Some Basic Concepts of Chemistry

Precision and accuracy of measurement

Accuracy: **It is the agreement of a particular value to the true value**.
Aim of any measurement is to get the actual value called true value or accepted value of a quantity. Nearness of the measured value to the true value is called the accuracy of measurement. Larger the accuracy smaller the error. Accuracy depends upon the sensitivity or least count (the smallest quantity that can be measured) of the measuring equipment.

Errors: It may be expressed as absolute or relative error.

Absolute error = Observed value - True value

Relative error is the ratio of an absolute error to the true value. It is expressed as a percentage.

Relative error = (Absolute error/ True value) x 100 % .

There can be error in a measurement due to a number of reasons including inefficiency of the person doing measurement.

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

Student	Measurements		Average
	1	2	
A	2.95	2.93	2.94
B	3.01	2.99	3
C	2.94	3.05	2.99

Case of student A : It is precise but not accurate since measurements one close but not accurate.

Case of student : Measurements are close (precise) and accurate.

Case of C student : Measurement are not close (not precise) and not accurate.

Significant Figures :

Uncertainty in measured value leads to uncertainty in calculated result. Uncertainty in a value is indicated by mentioning the number of significant figures in that value. Consider, the column reading 10.2 ± 0.1 mL recorded on a burette having the least count of 0.1 mL. Here it is said that the last digit '2' in the reading is uncertain, its uncertainty is ±0.1 mL. On the other hand, the figure '10' is certain.

The **significant figures** in a measurement or result are the number of digits known with certainty plus one uncertain digit.

Rules for deciding significant figures :

1. All non zero digits are significant; e. g. 127.34 g contains five significant figures which are 1, 2, 7, 3 and 4.

2. All zeros between two non zero digits are significant e. g. 120.007 m contains six significant figures.

3. Zeroes on the left of the first non zero digit are not significant. Such a zero indicates the position of the decimal point. For example, 0.025 has two significant figures.

4. Zeroes at the end of a number are significant if they are on the right side of the decimal point. Terminal zeros are not significant if there is no decimal point. (This is beacause the least count of an instrument contains decimal point) For example 0.400 g has three singnificant figures. The measurements here indicates that it is made on a weighing machine having least count of 0.001 g.

Significant figures are also indicated in scientific notation by means of decimal point. For example, the measurment 400 g has one significant figure. The measurement 4.0×10^2 g has two significant figures, wheras the measurment 4.00×10^2 g has three significant figures. The zeros after the decimal points in these cases indicates that the least counts of the weighing machines are 1 g, 0.1 g and 0.01 g, respectively.

5. In numbers written is scientific notation, all digits are significant. For example, 2.035×10^2 has four significant figures.

Rounding off :

The following rules are observed.

(a) 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).

(b) 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).

(c) 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.

e.g. 1.755 1.76 * = (* odd) & 1.765 1.76 * = (* even)

Calculations involving addition and substraction :

(i) *In case of addition and substraction the final result should be reported to the same number of decimal places as the number with the minimum number of decimal places e.g.* 34.72 (has two decimal places) + 8.1 (has one decimal place) = 42.82, but it should have only one decimal place so answer is 42.8.

(ii) 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. Example : $9.24 \times 3.6 = 33.264$ Rounded off to 33.

(iii) In case of division 5.235 /13.1 = 0.3996 , rounded off to 0.400.

Laws of Chemical Combination : The elements combine with each other and form compounds. This process is governed by five basic laws discovered before the knowledge of molecular formulae.

Law of conservation of mass: mass can neither be created nor destroyed, and Total mass of reactants = Total mass of products.

Law of Definite Proportions :A given compound always contains exactly the same proportion of elements by weight.

Law of multiple proportions :When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers

Law of Reciprocal proportions: If two different elements combine separately with a fixed mass of a third element, the ratio of the masses in which they do so are either the same as or a simple multiple of the ratio of the masses in which they combine with each other.

Gay Lussac Law of Gaseous Volume : When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all gases are at same temperature and pressure.

Avogadro Law : Equal volumes of all gases at the same temperature and pressure contain equal number of molecules.

Dalton's Atomic Theory :

1. Matter consists of tiny, indivisible particles called atoms.
2. All the atoms of a given elements have identical properties including mass. Atoms of different elements differ in mass.
3. Compounds are formed when atoms of different elements combine in a fixed ratio.
4. Chemical reactions involve only the reorganization of atoms. Atoms are neither created nor destroyed in a chemical reaction. Dalton's theory could explain all the laws of chemical combination.

Atomic Mass : Atomic mass is the mass of an atom. It is actually very very small and not easy to measure. In the present system, mass of an atom is determined relative to the mass of a carbon - 12 atom as the standard. The atomic masses are expressed in amu. One amu is defined as a mass exactly equal to one twelfth of the mass of one carbon-12 atom.

- 1 amu = $1/12 \times$ mass of one C-12
  $= 1/12 \times 1.992648 \times 10^{-23}$ g
  $= 1.66056 \times 10^{-24}$ g

Recently, amu has been replaced by unified mass unit called dalton (symbol 'u' or 'Da'), 'u' means unified mass.

Average Atomic Mass :Many naturally occuring elements exist as mixture of more than one isotope. Isotopes have different atomic masses. The atomic mass of such an element is the weighted average of atomic masses of its isotopes (taking into account the atomic masses of isotopes and their relative abundance i.e. percent occurrance). This is called average atomic mass of an element.
For example the average atomic mass of carbon
$= (12\ u)(98.892/100) + (13.00335\ u)(1.108/100) + (14.00317)(2 \times 10\text{-}10/100)$

Molecular Mass : Molecular mass of a substance is the sum of average atomic masses of all the atoms of elements which constitute the molecule. Molecular mass of a substance is the mass of one molecule of that substance relative to the mass of one carbon-12 atom.

Formula Mass: Some substances such as sodium chloride do not contain discrete molecules as the constituent units. In such compounds, cationic (sodium) and anionic (chloride) entities are arranged in a three dimensional structure, NaCl is the formula used to represent sodium chloride, though it is not a molecule. **The formula mass of a substance is the sum of atomic masses of the atoms present in the formula.**

Mole :One mole is the amount of a substance that contains as many entities or particles as there are atoms in exactly 12 g (or 0.012 kg) of the carbon -12 isotope.

Moles and gases : "One mole of any gas occupies a volume of 22.4 dm3 at standard temperature (00 C) and pressure (1 atm) (STP). The volume of
22.4 dm3 at STP is known as **molar volume** of a gas.

Molar Mass : The mass of one mole of a substance (element/compound) in grams is called its molar mass

Concentration of solution: The concerntration of a solution or the amount of substance present in given volume of a solution can be expressed in any of the following ways :
1. Mass percent or weight percent (w/w %)
2. Mole fraction
3. Molarity (M)
4. Molality (m)

Mass percent : It is obtained by using following relation:
Mass percent = (Mass of solute /Mass of solution) × 100 %

Mole fraction: It is the ratio of number of moles of a particular component of a solution to the total number of moles of the solution.

Molarity:It is the most wideley used unit and is denoted by M. It is defined as the number of moles of the solute present in 1 litre of the solution . Thus,
Molarity (M) = No. of moles of solute/ Volume of solution in litres

Molality: It is defined as the number of moles of solute present in 1 kg of solvent. It is denoted by m.
Molality (m) = No. of moles of solute / Mass of solvent in kg
Note that molality of a solution does not change with temperature since mass remains unaffected with temperature.

NORMALITY : It is the number of gram equivalents of a solute present in one litre of solution,
Normality = Gram equivalents of solute / Volume of solution in litre

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. C_6H_6 is molecular formula of benzene.
Molecular formula = n × empirical formula, where n is simple whole number

EQUIVALENT MASS : It is the number of parts by weight of the substance 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
(i) Equivalent mass for elements = Atomic mass/ Valency
(ii) Equivalent mass for acids = Molecular mass of acids / Basicity
(iii) Equivalent mass for bases = Molecular mass/ Acidity of base
Equivalent mass for salts= Formula mass / (Valency of cation)×(No. of cations)

METHODS OF DETERMINING EQUIVALENT MASSES :

(i) **Hydrogen displacement method** : It is used for metals which can displace H_2 from acids.
Equivalent mass of metal =(Weight of metal/Weight of displaced hydrogen)×1.008
=(Weight of metal in gram/Vol. of H2 in litre) × 11.2

(ii). **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.
$W_1/E_1 = w_2/E_2$

(iii). **Oxide formation method** : Equivalent mass of metal = (Weight of metal / Weight of oxygen)×8

(iv) **Chloride formation method** : Eqv. mass of metal = (Weight of metal / Weight ofchlorine)×35.5

(v) **Neutralisation method for acids and bases** :
Equivalent mass of acid (base) = Wt.of acid (base)/ {Normality of acid (base) ×Vol.of acid (base)in one litre required for neutralization}

(vi) **Conversion method** : When one compound of a metal is converted to another compound of similar metal then
Weight of first compound/ Weight of second compound
= (E + Eqv.mass of first radical)/ (E+ Eqv.mass of second radical), where E is the eqv. mass of the
metal.

DETERMINATION OF ATOMIC MASS :
(i) **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

(ii) **Vapour density method** is suitable for elements having volatile chlorides. Atomic mass = Eq. mass of metal × valency.

(iii) **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: $ZnSO_4 . 7H_2O$, $MgSO_4 .7H_2O$ and $FeSO_4 .7H_2O$ are isomorphous.

- **Stoichiometric problems**
- Generally problems based on stoichiometry are of the following types :
- a. Problems based on mass-mass relationship;
- b. Problems based on mass-volume relationship and
- c. Problems based on volume-volume relationship.
- d. Eudiometry or “gas analysis” involves a calculation based on gaseous reactions in which at least two components are gases & their amount is given in terms of volumes measured at same pressure & Temperature.

The relationship amongst gases, when they react with one another, is governed by two laws, namely Gay-Lussac law and Avogadro's law.

Limiting reagent: The reactant which gets consumed, limits the amount of product formed and is therefore, called the limiting reagent

The Principle of Atom Conservation (POAC) :
In chemical reaction atoms are conserved , so moles of atoms shall also be conserved. This is known as principle of atomic conservation. This principle is helpful in solving problems of nearly all stoichiometric calculations e.g.

$KClO_3(s) \rightarrow KCl(s) + O_2(g)$

Applying POAC for K atoms

Moles of K atoms in $KClO_3$ = Moles of K atoms in KCl

Since one mole of $KClO_3$ contains 1 mol of K atom. Similarly 1 mol of KCl contains one mole of K atoms.

$$1 \times n_{KClO_3} = 1 \times n_{KCl}$$

$$1 \times \frac{W_{KClO_3}}{M_{KClO_3}} = 1 \times \frac{W_{KCl}}{M_{KCl}}$$

(Mass-mass relationship)
Applying POAC for ' O ' atoms

Moles of O atom in $KClO_3$ = Moles of O atom in O_2

$3 \times n_{KClO_3} = 2 \times n_{O_2}$

$$3 \times \frac{W_{KClO_3}}{M_{KClO_3}} = 2 \times \frac{Volume\ of\ O_2\ at\ STP}{Standard\ Molar\ Volume}$$

(Mass volume relationship of reactant and product)

In this way applying POAC we can break the chemical equation into a number arithmetic equations without balancing the chemical equation.

Moreover number of reactions and their sequence from reactants to products are not required. It is important to note that POAC can be applied for the atoms which remain conserved in chemical reaction.

Multiple Choice Questions:

1. 25.4 g of I_2 and 14.2 g of Cl_2 are made to react completely to yield a mixture of ICl and ICl_3 . Calculate moles of ICl and ICl_3 formed

(a) 0.1, 0.1 (b) 0.2, 0.2 (c) 0.1, 0.2 (d) 0.2, 0.1

2. In the final answer of the expression

$(29.2 - 20.2)(1.79 \times 10^5) / 1.37$ the number of significant figures is

(a) 1 (b) 3 (c) 2 (d) 4

3. The vapour density of a gas is 11.2. The volume occupied by 11.2 g of the gas at NTP will be

(a) 22.4 L (b) 11.2 L (c) 1 L (d) 44.8 L

4. The percentage of Se in peroxidase anhydrous enzyme is 0.5% by weight (atomic weight = 78.4). Then minimum molecular weight of peroxidase anhydrous enzyme is

(a) 1.568×10^3 (b) 1.568×10^4 (c) 15.68 (d) 3.136×10^4

5. Equivalent weight of crystalline oxalic acid is

(a) 90 (b) 53 (c) 63 (d) 45

6. 3 g of an oxide of a metal is converted to chloride completely and it yielded 5 g of chloride. The equivalent weight of the metal is

(a) 3.325 (b) 33.25 (c) 12 (d) 20

7. Number of moles of $KMnO_4$ required to oxidize one mole of $Fe(C_2O_4)$ in acidic medium is

(a) 0.167 (b) 0.6 (c) 0.2 (d) 0.4

8. 100 cm^3 of 0.1 N HCl is mixed with 100 cm^3 of 0.2 N NaOH solution. The resulting solution is

(a) 0.1 N and the solution is basic (b) 0.1 N and the solution is acidic

(c) 0.05 N and the solution is basic (d) 0.05 N and the solution is acidic

9. In order to prepare one litre normal solution of $KMnO_4$, how many grams of KMnO4 are required if the solution is to be used in acid medium for oxidation?

(a) 158 g (b) 62.0 g (c) 31.6 g (d) 790 g

10. An aqueous solution of 6.3 g of oxalic acid dihydrate is made up to 250 ml. The volume of 0.1 N NaOH required to completely neutralise 10 ml of this solution is

(a) 20 ml (b) 40 ml (c) 10 ml (d) 4 ml

11. If potassium chlorate is 80% pure, then 48 gm of oxygen would be produced from (atomic mass of K =39)

(a) 153.12gm of $KClO_3$ (b) 122.5 gm of $KClO_3$ (c) 245 gm of $KClO_3$

(d) 98 gm of $KClO_3$

12. If 0.5 mol of $BaCl_2$ is mixed with 0.2 mol of Na_3PO_4 then maximum number of moles of $Ba_3(PO_4)_2$ that can be formed is

(a) 0.7 (b) 0.5 (c) 0.3 (d) 0.1

13. If 3.01×10^{20} molecules are removed from 98 mg of H_2SO_4, then the number of moles of H_2SO_4 left are

(a) 0.1×10^{-3} (b) 0.5×10^{-3} (c) 1.66×10^{-3} (d) 9.95×10^{-2}

14. Gastric juice contains 3.0 g of HCl per litre. If a person produces 2.5 litre of gastric juice per day. How many antacid tablets each containing 400 mg of $Al(OH)_3$ are needed to neutralize all the HCl produced in one day ?

(a) 18 (b) 14 (c) 20 (d) 17

15. A 100 ml solution of 0.1 N HCl was titrated with 0.2 N NaOH solution. The titration was discontinued after adding 30 ml of NaOH solution. The remaining titration was completed by adding 0.25 N KOH solution. The volume of KOH required for completing the titration is

(a) 16 ml (b) 32 ml (c) 35 ml (d) 70 ml

16. 6.02×10^{20} molecules of urea are present in 100 mL of its solution. The concentration of solution is :

(a) 0.01 M (b) 0.001 M (c) 0.1 M (d) 0.02 M

17. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g. of CO_2 . The empirical formula of the hydrocarbon is :

(a) C_2H_4 (b) C_3H_4 (c) C_6H_5 (d) C_7H_8

18. Two oxides of a metal contain 50% and 40% metal (M) respectively. If the formula of first oxide is MO_2 the formula of second oxide will be

(a) MO_2 (b) MO_3 (c) M_2O (d) M_2O_5

19. In which of the following number all zeros are significant?

(a) 0.0005 (b) 0.0500 (c) 50.000 (d) 0.0050

20. The number of moles of $KMnO_4$ reduced by one mole of KI in alkaline medium is:

(a) One (b) two (c) five (d) one fifth

Answer for MCQ

1 a	11 a
2 b	12 d
3 b	13 b
4 b	14 b
5 c	15 a
6 b	16 a
7 b	17 d
8 c	18 b
9 c	19 c
10 b	20 a