responsible for the replication of DNA. The complementary base pairs in the DNA structure are : Adenine and Thymine , Guanine and Cytosine.

   b. Ribonucleic Acid (RNA)
   In RNA, uracil is used in place of thymine.

   Messenger RNA (mRNA) copies genetic information from DNA in the cell nuclei to the ribosomes (where proteins are made.)

   Transfer RNA (tRNA) carries amino acids to the site of protein synthesis

   Ribosomal RNA (rRNA) provides the site for protein synthesis.

4. ***Lipids*** – fatty acids and their esters. Fatty acids are chemically carboxylic acids. If you eat too much, the food you overeat will be stored as lipids in adipose tissues. Two types of lipids are: the simple lipids and compound lipids.