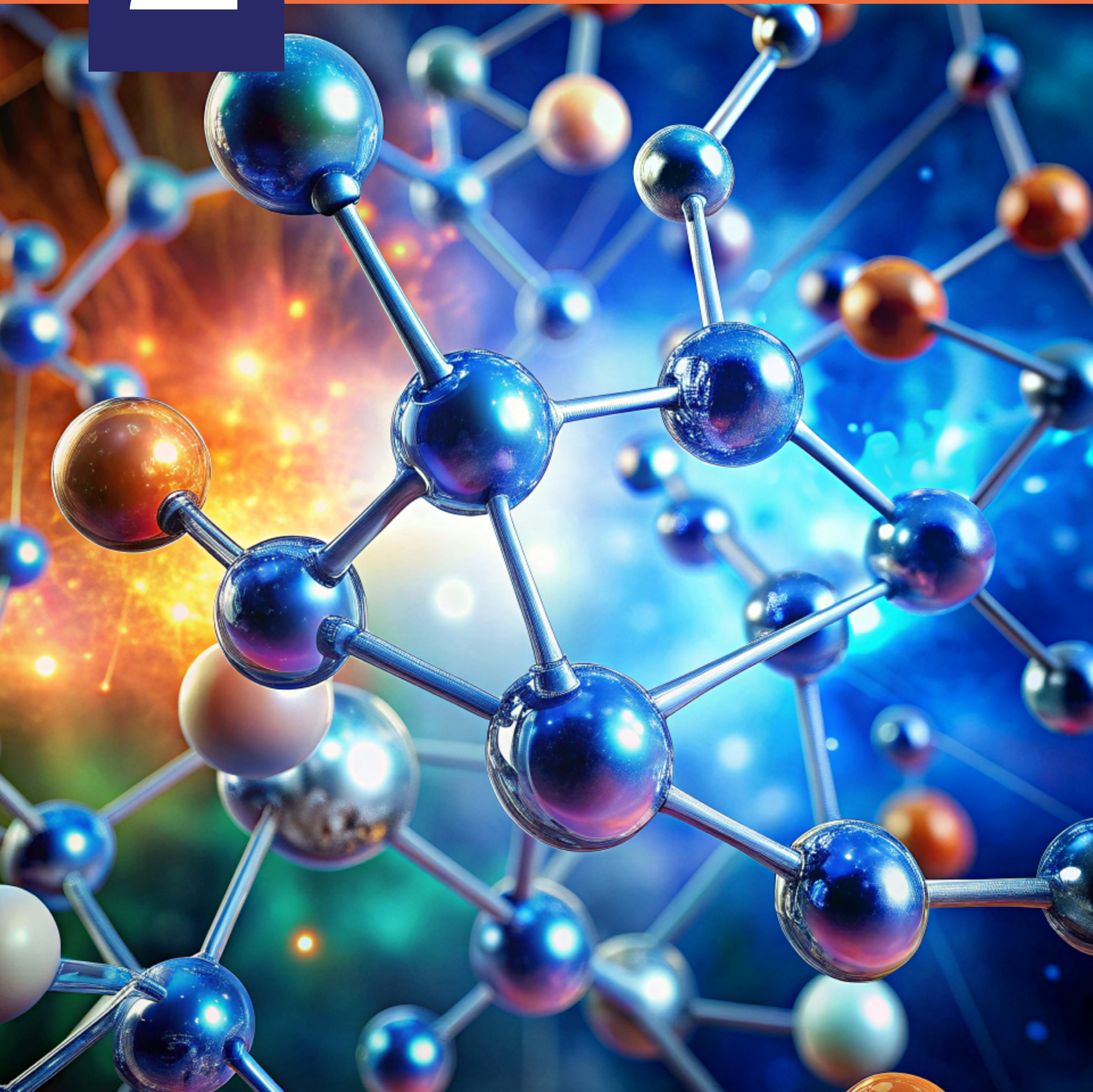
Chemistry Year 1

SECTION 2

THE CONCEPT OF THE MOLES



PHYSICAL CHEMISTRY

Matter and Its Properties

INTRODUCTION

In this section you will be introduced to fundamental concepts such as relative atomic mass, relative molecular mass, and the mole as a unit of amount of substance. You will also learn how to perform calculations based on the amount of substance and understand the importance of the mole concept in preparing standard solutions.

At the end of this section, you should be able to:

- Explain relative atomic mass and relative molecular mass.
- Describe the atomic mass unit as an average mass.
- Describe the mole as a unit of the amount of substance.
- Calculate different physical quantities (number of entities, mass and volume) based on the amount of substance.
- Explain the mole concept and its relevance in preparation of standard solutions.

Key Ideas

- **Atomic mass unit (amu)** is equal to 1/12th the mass of a carbon-12 atom.
- **Relative atomic mass** is the average mass of an element's atoms, considering isotopes.
- **Relative molecular mass:** the sum of the relative atomic masses of all atoms in a molecule.
- **Mole** is a way of measurement in chemistry.
- **Avogadro's constant** is number of units in one mole of any substance.
- **Entities** refers to any distinct atom, molecule, ion.
- **Avogadro's number** is the number of particles in one mole.

- **Avogadro's number** is 6.022×10^{23} particles (atoms, molecules, and ions)
- **Mole** (mol) is a unit of measurement.
- **Molar mass** is mass of mole a substance.

RELATIVE ATOMIC MASS (A_r)

Relative Atomic Mass (A_r) is defined as the average mass of one atom of the element compared to $1/12^{th}$ of the mass of one atom of carbon-12.

Mathematically Relative Atomic mass is represented as,

$$A_r = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12^{th}}\text{the mass of one atom of carbon} - 12}$$

This value is expressed in atomic mass units (amu). The relative atomic mass is determined by the abundance of each isotope of an element.

Worked Example 2.1

One atomic mass unit of Carbon – 12 is 1.6603×10^{-24} g. If the average mass of an atom of oxygen is 2.65659×10^{-23} g. Determine its relative atomic mass.

Solution:

$$A_r(\text{O}) = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12^{th}}\text{the mass of one atom of carbon} - 12}$$

$$= \frac{2.65659 \times 10^{-23}\ \text{g}}{1.6603 \times 10^{-24}\ \text{g}}$$

$$= 16.0$$

Relative Molecular Mass (M_r)

Relative molecular mass is defined as the average mass of one molecule of a substance compared with $\frac{1}{12^{th}}$ of the mass of one atom of carbon-12.

$$M_r = \frac{\text{average mass of one molecule of the substance}}{\frac{1}{12^{th}}\text{ the mass of one atom of carbon} - 12}$$

It also has no unit. For ionic compounds, the relative molecular mass is called its **Relative formula mass** (as ionic substances do not exist as molecules).

Relative molecular mass is the sum of the masses of the elements that make up the molecule.

Activity 2.1: Determining the mass of an element or compound using a beam balance

Materials needed:

Beam balances, standard carbon-12 samples (represented by 12 beads or any small, identical objects), samples of different elements or compounds (using different numbers of beads/objects), worksheets for recording observations and calculations, calculators

Steps:

Carry out this activity in small groups.

1. Set up the beam balance, get a standard carbon-12 sample (12 beads), and samples of other elements or compounds (different numbers of beads).
2. Place the carbon-12 sample in one pan of the beam balance.
3. Place the unknown element or compound sample in the other pan of the beam balance.
4. Adjust the number of beads/objects in the unknown sample until the beam balance is level, indicating that the masses are equal.
5. Record the number of beads/objects used for the unknown sample to balance the carbon-12 standard.
6. Calculate the relative mass of their unknown sample compared to carbon-12.

   Example Calculation:

   If the unknown sample balanced with 22 beads, the relative mass compared to carbon-12 (12 beads) is:

   $$A_r = \frac{\text{Average mass of one atom of the element}}{\frac{1}{12^{th}}\text{the mass of one atom of carbon} - 12}$$

   So then as a balance is set up:

   12 × (mass of 1 carbon atom) = 22 × (mass of 1 unknown atom)

   So: as we know the mass of a carbon atom is 1 amu we can state that:

   $$A_r = \frac{\text{Number of beads in carbon} - 12}{\text{Number of beads in unknown sample}} = \frac{12}{22} = 0.54$$

Therefore, the unknown sample has a relative mass of 0.54 times that of carbon-12.

7. Record your calculations and results on the worksheet.
8. Present your findings and explain your calculations.

Worked Example 2.2

Determining mass using a beam balance.

(a) Number of beads in carbon-12 sample: 12

(b) Number of beads in unknown sample: x

(c) Calculate the relative mass of the unknown sample compared to Carbon-12.

$$A_r = \frac{12}{\text{Number of beads in unknown sample}} = \frac{12}{x} = y$$

Therefore, the relative mass of the unknown sample is y.

Activity 2.2: Understanding relative molecular mass

Materials needed: different sets of coloured balls,

Steps:

Carry out this activity in small groups.

1. Provide different sets of coloured balls and connectors.
2. Construct different molecules such as H_2O, CO_2, and CH_4.
3. Create a "molecule" by connecting a few balls (e.g., two blue balls and one red ball to represent H_2O).
4. Calculate the relative molecular mass of the constructed molecules using the atomic masses provided.

Calculating the M_r of a compound using A_r:

Worked Example 2.3

Show the steps to calculate the relative molecular mass of water.

Solution:

1. Identify the chemical formula of water i.e. H_2O.
2. Using the relative atomic mass of each element multiply the relative atomic mass by the number of atoms of each element:
   a. In water (H_2O), there are 2 hydrogen atoms and 1 oxygen atom.
   b. The relative atomic mass of hydrogen = 1.0
   c. Number of hydrogen atoms = 2
   d. Contribution of hydrogen to the relative molecular mass = $1.0 \times 2 = 2.0$
   e. Relative atomic mass of oxygen = 16.0
   f. Number of oxygen atoms = 1
   g. Contribution of oxygen to the relative molecular mass = $16.0 \times 1 = 16.0$
3. Add the contributions from all elements:

   Sum the contributions of hydrogen and oxygen to get the relative molecular mass of water:

   $M_r(H_2O)$ = Contribution of hydrogen + Contribution of oxygen

   $M_r(H_2O) = 2.0 + 16.0 = 18.0$

Activity 2.3: Trial Question

Calculate the relative molecular masses (M_r) of the following substances:

For reference, the atomic masses are approximately: H = 1; O = 16; C = 12

1. Water (H_2O)
2. Carbon Dioxide (CO_2)
3. Methane (CH_4)
4. Glucose ($C_6H_{12}O_6$).

THE ATOMIC MASS UNIT (AMU)

The relative atomic mass scale is based on an isotope of carbon-12. Carbon-12 contains 6 protons and 6 neutrons and a mass of 12 atomic mass units. The carbon-12 scale is therefore defined as an atomic mass reference scale in which one atom of carbon-12 isotope has 12 units.

Therefore,

The mass of one carbon-12 = 12 amu.

$$1 \text{ amu} = \frac{\text{Mass of one carbon} - 12\_}{12}$$

NB*: One atomic mass unit of Carbon-12 is the same as* 1/12th of the mass of one atom of carbon-12.

Recall that most naturally occurring elements have different isotopes with different natural abundance and masses. Therefore, relative atomic mass is an average mass.

Applications of Relative Atomic Mass in everyday life

1. Relative atomic and relative molecular mass is used to calculate the concentration of a stock solution from chemical stores.
2. The idea of relative molecular mass or formula mass and the law of conservation of mass are used to do quantitative calculations in chemistry.
3. The idea of relative atomic mass is used to determine the empirical formula of a substance.

Activity 2.4: How to calculate the atomic mass unit (amu) of an individual particle (atom or molecule)

Materials needed:

- Access to a computer or smart device with internet
- Worksheet for notes and questions
- Access to an educational video on atomic mass units (amu) (e.g., a video science education channel on YouTube)

How to Calculate
Atomic Mass Practice

https://www.youtube.com/watch?v = ULRsJYhQmlo

Steps:

1. Watch the video and take notes on key points, especially on how the amu is defined and measured.
2. Pause the video at key moments to discuss important concepts and ensure understanding.
3. After watching the video, discuss the following questions:

   a. Write a brief definition of an atomic mass unit (amu).
   b. Explain how the carbon-12 isotope is used to define the amu.
   c. Describe how the mass of a single atom or molecule is measured in amu.
   d. Why is it important to have a standard unit like the amu in chemistry?

THE MOLE AS A UNIT OF THE AMOUNT OF SUBSTANCE

In everyday life, units such as pair and dozen are used to represent a specific number of items. Scientists use the term **mole** to represent a specific number of elementary entities (atoms, ions or molecules).

One mole of a substance is defined as the amount of substance that contains as many elementary entities as there are atoms in 12 g of the carbon-12 isotope.

The term "elementary entities" refers to the basic units that make up a substance. These can include atoms, molecules, ions, electrons, protons etc.

1 mole of every substance contains 6.02×10^{23} elementary entities.
For example, 1 mole of magnesium metal contains 6.02×10^{23} atoms inside it.

How do you determine the number of particles (N) of a substance contained in each number of moles (n)?

Number of formula units in 1 mole of any substance $= 1 \times 6.02 \times 10^{23}$
Number of formula units in 2 moles of any substance $= 2 \times 6.02 \times 10^{23}$

Number of formula units in n mole of any substance = $n \times 6.02 \times 10^{23}$

$N = n \times 6.02 \times 10^{23}$

But 6.02×10^{23} is termed **Avogadro's number** or **constant** and it is denoted by N_A or L.

Mathematically,

$$n = \frac{N}{N_A}$$

or

$$n = \frac{N}{L}$$

Where,

n = number of moles (amount of substance) measured in mol. The mole is the base unit of the fundamental quantity called the amount of substance.

N = number of entities

L = Avogadro's number expressed as defined particles mol^{-1}. L is a molar quantity, that is, a quantity expressed per mole.

Worked Example 2.4

Calculate the number of moles contained in 9.5×10^{23} molecules of oxygen.
$[L = 6.02 \times 10^{23}]$

Solution:

Use the problem-solving approach.

Analyse the question

Known

Number of molecules

Avogadro's Constant

Formula to use: $n = \frac{N}{L}$

Unknown

Amount of substance

Solve: Apply the formula

$N = 9.5 \times 10^{23}$

$L = 6.02 \times 10^{23}$

$n = ?$

$n = \frac{N}{L}$

Substituting the values,

$n = \frac{9.5 \times 10^{23}\ molecules}{6.02 \times 10^{23}\ molecules^{-1}}$

$n = 1.58\ mol$

Activity 2.5: Trial Questions

Find the answers to the following questions:

1. How many moles are there in 1.204×10^{24} molecules of water (H_2O)?
2. Calculate the number of moles in 3.011×10^{22} atoms of helium (He).
3. If you have 5.000×10^{23} molecules of carbon dioxide (CO_2), how many moles do you have?

Activity 2.6: Exploring the mole concept

Materials needed:

Sample substances (e.g., salt, water, sugar); calculators, 2 small cups or containers, electronic balance, periodic table, table salt (NaCl), water (H_2O)

Steps:

1. Use a large poster or chart to show what a mole represents: 6.02×10^{23} elementary entities.
2. Use the periodic table to calculate the molar masses of given compounds:

Sample substance	Formula of compound	Molar mass in g/mol	Actual mass as measured in (g)
water	H_2O	H_2O 2 H:$2 \times 1 = 2$ 1 O:$1 \times 16 = 16$ $= 18\ g\,mol^{-1}$	18 g
Sodium chloride	NaCl		

3. Measure and record the mass of water equal to 1 mol in a container.
4. Determine the molar mass of NaCl.
5. Measure and record the mass of NaCl equal to 1 mol in another container
6. Calculate the number of moles of (H_2O and NaCl):

$$n(H_2O) = \frac{m(H_2O)}{M(H_2O)} = \frac{18\ g}{18\ g\,mol^{-1}} = 1.0\ mol$$

$$n(NaCl) = \frac{m(NaCl)}{M(NaCl)} = \frac{\quad g}{\quad g\,mol^{-1}} = \qquad mol$$

7. Count the number of atoms in each molecule:
   - Water (H_2O) has 2 hydrogen atoms and 1 oxygen atom.
   - Sodium chloride (NaCl) has 1 sodium ion and 1 chloride ion.

Explanation:

1 mol of water contains 6.02×10^{23} molecules of water,

which means it contains $2 \times 6.02 \times 10^{23}$ hydrogen atoms

and 6.02×10^{23} oxygen atoms.

Discussion Questions:

1. Do 1 mol of H_2O and 1 mol of NaCl have the same mass?
2. Would 1.50 mol of H_2O have the same number of particles as 1.50 mol of NaCl

CALCULATING THE NUMBER OF ENTITIES

Hello, learner, you are about to be introduced to how the amount of substance can be used to calculate different quantities such as the number of entities, mass and volume of gases. You will also learn how to relate the mole concept in the preparation of standard solutions.

Recall that,

$$\text{Amount of substance} = \frac{\text{Number of entities}}{\text{Avogadro'sconstant}}$$

$$n = \frac{N}{N_A} \quad \text{or}$$

$$n = \frac{N}{L}$$

To calculate for the number of entities, multiply both sides of the equation by L

$N = n \times L$

Number of Entities = number of moles of substance × Avogadro′s Constant

Worked Example 2.5

Calculate the number of atoms contained in 0.25 mol of sodium.
[$L = 6.02 \times 10^{23}$]

Solution (Using problem-solving strategy):

i. Analyse the question

Known:

$n = 0.25\ mol$

$L = 6.02 \times 10^{23}$

Unknown:

$N = ?$

Use the formula: $N = n \times L$

Solve: Apply the formula

By definition,

$N = n \times L$

$N = 0.25\ \text{mol} \times\ 6.02 \times 10^{23}$

$N = 1.51 \times 10^{23}$

- Evaluate: Check to see if the answer makes sense and if the correct unit is stated.

Moles in Mass of Atoms or Molecules

The Molar mass (M) of a substance, is the relative atomic mass (A_r) or relative molecular mass (M_r) expressed in grams per mole; e.g. the $M(H_2)$is 2 g mol^{-1}, or the $M(CaCO_3)$is 100 g mol^{-1}

How do you determine the mass of a given number of moles of a substance?

Mass of 1 mole of atom **X** = 1 × M(**X**)

Mass of 1 mole of O = 1 × 16 = 16g

Mass of 2 moles of O = 2 × 16 = 32g

Mass of n moles of O = $\mathbf{n} \times \mathbf{M} = \mathbf{m}$ g

$m = n \times M$

Amount of substance $\left(n\right) = \frac{\text{mass of substance (m)}}{\text{Molar mass (M)}}$

Worked Example 2.6

Calculate the number of moles contained in 20 g of Aluminium atoms [Al = 27]

Solution:

Analyse the question

Known:

Mass (m)= 20 g

Relative atomic mass $\left(A_r\right) = 27$

Molar mass = 27 g mol^{-1}

Formula to use:

Amount of substance $\left(n\right) = \frac{\text{mass of substance (m)}}{\text{Molar mass (M)}}$

Unknown:

Number of moles (n)

Solve: Apply the formula

By definition,

Substituting the values,

$$\text{Amount of substance (n)} = \frac{20\text{ g}}{27\text{ gmol}^{-1}}$$

$$= 0.74\ mol$$

Therefore, the number of moles of aluminium atoms is 0.74 mol

Calculating for the mass of a given amount of substance

Recall that, Amount of substance $\left(n\right) = \frac{\text{mass of substance (m)}}{\text{Molar mass (M)}}$

Making m the subject yields,

$m = n \times M$

Worked Example 2.7

Calculate the mass of 0.50 mol of water H_2O. [$H = 1$, $O = 16$]

Solution:

Analyse the question

Known:

Number of moles (n) = 0.5 mol

Relative Atomic masses [$H = 1$, $O = 16$]

Formula to use:

$m = n \times M$

Unknown:

Mass (m) =?

Solve: Apply the strategy

Calculate the M

$M\ (H_2O) = 2(1) + 16 = 18\ \text{g mol}^{-1}$

Calculate the mass

By definition, $m = n \times M$

$m = 0.5 \times 18$

$m = 9\ g$

Therefore, the mass of water is 9g.

Activity 2.7: Trial Question

Calculate the number of moles in 36 g of water (H_2O).

Quantity of Substance and Molar Volume of Gases

The volume occupied by a gas depends on:

1. Quantity of substance
2. Temperature
3. Pressure of the gas

At standard temperature of 273 K and pressure of 101.3 kPa (known as standard temperature and pressure, or s.t.p.), the volume occupied by one mole of any gas is called molar volume, denoted by V_m. V_m is a constant and has a value of 22.4 dm^3mol^{-1} at s.t.p.

The Molar Volume V_m, the number of moles of substance **n** and volume of gas are related by the formula:

$$\text{Amount of substance (n)} = \frac{\text{volume of substance in dm}^3\text{(V)}}{\text{Molar volume }(V_m)}$$

Multiplying both sides of the equation by V_m gives

$V = n \times V_m$

NB:

a. This equation is used to calculate the volume of a gas at s.t.p., given the quantity of substance or number of moles. The equation ($V = n \times V_m$) **cannot** be used for any other values of temperature or pressure.

b. If the gas volume is measured in cm^3, convert to dm^3.

c. You can calculate the volume of a named gas, given the formula and relative atomic masses of the elements.

Worked Example 2.8

Calculate the volume occupied by 0.75 mol of ammonia gas (NH_3) at s.t.p.
[$V_m = 22.4\ dm^3 mol^{-1}$]

Solution:

Use the problem-solving approach

Known:

Mole (n) = 0.75 mol

$V_m = 22.4\ dm^3 mol^{-1}$

Formula to use: $V = n \times V_m$

Unknown:

Volume of gas, V =?

Solve: Apply the problem-solving approach

By definition,

$V = n \times V_m$

Substituting the values,

$V = 0.75\ mol \times 22.4\ dm^3 mol^{-1}$

$V = 16.8\ dm^3$

The concept map of the relationship between mole and other variables are:

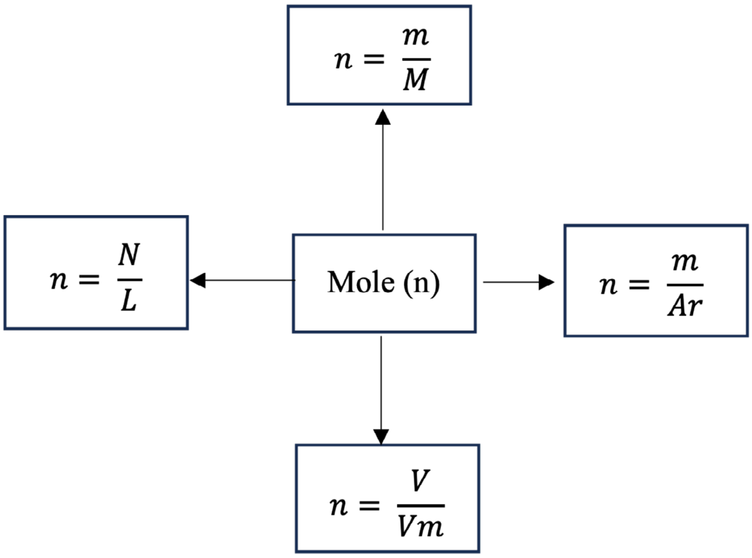


Fig. 2.1: A concept map showing the relationship between mole and other variables

Activity 2.8: Trial Question

Calculate the number of moles of oxygen gas (O_2) present in 44.8 dm^3 at standard temperature and pressure (stp).

Calculating the number of entities, mass and volume of a gas using mathematical equations

You can use a periodic table, calculator and worksheets with specific problems to help you to calculate or determine the number of entities (ion, atoms, molecules, etc), mole and mass of gases.

1. Conversions between Mass and Number of Particles

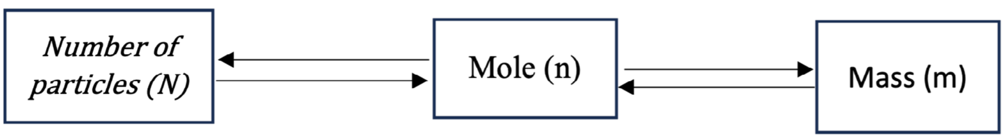


Fig. 2.2: Interrelationships between mass, number of particles and moles.

Figure 2.2 illustrates that mass, number of particles, and moles are all interrelated. To convert between mass and number of particles, a conversion to moles is required first.

2. Converting Mass to Number of Particles

Worked Example 2.9

How many molecules are present in a 17.5 g sample of P_4O_{10}? [P = 31, O = 16]

Solution:

Step 1: List the known quantities and plan the problem.

Known:

sample mass = 17.5 g P_4O_{10}

molar mass of O_2 = 284 g/mol

Unknown:

number of molecules of P_4O_{10}

Working:

- First, convert the mass of P_4O_{10} to moles.
- Second, convert moles of P_4O_{10} to the number of molecules.
- **Step 2: Calculate**

$$N\left(P_4O_{10}\right) = n\left(P_4O_{10}\right) \times L = \frac{m(P_4O_{10})}{M(P_4O_{10})}L$$

$$= \frac{17.5\text{ g}}{284\text{ g mol}^{-1}} \times 6.02 \times 10^{23}\text{ molecules mol}^{-1}$$

$$N(P_4O_{10}) = 3.7 \times 10^{22}\text{ molecules}$$

Worked Example 2.10

Calculate the number of sodium ions present in 2.5 g of Na_2SO_4.
[Na = 23, S = 32, O = 16, L = 6.02×10^{23} *particles mol*$^{-1}$]

Solution:

- **Step 1**: Identify the molar mass of Na_2SO_4
  Na: $2 \times 23.0 = 46.0$ g/mol
  SO_4^{2-}: $32. + (4 \times 16.0) = 96.0$ g/mol
  Total molar mass: $46.0 + 96.0 = 142.0$ g/mol

- **Step 2**: Calculate the number of moles of Na_2SO_4:

$$n(Na_2SO_4) = \frac{m(Na_2SO_4)}{M(Na_2SO_4)}$$

$$= \frac{2.5\ g}{142.0\frac{g}{mol}} = 0.0176\ mol$$

- **Step 3**: Identify the number of sodium ions (Na^+) in the formula (Na_2SO_4): 2
- **Step 4**: Calculate the total number of moles of sodium ions:

$$Na_2SO_4 \longrightarrow 2\ Na^+ + SO_4^{2-}$$

$$\frac{0.0176\ \text{mol}\ Na_2SO_4}{1\ \text{mol}\ Na_2SO_4} \times 2\ \text{mol}\ 2\ Na^+ = 0.0352\ \text{mol}\ Na^+$$

- **Step 5**: Convert moles of sodium ions to number of particles using Avogadro's number (L):

$$\text{Number of particles} = \text{Moles} \times \text{Avogadro's number (L)}$$

$$= 0.0352\ \text{mol} \times 6.02 \times 10^{23}\ \text{ions mol}^{-1}$$

$$= 2.12 \times 10^{22}\ \text{ions}$$

Therefore, there are 2.12×10^{22} sodium ions (Na^+) in 2.5 g of Na_2SO_4.

3. Converting Mass to Moles

Worked Example 2.11

Calculate the number of moles present in 2.5 g of $Na_2SO_{4.}$

Solution:

- **Step 1**: List the known quantities and plan the problem.

Known:

mass of Na_2SO_4 produced = 2.81 g

Unknown:

amount of Na_2SO_4 produced in moles

One conversion factor will allow us to convert from mass to moles.

- **Step 2: Calculate**.

First, it is necessary to calculate the molar mass of Na_2SO_4. The molar mass is 142 g/mol.

$$n(Na_2SO_4) = \frac{m(Na_2SO_4)}{M(Na_2SO_4)}$$

$$= \frac{2.5\ g}{142\frac{g}{mol}} = 0.0202\ g$$

4. Converting Moles to Mass

Worked Example 2.12

Calculate the mass of 0.25 mol of Na_2SO_4.

Solution:

- Step 1: List the known quantities and plan the problem.

  Known:

  0.25 mol mol Na_2SO_4

  molar mass of Na_2SO_4 = 142.00 g/mol

  Unknown:

  0.25 mol of Na_2SO_4 = ? g

  The molar mass of Na_2SO_4 will allow us to convert from moles of Na_2SO_4 to grams.

- **Step 2: Calculate**.

  $$n(Na_2SO_4)=\frac{m(Na_2SO_4)}{M(Na_2SO_4)}$$

  Make $m(Na_2SO_4)$ the subject:

  $$\begin{aligned} m(Na_2SO_4) &= n(Na_2SO_4) \times M(Na_2SO_4) \\ &= 0.25 \text{ mol} \times 142.00 \tfrac{\text{g}}{\text{mol}} \\ &= 35.5 \text{ g} \end{aligned}$$

5. Converting moles to Number of Particles

Worked Example 2.13

Calculate the number of oxide ions contained in 0.5 mol of Al_2O_3.
[$L = 6.02 \times 10^{23}$ *particles* mol^{-1}]

Solution:

- **Step 1**: The number of oxide ions (O^{2-}) in the formula = 3
- **Step 2**: Calculate the total number of moles of oxide ions using the balanced equation below:

  $$Al_2O_3 \longrightarrow 3\ O^{2-} + 2\ Al^{3+}$$

1 mol of Al_2O_3 contains 3 mols O^{2-}

$$\therefore 0.5 \text{ mol of } Al_2O_3 = \frac{0.5 \text{ mol of } Al_2O_3}{\text{mol of } Al_2O_3} \times 3 \text{ mols } O^{2-} = 1.5 \text{ mol } O^{2-}$$

- Step 3: Convert moles of oxide ions to number of particles using Avogadro's number (L):

  Number of particles = moles × Avogadro's number (L)

  $$N(O^{2-}) = n(O^{2-}) \times L$$
  $$= 1.5 \text{ mol} \times 6.02 \times 10^{23} \text{ ions mol}^{-1}$$
  $$= 9.03 \times 10^{23} \text{ ions}$$

  Therefore, there are 9.03×10^{23} oxide ions in 0.5 mol of Al_2O_3.

6. Conversions between Moles and Gas Volume

a. Converting Gas Volume to Moles

Worked Example 2.14

Calculate the number of moles contained in 250 cm^3 of carbon dioxide gas at s.t.p.

$[V_m = 22.4 \text{ dm}^3 \text{mol}^{-1}]$

Solution:

- **Step 1:** List the known quantities and plan the problem.
- **Step 2:** Convert the volume from cm^3 to litres

  $$\text{volume of } CO_2 = 250 \text{ cm}^3 = \frac{250}{1000} = 0.250 \text{ dm}^3$$

  Unknown:

  moles of CO_2

  Use the molar volume to convert from dm^3 to moles:

  1 mol = 22.4 dm^3
- **Step 3:** Calculate.

  $$n(CO_2) = \frac{V(CO_2)}{V_m} = \frac{0.250 \text{ dm}^3}{22.4 \text{ dm}^3 \text{mol}^{-1}} = 0.0112 \text{ mol}$$

Worked Example 2.15

Calculate the volume occupied by 0.25 moles of carbon dioxide gas at s.t.p. [$V_m = 22.4\ dm^3 mol^{-1}$]

Solution:

Multiply the number of moles by the molar volume to get the volume:

Volume = Moles × Molar volume

$= 0.25\ mol \times 22.4\ dm^3 mol^{-1}1$

$= 5.6\ dm^3$

Therefore, 0.25 moles of carbon dioxide (CO_2) occupy a volume of 5.6 dm^3 at stp.

b. Converting volume to number of particles

Worked Example 2.16

Calculate the number of carbon dioxide (CO_2) molecules in 500 cm^3 at stp, [$L = 6.02 \times 10^{23}$ *particles mol*$^{-1}$; $V_m = 22.4\ dm^3 mol^{-1}$]

Solution:

- **Step 1**: Convert the volume from cm^3 to dm^3: 500 cm^3 = 0.5 dm^3
- **Step 2**: Calculate the number of moles:

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

  $$= \frac{0.5\ dm^3}{22.4\ dm^3 mol^{-1}} = 0.0223\ mol$$
- **Step 3**: Convert moles to number of particles (molecules):

  Number of particles = moles x Avogadro's number (L)

  $= 0.0223\ mol \times 6.02 \times 10^{23}$ molecules mol^{-1}

  $= 1.34 \times 10^{22}$ molecules

  Therefore, there are 1.34×10^{22} CO_2 molecules present in 500 cm^3 at stp.

Activity 2.9: Trial Questions

A chemical compound has a molar mass of 180 g/mol. Calculate:

1. The mass of 2 moles of the compound.
2. The number of molecules in this amount.
3. The volume occupied by these molecules at s.t.p.

THE MOLE CONCEPTS AND THEIR RELEVANCE IN PREPARING STANDARD SOLUTIONS

A solution is a uniform mixture of a solute and a solvent. The quantity of solute per unit volume of solution is termed as concentration. To be able to compare the concentration of solutions, we use standard units.

Types of Concentration

1. Quantity of substance concentration (Molarity)

It is defined as the quantity of solute (number of moles of solute) dissolved in one cubic decimetre of the solution.

Mathematically, it is expressed as:

$$\text{Concentration in mol dm}^{-3}\left(\text{C}\right)=\frac{\text{amount of substance in moles (n)}}{\text{Volume of solution in dm}^{3}\text{ (V)}}$$

The number of moles of solute n can be made the subject as

$$n = C \times V$$

NB:

a. This equation can be used to calculate the number of moles of solute required for a given volume of a specified concentration.

b. The volume of solution required can be calculated given the number of moles and a specified concentration.

c. If the volume of solution is given in cm^3, it must be converted to dm^3 by dividing it by 1000.

Worked Example 2.17

Calculate the number of moles in 250 cm^3 of 0.500 mol dm^{-3} sulphuric acid solution.

Solution:

Use the problem-solving approach

- Analyse the question

  Known:

  Volume, V = 250 cm^3

  Concentration, C = 0.500 mol dm^{-3}

  Formula to use: $n = C \times V$

  Unknown:

  Number of moles, n =?

- Solve:

  Convert the volume to dm^3 by dividing it by 1000.

  $V = \frac{250}{1000} = 0.250\ dm^3$

  By definition,

  $n = C \times V$

  Substituting the values,

  $n = 0.500\ mol\ dm^{-3} \times 0.250\ dm^3$

  $n = 0.125\ mol$

2. Mass Concentration (Concentration in g dm^{-3})

It is defined as the mass of solute dissolved in one cubic decimetre of the solution. It is denoted by ρ (**not to be confused with density**).

Mathematically, it is defined as:

Concentration in g $dm^{-3}(\rho) = \frac{mass\ of\ solute\ in\ grams\ (m)}{Volume\ of\ solution\ in\ dm^3\ (V)}$

The mass of the solute m can be made the subject as follows:

$m = \rho \times V$

NB: This equation can be used to calculate:

a. The mass of solute in a given volume of a solution of a specified concentration.

b. The volume of solution needed when the mass of solute and specified concentration is given.

Mass of solute required to prepare a given volume of a standard solution

Recall that,

$C = \frac{n}{V}$....... (1)

$but n = \frac{m}{M}$........ (2)

Substituting (2) into (1) gives $C = \frac{m}{M \times V}$

Hence $m = C \times M \times V$

This equation is used to calculate the mass of solute required to prepare a given volume of a standard solution.

Worked Example 2.18

Calculate the mass of NaOH required to prepare 250cm^3 of 0.50mol dm^{-3} sodium hydroxide solution.
[Na = 23, H = 1, O = 16]

Solution:

Use the problem-solving approach:

- Analyse the question

  Known

  Volume of solution, V = 250 cm^3

  Concentration, C = 0.50 mol dm^{-3}

  Relative atomic masses: Na = 23, H = 1, O = 16

  Formula to use: $m = C \times M \times V$

  Unknown:

  Mass, m =?

- Solve: Apply the formula

  Calculate the M (NaOH) = 23 + 16 + 1 = 40 g mol^{-1}

  Convert the volume to dm^3 by dividing it by 1000.

  $V = \frac{250}{1000} = 0.250 \text{ dm}^3$

  Substituting the values into the formula,

  $m = 0.50 \times 40 \times 0.25 = 5g$

  Therefore, the mass of NaOH required is 5g.

Relationship between molar concentration and mass concentration

Consider the equation:

$$C = \frac{m}{M \times V}$$

Also recall that, $\rho = \frac{m}{V}$. Combining the two equations yields, $C = \frac{\rho}{M}$

Where,

c = Concentration in mol dm^{-3}; ρ = mass concentration in g dm^{-3}

M = Molar mass in $gmol^{-1}$

Activity 2.10: Trial Questions

Calculate the required quantity for the following questions:

1. A solution of hydrochloric acid (HCl) has a concentration of 0.5 mol/dm^3. Calculate the number of moles of HCl present in 250 cm^3 (0.25 dm^3) of this solution.
2. A solution of sodium hydroxide (NaOH) has a concentration of 0.1 mol/dm^3. Determine the volume of this solution needed to contain 0.4 moles of NaOH.
3. Calculate the concentration of a solution that contains 2 moles of sulphuric acid (H_2SO_4) in a volume of 500 cm^3 (0.5 dm^3).

PREPARATION OF STANDARD SOLUTIONS

A standard solution is a solution whose concentration is accurately known.

Primary Standard

A primary standard is a substance that is usually available in pure form or a state of known purity, which is used in preparing a standard solution. Examples are sodium carbonate and potassium iodate

Properties of a primary standard

It should be available in pure form or easily purified.

It must be stable, that is, it must not lose weight or take up water during weighing.

It must have a reasonably high relative formula mass.

It must react speedily without side reactions with the substance being standardised.

It should have high solubility

How to prepare a standard solution from a solid solute

1. Determine the mass of the solute required to make the appropriate concentration and volume of desired solution.
2. Weigh accurately the solute in a beaker.
3. Add distilled water and its contents to the beaker and swirl to dissolve the solid.

NB: The beaker must have a lower volume than the standard volumetric flask being used.

4. Transfer the solution to the required standard volumetric flask through a funnel.
5. Rinse the stirrer and the beaker used into the flask, then add more distilled water until the meniscus lies on the calibration mark.
6. Invert the stoppered flask a few times to mix.
7. Label the solution.

Activity 2.11: Practise the example of preparing a standard solution from a solid solute

Task Example: Prepare 250 cm^3 of 2.0 mol dm^{-3} NaOH solution.

Materials needed

Beakers, volumetric flask, burette, electronic balance, stirring rods, funnels, and wash bottles, safety goggles and lab coats, stock solution, distilled water, and calculator.

Steps:

1. Perform calculations to determine the required quantities:
   - Calculate the amount of the solute needed to prepare the solution.

     Use the formula:

     $$c(\text{NaOH}) = \frac{n(\text{NaOH})}{V(\text{NaOH})}$$

- Convert volume from cm^3 to dm^3

  $250\ cm^3 = \frac{250}{1000} = 0.250\ dm^3$

  $n(NaOH) = c(NaOH) \times V(NaOH)$

  $= 2.0\ mol\ dm^{-3} \times 0.250\ dm^3$

  $= 0.50\ mol$

- Convert moles to grams

  Find the molar mass of NaOH

  $M(NaOH) = 23 + 16 + 1$

  $= 40\ g\ mol^{-1}$

  $m(NaOH) = n(NaOH) \times M(NaOH)$

  $= 0.50\ mol \times 40.0\ g\ mol^{-1}$

  $= 20.0\ g$

2. Accurately weigh the calculated mass of sodium hydroxide using the electronic balance.
3. Transfer the weighed solute to a beaker and add a small volume of distilled water.
4. Stir with a stirring rod until the solute is completely dissolved.
5. Transfer the small solution in the beaker into the right size volumetric flask (e.g. 250 cm^3) with the aid of a funnel.
6. Rinse the beaker with distilled water and add the rinsing to the volumetric flask to ensure all solute is transferred.
7. Cover the volumetric flask and invert it several times to ensure thorough mixing of the solution.
8. Label the volumetric flask with the concentration, the solute, the date.
9. Place the prepared standard solution in a designated place for later use.

Preparation of Standard Solution from Concentrated Solution

1. Use the dilution formula ($C_1V_1 = C_2V_2$) to calculate the volume of the concentrated solution required.
2. Pour some distilled water into the required standard volumetric flask.
3. Measure the stock or concentrated solution and transfer it into the distilled water in the volumetric flask.

4. Swirl the flask and its content and top the solution to the calibration mark with distilled water.
5. Label the solution.

Determination of the concentration of a stock solution

The commercial stock solution usually contains chemical assay (that is the label on their container, specifying the purity, density, molecular mass, and other relevant information).

1. Calculate the mass of the substance in 1 dm^3.
2. Calculate the mass of the pure substance in 1dm^3 by multiplying by the percentage purity.
3. Divide this mass by the molar mass to get the concentration.

   Mathematically use the formula:

$$Concentration\left(C\right)=\frac{\text{Density }(\rho)\times 1000\times\text{percentage purity }(\%)}{Molar\ mass\ (M)\times 100}$$

Activity 2.12: Preparing standard solutions of various concentrations

Materials needed

Beakers, volumetric flask, burette, electronic balance, stirring rods, funnels, and wash bottles, safety goggles and lab coats, stock solution, distilled water, and calculator.

Steps:

1. Observe safety precautions, such as wearing goggles and lab coats, handling chemicals carefully, and using apparatus properly.
2. Identify and label the set of apparatus required.
3. Prepare solutions of specific concentration such as 0.10 mol dm^{-3}, 1.0 mol dm^{-3}, 0.50 mol dm^{-3} of sodium hydroxide, sodium chloride and hydrochloric acid.

Activity 2.13 Preparing a standard solution from a concentrated solution

Materials needed

Beakers, volumetric flask, burette, electronic balance, stirring rods, funnels, and wash bottles, safety goggles and lab coats, stock solution, distilled water, and calculator.

Task Example: Prepare a 250 cm^3 solution of HCl with a concentration of 2.0 mol/dm^3 using a stock solution of known density and percentage purity.

Steps:

Prepare a 250 cm^3 solution of HCl of concentration 2.0 mol dm^{-3} from a stock HCl solution of specifications:

Density = 1.19 g cm^{-3}

Percentage purity = 37%

Molar mass of HCl = 36.5 g mol^{-1}

1. Calculate the concentration of the stock solution:

- Calculate the mass of HCl in 1 cm^3 of the stock solution:

  Mass of solution = density × volume

  $= 1.19 \text{ g cm}^{-3} \times 1 \text{ cm}^{-3} = 1.19 \text{ g}$

  Given that the solution is 37% HCl by mass:

  Percentage purity of HCl = 37%

  Mass of HCl = 1.19 g × 0.37 = 0.4403 g

  Now concentration of the stock solution:

  $$n(\text{HCl}) = \frac{m(\text{HCl})}{M(\text{HCl})} = \frac{0.4403 \text{ g}}{36.5 \text{ g mol}^{-1}} = 0.01207 \text{ mol}$$

  Since this is the amount in 1 cm^3, the concentration of the stock solution is:

  1 cm^3 of the stock solution contains 0.01207 mol HCl

$$\therefore 1000 \text{ cm}^3 = \frac{0.01207 \text{ mol}}{1 \text{ cm}^3} \times 1000 \text{ cm}^3 = 12.07 \text{ mol HCl}$$

Concentration of the stock solution is 12.07 mol dm^{-3} HCl

2. Calculate the volume of stock solution needed.

We want to prepare 250 cm^3 (0.250 dm^3) of a 2.0 mol/dm^3 HCl solution.

Using the dilution formula:

$C_1V_1 = C_2V_2$

$C_1 = 12.07 \text{ mol dm}^{-3}$

$V_1 =$ volume of stock solution needed

$C_2 = 2.0 \text{ mol dm}^{-3}$

$V_2 = 250 \text{ cm}^3$

$$V_1 = \frac{C_2V_2}{C_1} = \frac{2.0 \text{ mol dm}^{-3} \times 250 \text{ cm}^3}{12.07 \text{ mol dm}^{-3}} = 41.4 \text{ cm}^3$$

3. Measure 41.4 cm³ of the stock solution using a burette.

4. Transfer the stock solution to a 250 cm³ volumetric flask.

5. Add distilled water to the flask up to the 250 cm³ mark.

6. Invert the flask to mix the solution thoroughly.

REVIEW QUESTIONS

Review Question 2.1

The atomic mass unit of Carbon-12 is 1.6603×10^{-24} g. If the average mass of an atom **X** is 6.63310×10^{-23} g. Determine its relative atomic mass.

Review Questions 2.2

1. Calculate the relative molecular masses (M_r) of the following substances:

a. Formic Acid (CH_2O_2)
b. Acetic Acid ($C_2H_4O_2$)
c. Ethanol (C_2H_6O)
d. Acetone (C_3H_6O)
e. Citric Acid ($C_6H_8O_7$)
f. Sucrose ($C_{12}H_{22}O_{11}$)

For reference, the atomic masses are approximately: H = 1; O = 16; C = 12

2. Determine the relative molecular mass of the following:

a. NH_3
b. CH_4
c. $C_{16}H_{16}F_3NO$
d. S^{2-}
e. SO_4^{2-}
f. $Na_2CO_3 . 10\, H_2O$

A_r: $H = 1.0; C = 12.0; N = 14.0; O = 16.0; F = 19.0; Na = 23.0; S = 32.0$

Review Questions 2.3

Find the answers to the following questions:

1. How many moles of sodium chloride (NaCl) are present in 1.505×10^{24} formula units of NaCl?

2. Determine the number of moles in 2.409×10^{23} atoms of gold (Au).

3. How many moles are in 8.436×10^{24} molecules of glucose ($C_6H_{12}O_6$)?

4. Calculate the number of moles in 6.022×10^{21} molecules of nitrogen gas (N_2).

5. What is a mole and why is it important in chemistry?

6. Explain how the mole relates to Avogadro's number.

7. The number of molecules of ammonia gas is 12.04×10^{23}. Calculate the number of moles of ammonia gas. [$L = 6.02 \times 10^{23}$]
8. Calculate the number of oxygen molecules in 0.5 mol of oxygen gas. [$L = 6.02 \times 10^{23}$]
9. Calculate the number of atoms in 16 g of copper, Cu. [Cu = 63.5, $L = 6.02 \times 10^{23}$]

Review Questions 2.4

1. Determine the number of moles in 88 g of carbon dioxide (CO_2).
2. How many moles are in 48 g of methane (CH_4)?
3. Calculate the number of moles in 180 g of glucose ($C_6H_{12}O_6$).
4. Determine the number of moles in 58.5 g of sodium chloride (NaCl).
5. Calculate the mass of 3 moles of water (H_2O).
6. Determine the mass of 2 moles of carbon dioxide (CO_2).
7. Find the mass of 4 moles of methane (CH_4).
8. Calculate the mass of 0.5 moles of glucose ($C_6H_{12}O_6$).
9. Determine the mass of 1.5 moles of sodium chloride (NaCl).

Review Questions 2.5

Calculate the required values for the following questions

1. Determine the volume occupied by 0.5 moles of carbon dioxide (CO_2) at STP, in dm^3.
2. How many moles of hydrogen gas (H_2) are there in a 67.2 dm^3 container at STP?
3. Calculate the volume occupied by 2 moles of nitrogen gas (N_2) at STP, in dm^3.

Review Questions 2.6

1. A sample of carbon dioxide (CO_2) has a molar mass of 44 g/mol. Calculate:
   a. The mass of 1.5 moles of carbon dioxide.
   b. The number of molecules in this amount.

c. The volume occupied by these molecules at STP.

2. A sample of propane (C_3H_8) has a molar mass of 44 g/mol. Calculate:
   a. The mass of 3 moles of propane.
   b. The number of molecules in this amount.
   c. The volume occupied by these molecules at STP.
3. The molar mass of CO_2 is 44 g/mol. How many moles of CO_2 are present in 124 g sample of CO_2.
4. What is the mass of 5.0×10^{23} molecules of NO_2?

   $[N = 14.0; O = 16.0 \mathrm{L} = 6.02 \times 10^{23}]$
5. a. How many molecules are there in 4.00 mol of glucose, $C_5H_{12}O_6$?
   b. How many atoms of carbon?
   c. How many atoms of hydrogen?

   $[\mathrm{L} = 6.02 \times 10^{23}]$

Review Questions 2.7

1. A solution of potassium permanganate ($KMnO_4$) has a concentration of 0.02 mol/dm³. Find the volume of this solution required to contain 0.1 moles of $KMnO_4$.
2. Determine the volume of a 0.5 mol/dm³ solution of glucose ($C_6H_{12}O_6$) needed to obtain 0.15 moles of glucose.
3. Calculate the mass of sodium chloride (NaCl) dissolved in 500 cm³ (0.5 dm³) of a solution with a concentration of 0.4 g/dm³.
4. Find the concentration of a solution if 30 g of potassium nitrate (KNO_3) is dissolved in 150 cm³ (0.15 dm³) of water.
5. Determine the volume of a solution with a mass of 25 g and a concentration of 0.1 g/dm³.
6. Calculate the mass of copper sulphate ($CuSO_4$) in 250 cm³ (0.25 dm³) of solution with a concentration of 0.8 g/dm³.
7. Find the concentration of a solution if 50 g of sucrose ($C_{12}H_{22}O_{11}$) is dissolved in 500 cm³ (0.5 dm³) of water.
8. A solution of hydrochloric acid (HCl) has a concentration of 0.4 mol/dm³. Calculate the mass of HCl in 300 cm³ (0.3 dm³) of this solution.

9. Calculate the volume of a 0.2 mol/dm^3 solution of sulphuric acid (H_2SO_4) needed to obtain 50 g of H_2SO_4.

10. Find the concentration of a solution if 150 g of sodium hydroxide (NaOH) is dissolved in enough water to make 500 cm^3 (0.5 dm^3) of solution.

11. Calculate the mass of a sample of copper sulphate ($CuSO_4$) in 200 cm^3 (0.2 dm^3) of solution with a concentration of 1.5 g/dm^3.

12. A solution of ethanol (C_2H_6O) has a concentration of 0.8 mol/dm^3. Find the volume of this solution required to obtain 100 g of ethanol.

13. List five (5) apparatus used in preparing a standard solution from solid solutes.

14. Work in small groups to prepare 250 cm^3 of 2.0 mol dm^{-3} solution of sodium hydroxide in the laboratory. [H = 1.0, O = 16.0, Na = 23.0]

15. What are some potential sources of error when preparing standard solutions, and how can they be minimised?

SUGGESTED ANSWERS TO REVIEW QUESTIONS

Answer to review question 2.1

$A_r(\mathbf{X}) = 39.95$

Answers to Review questions 2.2

1. a. Formic Acid (CH_2O_2) = 46
   b. Acetic Acid ($C_2H_4O_2$) = 60
   c. Ethanol (C_2H_6O) = 46
   d. Acetone (C_3H_6O) = 58
   e. Citric Acid ($C_6H_8O_7$) = 192
   f. Sucrose ($C_{12}H_{22}O_{11}$) = 342
2. a. $M_r(NH_3) = 17$;
   b. $M_r(CH_4) = 16$;
   c. $M_r(C_{16}H_{16}F_3NO) = 295.0$
   d. $M_r(S^{2-}) = 32.0$;
   e. $M_r(SO_4^{2-}) = 96.0$;
   f. $M(Na_2CO_3 . 10\,H_2O) = 286$

Answers to Review Questions 2.3

1. 2.500 mol.
2. 0.400 mol.
3. 14.00 mol.
4. 0.0100 mol.
5. The mole is a fundamental concept in chemistry that enables chemists to count particles by weighing them and to relate macroscopic quantities to atomic-scale quantities. It is essential for understanding chemical reactions,

preparing solutions, and applying gas laws, making it a cornerstone of quantitative chemistry.

6. 1 mole of every substance contains 6.02×10^{23} elementary entities. This relationship forms the basis of quantitative chemistry, allowing chemists to bridge the gap between the microscopic and macroscopic worlds and enabling precise measurements and calculations in chemical reactions and analyses.
7. $n(NH_3) = 2\ mol$
8. $n(O) = 3.01 \times 10^{23}$ molecules
9. $n(Cu) = 0.0913$ atoms

Answers to Review Questions 2.4

1. 2 mol.
2. 3 mol.
3. 1 mol.
4. 1 mol.
5. 54 g
6. 88 g
7. 64 g
8. 90 g
9. 87.75 g

Answers to Review questions 2.5

1. 11.2 dm^3
2. 3 moles
3. 44.8 dm^3

Answers to Review questions 2.6

1. a. Mass of 1.5 moles of carbon dioxide = 66 g
   b. Number of molecules in this amount = 9.033×10^{23} molecule
   c. Volume occupied by these molecules at STP = 33.6 dm^3

2. a. Mass of 3 moles of propane = 132 g
   b. Number of molecules in this amount = 1.8066×10^{24} molecule
   c. Volume occupied by these molecules at STP = 67.2 dm^3
3. 2.82 mol
4. 38.2 g
5. (a) 2.41×10^{23} molecules ;
   (b) 1.20×10^{25} molecules ;
   (c) 289×10^{25} molecules

Answers to Review Questions 2.7

1. 5 dm^3
2. 0.3 dm^3
3. 0.2 g
4. 0.2 g/dm^3
5. 250 dm^3
6. 200 g
7. 0.1 g/dm^3
8. 0.12 g
9. 0.25 dm^3
10. 0.3 g/dm^3
11. 0.3 g
12. 125 dm^3
15. Inaccurate weighing, incorrect volume, impure solute, equipment errors, inaccurate pipettes, burettes, or volumetric flasks, not measuring volumes on a flat surface at eye level with the meniscus and contamination calculation mistakes.

EXTENDED READING

Click on the link below and read the recommended books for further information on topics discussed in this section.

https://www.khanacademy.org/science/ap-chemistry-beta/x2eef969c74e0d802:atomicstructure-and-properties

Chang Reymond. (2017). General Chemistry – Essential Concepts. The McGraw

Hill Companies

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