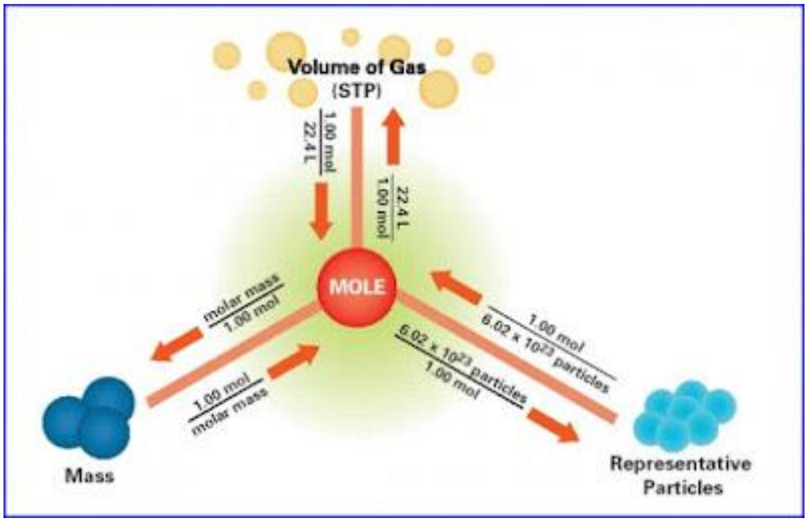
MITIHANI.CO.TZ



TOPIC 4: THE MOLE CONCEPT AND RELATED CALCULATIONS

The Mole as a Unit of Measurement

The Mole with Other Units of Measurements

Compare the mole with other units of measurements

When carrying out an experiment, a chemist cannot weigh out a single atom, ion, electron, proton or molecule of a substance. These particles are simply very small. A counting unit that is useful in practical chemistry must be used.

The standard unit is called one **mole** of the substance. One mole of each of these different substances contains the same number of the particles (atoms, molecules, ions, electrons, protons, neutrons, etc). That number per mole has been worked by several different experimental methods and is found to be 6.0×10^{23}. The value 6.0×10^{23} is called **Avogadro's constant** or Avogadro's number and is abbreviated as L. It is named after the nineteenth-century Italian chemist, Amedeo Avogadro.

The value 6.0×10^{23} is obtained through the following relationship.The mass of one atom of carbon-12 is 1.993×10^{-23}g. Then, the number of atoms present in 12g of carbon-12 is derived as follows:

1 atom = 1.993×10^{-23}g

X atoms = 12g

$$\text{X atoms} = \frac{12g \times 1}{1.993 \times 10^{-23}} = 6.02 \times 10^{23}$$

X = 6.0×10^{23} atoms.

Therefore, the number of atoms in 12g of carbon-12 and hence the number of particles in a mole are 6.02×10^{23} atoms.

Hence, Avogadro's number is the number of atoms in exactly 12g of carbon-12 isotope.One mole of any substance contains as many as many elementary particles as the Avogadro's number (constant).

So, from the above explanation, the **mole** can be defined as *the amount of a substance that contains as many elementary particles as the number of atoms present in 12g of carbon-12 isotope.*

Substance	**Formula**	**Relative formula mass, M_r**	**Mass of one mole (molar mass)**	**This mass (1 mole) contains**
Carbon	C	12	12g	6.0×10^{23} carbon atoms
Iron	Fe	56	56g	6.0×10^{23} iron atoms
Hydrogen	H_2	$2 \times 1 = 2$	2g	6.0×10^{23} molecules
Oxygen	O_2	$2 \times 16 = 32$	32	6.0×10^{23} molecules
Water	H_2O	$(2\times1) + 16 = 18$	18g	6.0×10^{23} formula units
Magnesium oxide	MgO	$24 + 16 = 40$	40g	6.0×10^{23} formula units
Calcium carbonate	$CaCO_3$	$40+12+(3\times16) = 100$	100g	6.0×10^{23} formula units
Silicon oxide	SiO_2	$28 + (2 \times 16) = 60$	60g	6.0×10^{23} formula units
Fe^{3+}	Fe^{3+}	56	56g	6.0×10^{23} iron(III) ions
Cl^-	Cl^-	35.5	35.5g	6.0×10^{23} chloride ions
e^-	e^-	–	–	6.0×10^{23} electrons

The other substances, which also exist as molecules, include ozone molecule (gas), O_3; phosphorus molecule (solid), P_4; sulphur molecule, S_8, etc.

In real life, when dealing with large numbers of small objects, it is usual to count them in groups. The objects are grouped and counted in unit amounts. For example, we buy a carton of soap, a gallon of kerosene, a crate of soda, a dozen of pencils, a ream of papers, etc.

Some units of measurement

Unit	Number of objects per unit
Pair	1 pair = 2 objects, e.g. gloves, shoes, socks, scissors, etc are always sold in pairs.
Dozen	1 dozen = 12 objects e.g. a dozen of cups, plates, spoons, etc.
Gross	1 gross = 144 objects, e.g. a box of blackboard chalk contains 144 pieces of chalk.
Ream	1 ream = 500 objects, e.g. papers are sold in reams of 500 sheets.
Mole	1 mole = 6.02×10^{23} particles. In chemistry, extremely small particles are expressed in moles. For example:1 mole of atoms = 6.02×10^{23} atoms1 mole of electrons = 6.02×10^{23} electrons1 mole of protons = 6.02×10^{23} protons1 mole of ions = 6.02×10^{23} ions1 mole of molecules = 6.02×10^{23} molecules

Molar Quantities of Different Substances

Measure molar quantities of different substances

The mass of one mole of any substance (its molecular mass) is the atomic mass or molecular mass expressed in grams (or kilograms). For convenience, chemists prefer to express mass in grams, although the SI unit of mass is the kilogram. This is because the amount of substances which chemists usually work with in science laboratories, is quite small and if their masses are expressed in kilograms, the numbers used would be extremely small.

You can calculate the molar mass (M) of any substance by summing up the relative atomic weights of its constituents atoms. For example, ethanol, C_2H_5OH, contains two carbon atoms, six hydrogen atoms and one oxygen atom. So, the molar mass of ethanol can be calculated thus: Molar mass of C_2H_5OH = (2 × 12) + (6×1) + 16 = 46g.

In a similar way, molar masses of other compounds can be calculated. For example, the molar mass of sodium chloride, NaCl, is calculated by adding together the relative atomic masses of the constituents elements (Na = 23 and Cl = 35.5) = 23 + 35 = 58.5g (g mol^{-1}).

It is important to note that *relative atomic mass* or *relative molecular mass* has *no unit* while molar masses are always expressed in grams or kilograms.

The *molar mass of a compound* is the same as the *relative molecular mass* and the **molar mass of an element** is the same as the **relative atomic mass** (A_r) of that element. The only difference lies in the units.

Example 1

- $M(CO_2)$ = 44g (or g mol^{-1}) = molar mass of carbon dioxide
- $M_r(CO_2)$ = 44 = relative molecular mass of carbon dioxide
- M(Fe) = 56g (or g mol^{-1}) molar mass of iron
- M_r(Fe) = 56 = Relative atomic mass of iron

Similarly, the molar masses of each of the following substances can be calculated using values for the relative atomic masses of the elements.

Molar masses of different substances

Substance	**Formula**	**Molar mass**
Ammonia	NH_3	14 + 1×3 = 17g
Ammonium chloride	NH_4Cl	14 + (1×4) + 35.5 = 53.5g
Lead (II) nitrate	$Pb(NO_3)_2$	207 + (14×2) + (16×6) = 331g
Sulphuric acid	H_2SO_4	(1×2) + 32 + (16×4) = 98g
Calcium carbonate	$CaCO_3$	40 + 12 + (16×3) = 100g
Potassium dichromate	$K_2Cr_2O_7$	(39×2) + (52 ×2) + (16×7) = 294g

Application of the Mole Concept

Known Masses of Elements, Molecules or Ions to Moles

Convert known masses of elements, molecules or ions to moles

In experimental work, chemists work with varying masses. They cannot always use one mole of a substance. The equation that links the mass of a substance to the number of moles present is:

$$\text{Number of moles (n)} = \frac{\text{Mass}}{\text{Molar mass}}$$

Example 2

Convert 49g of sulphuric acid, H_2SO_4, into moles.*Given:*Mass = 49g; molar mass = 98g

Formula:

$$\text{Number of moles (n)} = \frac{\text{Mass}}{\text{Molar mass}}$$

*Solution:*49g of H_2SO_4= 49/98= 0.5 mol.

Known Volumes of Gases at S.T.P to Moles

Convert known volumes of gases at S.T.P to moles

The volume occupied by one mole of a gas at standard condition of temperature and pressure has been scientifically determined, and it is found to be 22.4 dm^3. This volume is called the molar volume of a gas. The molar volume of a gas, therefore, has the value of 22.4 dm^3 at s.t.p. Remember that 1 dm^3 (1 litre) = 1000 cm^3. One important thing about this value is that it applies to all gases. Therefore, at s.t.p. 32g of oxygen (O_2) or 17g of ammonia (NH_3) or 44g of carbon dioxide (CO_2) or 40g of argon (Ar) will occupy a volume of 22.4 dm^3. This makes it easy to convert the volume of any gas at s.t.p. into moles, or moles into volume. However, it is important to note that as the conditions of temperature and pressure change the molar volume will also change.

The number of moles of a given sample of gas is obtained by dividing the volume of the gas by molar volume (22.4 dm^3).

$$Number\ of\ moles = \frac{\text{Volume}}{\text{Molar volume}}$$

For example**,** 4.4d m^3 of carbon dioxide gas at s.t.p. = 4.4/22.4= 0.196 mol.Similarly, 2.24 dm^3 of neon gas at s.t.p. = 2.24/22.4= 0.1 mol.

If the volume of the gas is given in cm^3, then it should be divided by the molar volume of a gas expressed in cm^3. For example, 560 cm^3 of nitrogen gas = $560cm^3/22400cm^3$ mol= 0.025 mol.

Alternatively, the volume may, first, be converted to dm^3 and then divides by the molar volume, expressed in dm^3, that is, $0.46dm^3/22.4dm^3$ = 0.25mol

Masses of Solids or Volumes of Known Gases to Actual Number of Parties

Change masses of solids or volumes of known gases to actual number of parties

The number of particles in one mole of any substance is 6.02×10^{23}. To find the number of particles in a substance, we use the expression:

- N = n.L, where
- N = the number of particles in that substance;
- n = the amount of substance (moles); and
- L = the Avogadro’s constant (6.02×10^{23}).

This conversion requires two steps: first convert the mass of solid or volume of gas to moles, and then multiply the number of moles by the Avogadro’s constant. For example, to convert 5.6 dm^3 of ammonia gas to the actual number of ammonia (NH_3) molecules, change 5.6 dm^3 of ammonia to moles =$0.46dm^3/22.4dm^3$=0.25 mol. Then multiply by the Avogadro’s constant to get the total number of molecules$0.25 \times 6.02 \times 10^{23} = 1.5 \times 10^{23}$ molecules

Similarly, 1.12 dm^3 of hydrogen gas = 1.12/22.4= 0.05 mol. This is equal to $0.05 \times 6.02 \times 10^{23} = 3.0 \times 10^{22}$ molecules

Alternatively, we may find out the number of particles by converting the given volume to the number of molecules straight forward without passing through the number of moles first. We know that one mole (22.4 dm^3) of a gas at s.t.p. = 6.02×10^{23} molecules. So, 5.6 dm^3 = $5.6 \times 6.02 \times 10^{23}/22.4 = 1.5 \times 10^{23}$ molecules

Molar Solutions of Various Soluble Substances

Prepare molar solutions of various soluble substances

A molar solution is a solution which contains one of the compound in one litre (1 dm^3 or 1000 cm^3) of the solution.Let us consider the case of sodium hydroxide, NaOH. The molecular weight of this compound is 40g. Therefore, a molar solution of sodium hydroxide will contain 40g in 1000 cm^3 (1 dm^3) of the solution.

Also, consider anhydrous sodium carbonate, Na_2CO_3. 1 mole of this carbonate weights 106g. Hence, its molar solution will contain 106g of the anhydrous salt in 1000 cm^3 of solution.If, however, 0.1 moles (10.6g) of the solute is dissolved in 1.0 dm^3, the solution is 0.1 molar. But if 0.1 moles is dissolved in 0.1 dm^3 of the solution, the solution is still 1.0 molar (since 1 dm^3 of solution would contain 1.0 mole of the solute).

The molecular weights of some common substances are shown below:

Compound	**Molecular weight (1 mole)**
Potassium hydroxide, NaOH	56g
Hydrochloric acid, HCl	36.5g
Sulphuric acid, H_2SO_4	98g
Sodium chloride, NaCl	58.5g
Sodium bicarbonate, $NaHCO_3$	84g
Calcium hydroxide, $Ca(OH)_2$	74g

The molar solution of each of these substances can be prepared by dissolving one mole of each substance in 1000 cm^3 (1 dm^3) of distilled water. We see, therefore, that 40g of sodium hydroxide in 1000 cm^3 of solution will give a 1.0M solution. Hence, 20g of the hydroxide should give a 0.5M solution. In a similar way, we can make derivative solution concentrations ranging as follows: 0.1M, 0.2M, 0.3M, 0.4M….1M, 2M, etc.

However, in each case the amount of solution should always be 1000 cm^3. The concentration ranges like these are known as *molarities* of solutions. Hence, 0.5M sodium carbonate can also be read as “a sodium carbonate solution with a molarity of 0.5M.”

The Concentration of Solutions

When a chemical substance (the solute) is dissolved in a given volume of solvent, we can measure the “quantity” of solute in two ways; we can measure either its **mass** (in grams) or its **amount** (in moles). The final volume of the solution is usually measured in dm^3.

When we measure the **mass** of the solute in **grams**, we obtainthe **mass concentration** in g/dm^3

$$\text{Concentration } (g/dm^3) = \frac{\text{Mass of solute(g)}}{\text{Volume of solution}(dm^3)}$$

Example 3

Calculate the concentration (g/dm^3) of sodium chloride solution (NaCl) that contains 20g of sodium chloride in a final solution of 100 cm^3

Solution

First, convert the given volume to dm^3

Volume (dm^3) = 100/1000= 0.1 dm^3

Then, work out the concentration of the solution by dividing the mass (weight) of solute (g) by the volume (dm^3).

$$\text{Concentration } (g/dm^3) = \frac{\text{Mass of solute(g)}}{\text{Volume of solution}(dm^3)}$$

$=20g/0.1dm^3$

$= 200g/dm^3$

Alternatively, we could calculate the concentration straightforward without having to convert the given volume into dm^3, e.g.:If 20g of the solution are contained in 100 cm^3 of the solution, then the amount of solute in 1000 cm^3 (1 dm^3) of the solution would be

$1000 \times 20/100 = 200g/dm^3$

Calculations Based on the Mole Concept

Perform calculations based on the mole concept

A chemist always wants to know how much of one substance would react with a given amount of another substance. This is achieved by use of balanced chemical equations. Such equations are called stoichiometric equations.

A stoichiometric equation is the one in which the reactants and the products are correctly balanced; all the atoms, ions and electrons are conserved. Such an equation gives correct mole ratios of reactants and products in chemical reactions. This quantitative relationship is called stoichiometry.

Consider an equation for the reaction between hydrogen and nitrogen to produce ammonia:

$3H_{2(g)} + N_{2(g)} \rightarrow 2NH_3$

This can be read as follows:

three *moles of hydrogen reacts with* ***one*** *mole of nitrogen to yield* ***two*** *moles of ammonia.*

The numbers 3, 1 and 2 are called stoichiometric coefficients. They tell us the proportions in which the substances react and in which the products are formed.

Example 4

What volume of carbon dioxide (CO_2) measured at s.t.p. will be produced when 21.0g of sodium hydrogencarbonate ($NaHCO_3$) is completely decomposed according to the equation.$2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(l)}$

Solution

First, find the weight of carbon dioxide that will be produced by the hydrogencarbonate.

- Mass of $2NaHCO_3 = 2 \times 84 = 168g$
- Mass of $CO_2 = 44g$

The weight of carbon dioxide produced can be obtained from the following relation:

$168g \equiv 44g$

$21g \equiv X$

$X = 21 \times 44/168 = 5.5g$

The weight of carbon dioxide produced = 5.5g

Then, convert this weight of CO_2 to volume at s.t.p.We know that one mole (44g) of carbon dioxide at s.t.p. occupies 22.4 dm^3

That is, $44g \equiv 22.4dm^3$

$5.5g \equiv X\ dm^3$

$X = 5.5 \times 22.4/44 = 2.8 dm^3$