

ST. LAWRENCE HIGH SCHOOL A JESUIT CHRISTIAN MINORITY INSTITUTION STUDY MATERIAL FOR CHEMISTRY (CLASS-11) TOPIC- BASIC CONCEPTS IN CHEMISTRY PREPARED BY: MR. ARNAB PAUL CHOWDHURY SET NUMBER-02 DATE: 06.07.2020



SOME BASIC CONCEPTS OF CHEMISTRY

PHYSICAL QUANTITIES AND SI UNITS

The 11th general conference of weights and measures in 1960 recommended the use of international system of units. Abbreviated as SI Units (after the French expression La System International de units).

FUNDAMENTAL UNITS

The SI system has seven basic units of physical quantities as follows:

Physical Quantity	Abbreviation	Name of unit	Symbol
time	t	second	S
mass	m	kilogram	kg
length	1	metre	m
temperature	Т	kelvin	К

electric current	1	ampere	A
light intensity	lv	candela	Cd
amount of substance	n	mole	mol

DERIVED UNITS

The units obtained by combination of basic units are known as derived units e.g. velocity is expressed as distance/time. Hence unit is m/s or ms-1. Some common derived units are:

Physical Quantity	Definition	SI Unit
volume	length cube	m3
area	length square	m2
speed	distance travelled	ms–1 per unit time
acceleration	speed changed	ms–2 per unit time
density	mass per unit volume	kg m–3
pressure	force per unit area	kgm–1s–2 or Nm–2 (pressure =

		Pa)
force	mass times acceleration of object	kgms–2 (Newton N)
energy	force times distance travelled	kgm2s–2(Joule J)
frequency	cycles per second	s—1 (hertz = Hz)
power	energy per second	kgm2s–3 or Js–1 (Watt = W)
electric charge	ampere times second	As (coloumb = C)
electric potential	energy per unit	JA–1s–1 or kgm2s–3
difference	charge	A–1 (volt = V)

SOME NON-SI UNITS IN COMMON USE

Quantity	Unit	Symbol	SI definition	SI Name
Length	angstrom	Å	10–10 m	0.1 nanometers (nm)
Volume	litre	L	10–3 m3	1 decimeter (dm3)
Energy	calorie	cal	kg m2s–2	4.184 Joule (J)

STANDARD PREFIXES FOR EXPRESSING THE DECIMAL FRACTIONS OR MULTIPLES OF FUNDAMENTAL UNITS

Fraction	Prefix	Symbol	Multiple	Prefix	Symbol
10–1	deci	d	101	Deka	da
10–2	centi	С	102	Hecta	h
10–3	milli	m	103	kilo	k
10–6	micro	m	106	Mega	М
10—9	nano	n	109	Giga	G
10–12	pico	р	1012	Tera	Т
10–15	femto	f	1015	Peta	Р
10–18	atto	a	1018	Exa	E

PRECISION AND ACCURACY

Precision: It is the closeness of various measurements for the same quantity.

Accuracy: It is the agreement of a particular value to the true value.

Example: Let the true weight of a substance be 3.00g. The measurement reported by three students are as follows

Student	Measurements/g		Average/g
	1	2	
Α	2.95	2.93	2.94
В	3.01	2.99	3
С	2.94	3.05	2.99

<u>Case of A student</u>: It is precision but no accuracy since measurements one close but not accurate.

Case of B student: Measurements are close (precision) and accurate (Accuracy)

<u>Case of C student</u>: Measurement are not close (no precision) and not accurate (no accuracy)

STOICHIOMETRY

<u>Stoichiometry</u>: It is calculation of masses or volumes of reactants and products involved in a chemically balanced reaction. Consider the formation of ammonia.

 $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

All are gases indicated by letter (g) and coefficients 3 for H_2 and 2 for NH_3 are called stoichiometric coefficients. The formation of ammonia can be interpreted in many ways:

- One mole of $N_2(g)$ reacts with three moles of $H_2(g)$ to give two moles of $NH_3(g)$.
- 28g of $N_2(g)$ reacts with 6g at $H_2(g)$ to give 34g of $NH_3(g)$.
- 22.4L of $N_2(g)$ reacts with 67.2L of $H_2(g)$ to give 44.8L of $NH_3(g)$

SIGNIFICANT FIGURES

The weight 7.52 gm of a substance indicates that it is reliable to the nearest hundredth of a gram and may be expressed as 7.52 \pm 0.01. It means slightest variation may occur at the second place of decimal or we can say that uncertainty is \pm 0.01 g.

Now consider the weight 6.4234 g. It may correctly be expressed as 6.4234 ± 0.001 g.

In the first case the weight contains three significant figures and in the second case weight contains five significant figures.

Significance of zero : If zero is used to locate the decimal point it is not considered as significant figure. Thus in 0.0072 there are only two significant figures whereas in 70.40, there are four

significant figures since zero is after 4. Again in 0.0070 there are two significant figures, since zero after 7 is significant for it has a meaning when written in exponentials. If we compare $7.0 \times 10-3$ and $7 \times 10-3$, the first term has uncertainty of one in seventy and second has uncertainty of one in seven. The exponential term does not add to number of significant figures.

Addition and subtraction of quantities: In this case the uncertainty in the result is equal to the sum of the uncertainties of the individual quantities.

Multiplication and division: In this case the uncertainty in the result is equal to the sum of the percentage of individual uncertainties.

Rounding off: The following rules are observed.

If the digit after the last digit to be retained is less than 5, the last digit is retained as such.
e.g. 1.752 = 1.75 (2 is less than 5).

2) If the digit after the last digit to be retained is more than 5, the digit to be retained is increased by 1. e.g. 1.756 = 1.76 (6 is more than 5).

3) If the digit after the last digit to be retained is equal to 5, the last digit is retained as such if it is even and increased by 1 if odd.

Calculations involving addition and subtraction : In case of addition and subtraction the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places .

Calculations involving multiplication and division : In this case the final result should be reported having same number of significant digits as that of the number having least significant digits.

MATTER

Anything which occupies space, possesses mass and can be felt is called matter.

CLASSIFICATION OF MATTER



ELEMENT

Pure substance consisting of one type of particles in the form of atoms eg. Cu, Na, Fe or molecules eg. H_2 , O_2 etc.

COMPOUND

Pure substance consisting of molecules formed by the combination of atoms of different elements eg. CO_2 , H_2O etc.

MIXTURES

Mixtures are substances made of two or more elements or compounds in any proportion. They may be homogeneous or heterogeneous.

SEPARATION OF MIXTURES

Mixtures can be separated into constituents by following methods:

1) Filtration can separate those mixtures whose one component is soluble in a particular solvent and other is not.

2) Distillation can be used to separate constituents of mixtures having different boiling points.

3) Extraction dissolves one out of several components of mixture.

4) Crystallisation is a process of separating solids having different solubility in a particular solvent.

5) Sublimation separates volatile solids which sublime on heating from non-volatile solids.

6) Chromatography is the technique of separating constituents of a mixture which utilises the property of difference of adsorption on a particular adsorbent.

7) Gravity separation separates constituents having different densities.

8) Magnetic separation can separate magnetic components from non-magnetic ones.

PHYSICAL AND CHEMICAL CHANGES

A change which does not affect chemical composition and molecular structure is a physical change and the one that involves alteration of chemical composition and molecular structure is a chemical change.

1) Chemical Combination is reaction between two or more elements or compounds to form a single substance.

2) Displacement means replacement of one element of compound by another.

3) Decomposition involves splitting of a compound to form two or more substances.

4) Combustion is a complete and fast oxidation of a substance.

5) Neutralisation is the reaction between acid and base to form a salt.

6) Polymerisation is the combination of molecules of same or different substances to form a single molecule called polymer

7) Photochemical changes occur in presence of visible or ultraviolet light.

8) Double decomposition or metathesis is the exchange of oppositely charged ion on mixing two salt solutions.

9) Hydrolysis involves reaction of salts with water to form acidic or basic solutions.

LAWS OF CHEMICAL COMBINATIONS

A) <u>Law of conservation of mass</u>: This law was given by French chemist A. Lavoisier (1774) which states that "during any physical or chemical change, the total mass of products is equal to the total mass of reactants". It is also called law of indistinctibility. It does not hold good for nuclear reaction.

B) <u>Law of definite proportions</u> : This law was given by Proust (1799) and states that "a chemical compound always contains some elements combined together in same proportion by mass". For example different samples of pure CO2 always have carbon and oxygen in 3 : 8 ratio by mass.

C) <u>Law of multiple proportions</u>: This law was given by John Dalton (1803) and states that "when two elements combine to form two or more compounds, the different mass of one of the elements and the fixed mass of the one with which it combines always form a whole number ratio". This law explains the concept of formation of more than one compound by two elements.

D) <u>Law of reciprocal proportions</u>: This was given by Richter (1792) and states that "when two elements combine separately with a fixed mass of third, the ratio of masses in which they do so is same or whole number multiple of the ratio in which they combine with each other." This law is also called law of equivalent proportions and is helpful in determining equivalent weights.

E) <u>Gay Lussac's law of combining volumes</u>: This law states that when gases react with each other, their volumes bear a simple whole no. ratio to one another and to volume of products (if gases) and similar conditions of pressure and temperature.

Dalton's atomic theory:

Proposed by John Dalton in 1808. Main points are : Matter is made up, by indivisible particles called atoms Atoms of same elements are identical in physical and chemical properties. Atoms of different substances are different in every respect Atoms always combine in whole numbers to form compounds Atoms of resultant compounds possess similar properties

Drawbacks of Dalton's theory:

Does not explain structure of atom. Fails to explain binding forces between atoms in compounds. Does not explain Gay Lussac's law. Does not differentiate between atom and molecule.

Avogadro's law :

It states that "equal volumes of all gases, under similar conditions of temperature and pressure contain equal number of molecules". Applications are:

- Deducing atomicity of elementary gases
- Deriving relationship between molecular mass and vapour density

- Deriving formula of substances
- Determining molecular wt. of a gas
- Deducing the gram molecular volume.

ΑΤΟΜ

Atom is the smallest particle of element which might not be able to exist independently.

MOLECULE

Molecule is the smallest particle of the substance which can exist independently. It can be subdivided as

 Homoatomic molecules are molecules of same element and can be further divided as monoatomic, diatomic and polyatomic molecules depending upon number of atoms. eg: He, O₂, P₄ etc.

2) Heteroatomic molecules are molecules of compound. They can be diatomic and polyatomic. eg: H₂O, PCI₅, H₂SO₄, NO etc.

ATOMIC MASS UNIT (A.M.U.)

It is the unit of representing atomic masses. 1 a.m.u. = th the mass of C-12.

MOLE

It is a unit which represents 6.023 × 1023 particles. The number 6.023 × 1023 is called Avogadro's number and is represented by N0 or NA. Avogadro's number of gas molecules occupy a volume of 22400 cm3 at N.T.P. Number of molecules in 1 cm3 of gas at NTP is Loschmidt N0. With value 2.688 × 1019.

ATOMIC MASS

"It is the number of times the atom of the element is heavier than H atom" was the first proposed definition. Later on oxygen was preferred as standard. In 1961 C-12 was chosen as standard and thus "the number of times the atom of an element is heavier than 12th part of C-12 is called atomic mass of the element.

Atomic mass =
$$\frac{\text{Mass of an atom of the element}}{\frac{1}{12} \times \text{mass of C-12 atom}}$$

AVERAGE ATOMIC MASS

It is the mass of each isotope determined separately and then combined in ratio of their occurrence. Suppose a and b are two isotopes of an element with their occurence ratio p : q then

 $p \times a + q \times b$

Average atomic mass = p+q

DETERMINATION OF ATOMIC MASS

1) Dulong and petit's rule : It is based on experimental facts. "At ordinary temperature, product of atomic mass and specific heat for solid elements is approximately 6.4 and this product is known as atomic heat of the element"

Atomic mass × specific heat = 6.4

The law is valid for solid elements except Be, B, Si and C.

Correct At. Mass = Eq. mass × valency

2) Specific heat method: This method is for gases. $\gamma = \frac{C_p}{C_v}$, where Cp = specific heat at constant pressure and Cv = specific heat at constant volume. the ratio g is a constant = 1.66 for monoatomic, 1.40 for diatomic, 1.33 for triatomic gas and atomic mass of gaseous element $\frac{Mol. mass}{atomicity} = \frac{Mol. mass}{\gamma}$

3) Chloride formation method: This method converts the element (whose mass is to be determined) into volatile chloride whose vapour density is found by Victor Mayer method. Molecular mass = 2 × V.D.

4) Vapour density method is suitable for elements having volatile chlorides.

Atomic mass = Eq. mass of metal × valency.

5) Mitscherlich's law of isomorphism : It states that isomorphous substances have similar chemical constitution. Isomorphous substances form crystals of same shape and valencies of elements forming isomorphous salts are also same. eg: ZnSO4. 7H2O, MgSO4.7H2O and FeSO4.7H2O are isomorphous.

GRAM ATOMIC MASS (GAM)

Is the mass of an atom expressed in gms.

No. of Gm-atoms of element = $\frac{Mass \text{ of element in grams}}{GAM \text{ of element}}$

MOLECULAR MASS:

It is the average relative mass of the molecule as compared with mass of C-12 atom.

Average relative mass of one molecule

 $\frac{1}{12}$ × Mass of C-12 atom

Molecular mass =

CALCULATION OF MOLECULAR MASS:

1) Graham's law of diffusion: It states that rate of diffusion of two gases is inversely proportional to the square root of ratio of their molecular weights.

$$\frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

2) Victor meyer method : This method can determine the molecular mass as

Molecular mass = $w/v \times 22400$

Where, W is the mass of liquid in gm. occupying a volume V ml at STP.

3) Vapour density method : Vapour density is the ratio of volume of a gas to the mass of same volume of hydrogen under identical conditions.

$$V.D. = \frac{\text{Mass of V litre of gas}}{\text{Mass of V litre of hydrogen}}$$
$$V.D. = \frac{1}{2} \frac{\text{Weight of volatile substance}}{\text{Volume at STP}} \times 22400$$

Thus molecular mass = $2 \times V.D$.

4) Colligative properties method : This method can be helpful in determining molecular mass

 $(\Delta T_b) = \frac{K_b \times w \times 1000}{W \times mol. mass}$

Where, T_b is elevation in b.pt. K_b is molal elevation constant w is wt. of solute W is wt. of solvent

$$(\Delta T_f) = \frac{K_f \times w \times 1000}{W \times mol. mass}$$

Depression in freezing point

GRAM MOLECULAR MASS OR MOLAR MASS:

That amount of substance whose mass in grams is equal to its molecular mass or the equivalently molecular mass of a substance expressed in grams is called gram molecular mass. Gram molecular mass is also called one gram molecule. Thus,

No. of gm molecules = $\frac{Wt. of substance (gms.)}{GMM of substance}$

EQUIVALENT MASS:

It is the number of parts by weight of the substance(at S.T.P.) that combines or displaces, directly or indirectly, 1.008 parts by mass of hydrogen or 8 parts by mass of oxygen or 35.5 parts by mass of chlorine. It can be calculated as-

Equivalent mass for elements = Atomic Mass/ Valency Equivalent mass for acids = Molecular Mass/ Basicity Equivalent mass for bases = Molecular/ Acidity Equivalent mass for salts = Formula Mass/(valency of the cation x number of cation) Equivalent mass for oxidizing agents = Formula Mass/no, of electrons gained per molecule Equivalent mass for reducing agents = Formula Mass/no, of electrons lost per molecule Equivalent weight of radicals = Formula Mass of the radical/No. of units of charge

FORMULA MASS:

It is obtained by adding atomic masses of various atoms present in the formula and this term replaces molecular mass in ionic compounds.

ACIDITY:

It is the number of OH– ions that can be displaced from one molecule of a substance.

BASICITY:

It is the number of H+ ions that can be displaced from one molecule of a substance.

GRAM EQUIVALENT MASS (GEM):

It is the mass of a substance expressed in grams or equivalently the quantity of substance whose mass in grams is equal to its equivalent mass is called one gram equivalent or gram equivalent mass.

No. of gm equivalents = $\frac{Mass in gm}{GEM}$.

METHODS OF DETERMINING EQUIVALENT MASSES:

1) Hydrogen displacement method : It is for metals which can displace H2 from acids.

Equivalent mass of metal

Weight of metal Weight of displaced hydrogen ×1.008

 $\frac{\text{Weight of metal in gram}}{\text{Vol. of H}_2 \text{ in litre}} \times 11.2 \text{ litre}$

2) Metal displacement method : It utilises the fact that one GEM of a more electropositive

metal displaces one GEM of a less electropositive metal from its salt $\frac{W_1}{W_2} = \frac{E_1}{E_2}$

3) Conversion method : When one compound of a metal is converted to another compound of

similar metal then $\frac{\text{Weight of first compound}}{\text{Weight of second compound}} = \frac{\text{E} + \text{Eqv. mass of first radical}}{\text{E} + \text{Eqv. mass of second radical}}$

Where, E is the equivalent. Mass of the metal.

4) Electrolytic method:

It states that the quantity of substance that reacts at electrode when Faraday of electricity is passed is equal to its GEM.

GEM = Electrochemical equivalent × 96500

and ratio of weights deposited by equal amount of electricity is in ratio of their equivalent masses.

5) Oxide method :

Equivalent mass of metal = $\frac{\text{Weight of metal}}{\text{Weight of oxygen}} \times 8.0 = \frac{\text{Weight of metal in gram}}{\text{Vol. of O}_2 \text{ in litre}} \times 5.6 \text{ litre}$

6) Double decomposition :

$$\frac{AB + CD \longrightarrow AD \downarrow + CB}{\text{Weight of compound AB}} = \frac{\text{Eqv. mass of A + Eqv. mass of B}}{\text{Eqv. mass of A + Eqv. mass of D}} \Rightarrow \frac{\text{Wt. of salt}}{\text{Wt. of ppt.}} = \frac{\text{Eqv. mass of salt}}{\text{Eqv. mass of salt in ppt}}$$

7) Neutralisation method for acids and bases :

	Wt. of acid (base)		
	Vol. of acid (base) in one litre required	for neutralization	
Equivalent mass of acid (base)=	× Normality of acid (base)		

8) Silver salt is method commonly used for organic acids.

 $\frac{108 \times Mass \text{ of silver salt}}{\text{Mass of Ag metal}} - 107$

Eqv. mass of acid =

Mol. mass of acid = Eqv. mass of acid × Basicity

9) Platinichloride method for bases :

195×Mass of platinum salt - 205 Mass of pt. metal × 2

Eqv. mass of base=

Mol. mass of base = Eqv. mass of base × Acidity

10) Chloride method :

Eqv. mass of metal = $\frac{\text{Weight of metal}}{\text{Weight of chlorine}} \times 35.5 = \frac{\text{Weight of metal in gram}}{\text{Vol. of Cl}_2 \text{ in litre}} \times 11.2 \text{ litre}$

11) Volatile chloride method :

Valency of metal

 $\frac{2 \times VD \text{ of chloride}}{\text{Eq. mass of metal chloride}} = \frac{2 \times VD}{E + 35.5} \implies E = \frac{2 \times VD \text{ of chloride}}{Valency} - 35.5$

CHEMICAL EQUATION :

It is the equation representing chemical change in terms of formula of reactants and products.

1) An equation which has not been equalised in terms of number of atoms of reactants and products is called a skeleton equation.

2) An equation having equal number of atoms of various kinds on both sides is a balanced equation.

EMPIRICAL FORMULA:

It is the simplest formula of a compound giving simplest whole number ratio of atoms present in one molecule. e.g. CH is empirical formula of benzene.

MOLECULAR FORMULA:

It is the actual formula of a compound showing the total number of atoms of constituent elements e.g. C6H6 is molecular formula of benzene.

Molecular formula = n × empirical formula, where n is simple whole number.

SOLUTION :

It is a homogenous mixture of two or more substances. The component of solution having larger proportion is solvent and others are solute.

MOLE FRACTION :

It is the ratio of moles of a constituent to the total number of moles in a solution.

Let A be solute & B is solvent then mole fraction of solute (xA)

= $\frac{n_A}{n_A + n_B}$, where n is the number of moles.

Mole fraction of solution $x_A + x_B = 1$

MASS PERCENTAGE :

It is the number of parts by mass of solute per hundred parts by mass of solution. If WA is mass of solute and WB the mass of solvent, then

Mass percentage of A = . $\frac{W_s}{W_s + W_s} \times 100$

VOLUME PERCENTAGE :

It is the number of parts by volume of solute per hundred parts by volume of solution. If VA is volume of solute and VB is the volume of solvent then

Volume percentage of A = $\frac{V_A}{V_A + V_B} \times 100$

PARTS PER MILLION (PPM) :

It is the mass of solute present in one million parts by mass of solution.

 $ppm = \frac{\text{Mass of solute}}{\text{Mass of solution}} \times 10^6$

NORMALITY :

It is the number of gram equivalents of a solute present in one litre of solution.

 $Normality = \frac{\text{Gram equivalents of solute}}{\text{Volume of solution in litre}} = \frac{\text{Mass of solute in gms.}}{\text{GEM of solute × volume of solution in litre}}$

Normality depends on temperature. Also if strength is given in normalities, N1 of A & N2 of B

Then N1V1 = N2V2.

MOLARITY :

It is the number of moles of solute present in one litre of solution.

 $M = \frac{\text{Moles of solute}}{\text{Volume of solution (L)}} = \frac{\text{Weight of solute}}{\text{GMM of solute} \times \text{Volume of solution (L)}}$

and millimoles = $M \times V$ (in ml).

Molarity and mass percentage have the relation M

$$\implies Moles = M \times V_{(in \ litre)} = \frac{mass \ percentage \times 10 \times d}{GMM \ of \ solute}, \text{ where } d = density$$

If a solution of molarity M1 and volume V1 adds up with a solvent to a final volume V2, then molarity M2 is given by

$$M_2 = \frac{M_1 V_1}{V_2}$$

If two different solutions (M1, V1) and (M2, V2) are mixed then molarity of resulting solution is

$$M = \frac{M_1V_1 + M_2V_2}{V_1 + V_2}$$

Also, Molarity × GMM of solute = Normality × GEM of solute

MOLALITY :

It is the number of moles of solute in 1 kg of solvent.

 $\frac{\text{Weight of solute} \times 1000}{\text{Mol. wt. of solute} \times \text{Wt. of solvent}} = \frac{\text{Weight of solute} \times 1000}{\text{Mol. wt. of solute} \times \text{Wt. of solvent}}$

Molality is independent of temperature.

FORMALITY (F) :

It is the number of gram formula mass of ionic solute dissolved in 1 litre of solution.

Formality = Mass of solute (gm) Volume of solution (litre)×GFM of solute

LIMITING REAGENT:

What is Limiting Reagents?

The reactant that is entirely used up in a reaction is called as limiting reagent. Limiting reagents are the substances that are completely consumed in the completion of a chemical reaction. They are also referred to as limiting agents or limiting reactants. According to the stoichiometry of chemical reactions, a fixed amount of reactants are required for the completion of the reaction. Let us consider the following reaction of formation of ammonia:

 $3H_2 + N_2 \rightarrow 2NH_3$

In the reaction given above, 3 moles of Hydrogen gas are required to react with 1 mole of nitrogen gas to form 2 moles of ammonia. But what if, during the reaction, only 2 moles of hydrogen gas are available along with 1 mole of nitrogen.

In that case, the entire quantity of nitrogen cannot be used (because the entirety of nitrogen requires 3 moles of hydrogen gas to react). Hence, the hydrogen gas is limiting the reaction and is therefore called the limiting reagent for this reaction.

Limiting Reagent Explanation

This reactant generally determines when the reaction will stop. The exact amount of reactant which will be needed to react with another element can be calculated from the reaction stoichiometry. The limiting reagent **depends on the mole ratio, not on the masses of the reactants present.**



Limiting Reagent Before and After Reaction

From the illustration shown above, it can be observed that the limiting reactant is the reason the reaction cannot continue since there is nothing left to react with the excess reactant. it is the reactant that entirely consumed over the course of the reaction.

Limiting Reagent Examples

Consider 1 mol of oxygen and 1 mol of hydrogen are present to undergo the following reaction.

$2H_2 + O_2 \rightarrow 2H_2O$

Since the reaction uses up hydrogen twice as fast as oxygen, the limiting reactant would be hydrogen.

Example: 100g of hydrochloric acid is added to 100g of zinc. Find the volume of hydrogen gas evolved under standard laboratory conditions.

Solution:

The chemical equation for this reactions is given below.

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

Zinc chloride is formed in excess so the limiting reagent here is hydrochloric acid.

73g of HCl = 24.5l of H₂ 100g of HCl = yL of H₂

y/24.5 = 100/73 y = (100 x 24.5)/73 y = 33.6L

Therefore 33.6L of H₂ is produced under standard laboratory conditions.

How to find Limiting Reagent?

The determination of the limiting reactant is typically just a piece of a larger puzzle. In most limiting reactant stoichiometry problems, the real goal is to determine how much product could be formed from a particular reactant mixture. The limiting reactant or reagent can be determined by two methods.

- 1. Using the mole ration
- 2. Using the product approach

In order to calculate the mass of the product first, write the balanced equation and find out which reagent is in excess. Using the limiting reagent calculate the mass of the product.

The following points should be considered while attempting to identify the limiting reagent:

- When there are only two reactants, write the balanced chemical equation and check the amount of reactant B required to react with reactant A. When the amount of reactant B is greater, the reactant A is the limiting reagent.
- The reactant which is in a lesser amount than is required by stoichiometry is the limiting reactant.
- In an alternate method of finding the limiting agent, the amount of product formed by each reactant is calculated.
- The limiting reactant is the reactant from which the minimum amount of product is formed.
- Also if we calculate the amount of one reactant needed to react with another reactant, then the reactant which is in shortage would be the required limiting reactant.

Thus the required limiting reagent for the reaction can be identified using the points provided above. These reagents are very important while calculating the percentage yield of a given reaction.

Important Questions:

1. What is the limiting reagent?

Ans. A reactant in a chemical reaction that determines the amount of product that is produced is the limiting reactant or limiting reagent. The reason for a limiting reactant is that elements and compounds react in a balanced chemical equation according to the mole ratio between them.

2. What is the difference between limiting reactant and limiting reagent?

Ans. A reactant is a substance that reacts directly when the reaction is initiated, whereas a reactant is a substance that participates in a chemical reaction.

3. Why is the limiting reactant important?

Ans. The limiting reagent/reactant is important because it can tell a chemist that only x moles of compounds can form when the perfect quantity is used with how much of this material they use, because it restricts the reaction, instead of the hypothetically.

4. Are limiting reactants present in all reactions?

Ans. The limiting reactant determines the maximum amount of product that can be formed from the reactants when reactants are not present in stoichiometric quantities.

5. What is the benefit of having a limiting reagent?

Ans. In a chemical reaction, the task of limiting reagent or reactant is significant because it can help the chemist predict the maximum quantity of reactant is consumed, since it restricts the reaction, only the necessary moles of products can be produced instead of the hypothetical yield where the perfect quantity is used.

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