C H E M I S T R Y

MOLE CONCEPT

WHY THE CONCEPT OF MOLE WAS INVENTED?

NOTES

What you will learn

- Order of magnitude
- Why do we study the mole concept
- What is a mole
- Mole - particle conversion

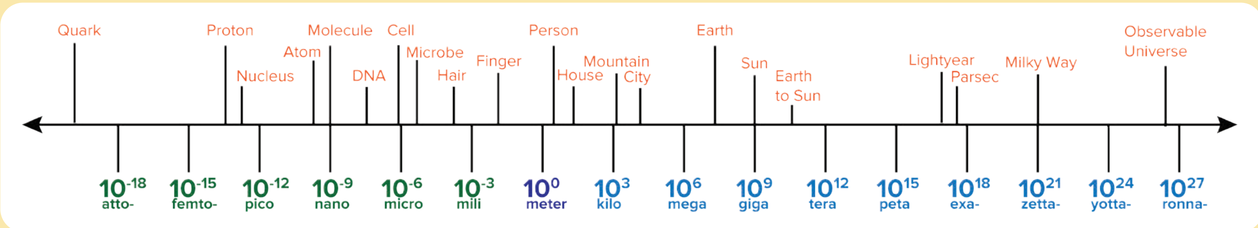
BOARDS

Order of Magnitude

When a number is written in its scientific notation, the exponent to which 10 is raised gives the order of magnitude.

Examples:

- The order of magnitude of 12 is 1 because it is around 10 in size.
- A kilometre (1000 meters) is three orders of magnitude greater than a metre.



Why do we study the Mole Concept?

- To convert the theoretical particle information going on at the molecular level into physically observable information.

  For example, in the formation of ammonia from N_2 and H_2

  $$N_2 + 3H_2 \rightarrow 2NH_3$$

  We cannot take 1 molecule of N_2 and 3 molecules of H_2 to proceed with this reaction in the laboratory because their masses are not physically observable.

- So, the study of the mole concept helps us to connect the microworld to the physically observable macroworld.

  For example, physical quantities like weight, volume, density, pressure are not measurable at particle level but can be measured after the study of the mole concept.

What is a Mole?

Mole is an unit to measure a relatively large number of entities such that 1 mole of an item contains 6.022×10^{23} entities.

Avogadro's number

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Avogadro's number is another number used to represent quantities. Its value is same as the number of entities of an item present in 1 mole of that item, i.e., 6.022×10^{23}. This number is represented by N_A.

1 kilo of chocolates = 10^3 chocolates

N_A chocolates = 6.022×10^{23} chocolates

- Avogadro's number does not have unit.
- Mole is just a number and does not represent any physical quantity.
- The value of $N_A = 6.022 \times 10^{23}$ is an approximate value.

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Mole - Particle Conversion

Dozen - item analogy

Number of dozens of the item

$$= \frac{\text{number of items}}{12}$$

- 24 chocolates = $\frac{24}{12}$ = 2 dozens of chocolates
- 6 marbles = $\frac{6}{12} = \frac{1}{2}$ dozens of marbles
- Number of moles = $\frac{\text{Number of items}}{N_A}$

Number of items = number of dozens of item × 12

- 2 dozens of chocolates = 2 x 12 = 24 chocolates
- $\frac{1}{2}$ dozens of marbles = $\frac{1}{2} \times 12 = 6$ marbles
- Number of items = Number of moles × N_A

Conversion: Mole to particle

Find **the number of CO_2 molecules in 2 moles of** a pure sample **of CO_2 molecules.**

Solution

Given,

Number of moles of CO_2 molecules = 2 *mol*

Number of particles = Number of moles × $N_A = 2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$

Conversion: Mole to particle

Find the **number of He atoms** in $\frac{8}{3}$ **moles** of a pure sample **of He atoms.**

Solution

Given,

Number of moles of He atom = $\frac{8}{3}$ *mol*

Number of particles = Number of moles × N_A = $\frac{8}{3}$ × 6.022 × 10^{23} = 1.606 × 10^{24}

Conversion: Mole to particle

Find the **number of H_2SO_4 molecules** in **10^{-13} moles of a pure H_2SO_4 sample.**

Solution

Given,

Number of moles of H_2SO_4 = 10^{-13} *mol*

Number of particles = Number of moles × N_A = 10^{-13} × 6.022 × 10^{23} = 6.022 × 10^{10}

Conversion: Mole to particle

Find the **number of $Na_2SO_4.5H_2O$ formula units** in **5 kilomoles of a pure $Na_2SO_4.5H_2O$ sample.**

Solution

Step 1:

Convert kilomoles to moles.

1 kilomole = 10^3 *mol*

Number of moles of $Na_2SO_4.5H_2O$

= 5 × 1000 = 5 × 10^3 *mol*

Step 2:

Find the number of particles from the given amount of moles.

Number of particles = Number of moles × N_A

= 5 × 10^3 × 6.022 × 10^{23} = 3.011 × 10^{27}

Conversion: Mole to particle

Find **the number of moles of CO_2 molecules** in a sample containing **3.011 × 10^{23} CO_2 molecules.**

Solution

Given,

Number of molecules of CO_2 = 3.011×10^{23}

Number of moles CO_2 = $\dfrac{\text{Number of particles}}{N_A} = \dfrac{3.011 \times 10^{23}}{6.022 \times 10^{23}} = 0.5\ mol$

Conversion: Mole to particle

Find the **number of moles of He atoms** in a sample containing **3.011 × 10^{40} He atoms.**

Solution

Given,

Number of atoms of He = 3.011×10^{40}

Number of moles of He = $\dfrac{\text{Number of particles}}{N_A} = \dfrac{3.011 \times 10^{40}}{6.022 \times 10^{23}} = 5 \times 10^{16}\ mol$

Conversion: Mole to particle

Find the **number of moles of H_2SO_4 molecules** in a sample **containing 10^{10} H_2SO_4 molecules.**

Solution

Given,

Number of molecules of H_2SO_4 = 10^{10}

Number of moles of H_2SO_4 = $\dfrac{\text{Number of particles}}{N_A} = \dfrac{10^{10}}{6.022 \times 10^{23}} = 1.661 \times 10^{-14}\ mol$

Conversion: Mole to particle

Find the number of **moles of H_2SO_4 molecules** in a sample containing 10 **H_2SO_4 molecules.**

Solution

Given,

Number of molecules of H_2SO_4 = 10

Number of moles H_2SO_4 = $\dfrac{\text{Number of particles}}{N_A} = \dfrac{10}{6.022 \times 10^{23}} = 1.661 \times 10^{-23}\ mol$

MOLE CONCEPT

PLAYING WITH THE BASICS

What you already now

- Order of magnitude
- Mole-particle conversions
- Avogadro's number

What you will learn

Conversion: Number of molecules/formula units to number of particles

Conversion: Number of particles to number of molecules/formula units

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Car-Tyre Analogy

- We know that a car has 4 tyres. Let us assume that 1 formula unit of H_2SO_4 is the car and each oxygen atom is a tyre.
- We can say that each car will have 4 tyres or that 2 cars will have 8 tyres. By a similar analogy, we can say that each H_2SO_4 unit has 4 O atoms. Hence, n units of H_2SO_4 will have 4n O atoms.

1 car	4 tyres
22 cars	88 tyres
2 dozen cars	96 tyres

Similarly,

1 molecule of H_2SO_4	4 oxygen atoms
5 molecules of H_2SO_4	20 oxygen atoms
2 moles of H_2SO_4	8 N_A oxygen atoms

20 tyres	5 cars
2 dozen tyres	6 cars

Similarly,

20 oxygen atoms	5 molecules of H_2SO_4
4 N_A oxygen atoms	1 *mol* of H_2SO_4

BOARDS MAIN

Interconversion: Particles Within a Particle

Case 1:

Calculating the number of particular entities from the given number of molecules or formula units of a chemical species.

Step 1:

Calculate the number of the desired entities in 1 molecule or 1 formula unit of the chemical species.

Step 2:

Multiply the number of desired entities present in 1 molecule or 1 formula unit of the chemical species by the total number of the chemical species.

Conversion: Particle within a particle

a. Find the number of O atoms in 5 H_2SO_4 molecules.

b. Find the number of O atoms in 5 $Na_2SO_4.5H_2O$ formula units.

Solution

a. Number of O atoms in 1 molecule of H_2SO_4 = 4
   Number of O atoms in 5 molecules of H_2SO_4 = 5 × 4 = 20

b. Number of O atoms in 1 formula unit of $Na_2SO_4.5H_2O$ = 4 + 5 × 1 = 9
   Number of O atoms in 5 formula units of $Na_2SO_4.5H_2O$ = 9 × 5 = 45

Conversion: Particle within a particle

a. Find the number of O atoms in 5×10^5 CO_2 molecules.

b. Find the number of O atoms in 5×10^{13} $Mg(NO_3)_2$ formula units.

Solution

a. Number of O atoms in 1 molecule of CO_2 = 2

Number of O atoms in 5×10^5 molecule of $CO_2 = 2 \times 5 \times 10^5 = 10^6$

b. Number of O atoms in 1 formula unit of $Mg(NO_3)_2 = 3 \times 2 = 6$

Number of O atoms in 5×10^{13} formula units of $Mg(NO_3)_2 = 6 \times 5 \times 10^{13} = 3 \times 10^{14}$

Case 2:

Calculating the number of molecules or formula units of a chemical species from the given number of entities.

Step 1:

Calculate the number of the desired entities in 1 molecule or 1 formula unit of the chemical species.

Step 2:

Divide the total number of entities by the number of entities present in 1 molecule or 1 formula unit of the chemical species.

Conversion: Particle within a particle

If we have a pure CO_2 sample and it contains a total of 10^{10} O atoms, find the number of molecules of CO_2 present in the sample.

Solution

Step 1:

Finding the number of O atoms present in CO_2 molecules.

From the chemical formula,

1 formula unit of CO_2 contains 2 O atoms

Hence, two atoms of oxygen correspond to one molecule of CO_2

Step 2:

Finding the number of CO_2 molecules.

Given number of O atoms = 10^{10}

Hence, from step 1,

Number of CO_2 molecules that contain 10^{10} O atoms = $\frac{10^{10}}{2} = 5 \times 10^9$

Conversion: Particle within a particle

If we have a pure CO_2 sample and it contains a total of 10^{10} O atoms, find the number of molecules of CO_2 present in the sample.

Solution

Step 1:

Finding the number of CO_2 molecules.

Number of O atoms in 1 molecule of CO_2 = 2

Number of CO_2 molecules from 10^{10} O

O atoms = $\frac{10^{10}}{2}$ = 5×10^9

Step 2:

Finding the number of moles of CO_2.

1 *mol* of CO_2 molecules consists of 6.022×10^{23} molecules.

Moles of CO_2 in sample

= $\frac{5 \times 10^9}{6.022 \times 10^{23}}$ = 8.303×10^{-15} *mol*

Conversion: Particle within a particle

Find the number of O atoms in 6.023×10^{23} $Na_2SO_4.5H_2O$ formula units.

Solution

Number of O atoms in 1 formula unit of $Na_2SO_4.5H_2O$ = 4 + (5 × 1) = 9

Number of O atoms in 6.022×10^{23} formula units of $Na_2SO_4.5H_2O$ = $9 \times 6.022 \times 10^{23} = 5.420 \times 10^{24}$

Conversion: Particle within a particle

Find the number of O atoms in 3 moles of CO_2 molecules.

Solution

Number of O atoms in 1 molecule of CO_2 = 2

Number of molecules in 1 mole of CO_2 = 6.022×10^{23}

Number of O atoms in 3 moles of CO_2 = $2 \times 3 \times 6.022 \times 10^{23} = 3.613 \times 10^{24}$

Conversion: Particle within a particle

Find the number of O atoms in 0.0003 moles of $Na_2SO_4.5H_2O$ formula units.

Solution

Number of O atoms in 1 molecule of $Na_2SO_4.5H_2O$ = 9

Number of formula units in 1 mole of $Na_2SO_4.5H_2O$ = 6.022×10^{23}

Number of O atoms in 0.0003 moles of CO_2 = $9 \times 0.0003 \times 6.022 \times 10^{23} = 1.626 \times 10^{21}$

Conversion: Particle within a particle

Find the number of O atoms in $\frac{1}{3}$ moles of $Mg(NO_3)_2$ formula units.

Solution

1 formula unit of the compound contains 6 O atoms
1 mole of the compound contains $6 \times 6.022 \times 10^{23}$ O atoms
Number of O atoms in $\frac{1}{3}$ mole of compound = $\frac{1}{3} \times 6 \times 6.022 \times 10^{23} = 1.2044 \times 10^{24}$

Conversion: Particle within a particle

Find the number of moles of $Na_2SO_4.5H_2O$ formula units in a pure sample that contains 1.8 x 10^{-3} moles of O atoms.

Solution

9 atoms of O are present in 1 formula unit of the compound
9 moles of O atoms are present in 1 mole of the formula unit
1.8×10^{-3} moles O atoms will be present in $\frac{1 \times 1.8 \times 10^{-3}}{9} = 2 \times 10^{-4}$ *mol* formula units

Conversion: Particle within a particle

Find the number of electrons in 6.023×10^{19} H_2O molecules.

Solution

1 molecule of H_2O contains $(2 \times 1 + 8)$ e^- = 10 electrons
6.023×10^{19} molecules will contain $10 \times 6.023 \times 10^{19} = 6.023 \times 10^{20}$ electrons

Conversion: Particle within a particle

If we have a pure CO_2 sample and it contains a total of 440 electrons, find the number of molecules of CO_2 present in the sample.

Solution

1 molecule of the compound contains $(6 + (2 \times 8)) = 22$ electrons

22 electrons are present in 1 molecule of the compound.

440 electrons will be present in $\frac{1}{22} \times 440 = 20$ molecules.

Conversion: Particle within a particle

If we have a pure CO_2 sample and it contains a total of 2.2×10^{10} electrons, find the number of molecules of CO_2 present in the sample.

Solution

22 electrons are present in 1 molecule of the compound

2.2×10^{10} electrons will be present in $\frac{1}{22} \times 2.2 \times 10^{10} = 10^9$ molecules

Conversion: Particle within a particle

If we have a pure $Na_2SO_4 . 5H_2O$ sample and it contains a total of 2.4×10^{10} electrons, find the number of formula units of $Na_2SO_4.5H_2O$ present in the sample.

Solution

1 formula unit of compound contains $(2 \times 11) + (1 \times 16) + (4 \times 8) + (5 \times 10) = 120$ electrons
120 electrons are present in 1 formula unit of the compound.

2.4×10^{10} electrons will be present in $\frac{1}{120} \times 2.4 \times 10^{10}$ molecules $= 2 \times 10^8$ molecules

Conversion: Particle within a particle

Find the number of electrons in 5 moles of H_2SO_4 molecules.

Solution

1 molecule of H_2SO_4 has $(2 + 16 + 4(8)) = 50$ electrons.

So, 1 mole of H_2SO_4 will have 50 N_A electrons.

Similarly, using the unitary method, the number of electrons in 5 moles of H_2SO_4
$= 250\ N_A = 250 \times 6.022 \times 10^{23} = 1.505 \times 10^{26}$ electrons.

Conversion: Particle within a particle

Find the number of electrons in $\frac{5}{8}$ moles of H_2SO_4 molecules.

Solution

1 molecule of H_2SO_4 has (2 + 16 + 4(8)) = 50 electrons.

So, 1 mole of H_2SO_4 will have 50 N_A electrons.

Similarly, using the unitary method, the number of electrons in $\frac{5}{8}$ moles of H_2SO_4
$= \frac{5}{8} \times 50 \times 6.022 \times 10^{23} = 1.881 \times 10^{25}$ electrons.

Conversion: Particle within a particle

Find the number of electrons in $\frac{5}{8}$ moles of SO_4^{2-} ions.

Solution

1 ion of SO_4^{2-} has (16 + 4(8) + 2) = 50 electrons
So 1 mole of SO_4^{2-} will have $50N_A$ electrons.
Similarly, using the unitary method, the number of electrons in $\frac{5}{8}$ moles of SO_4^{2-}
$= \frac{5}{8} \times 50 \times 6.022 \times 10^{23} = 1.881 \times 10^{25}$ electrons.

Conversion: Particle within a particle

If we have a pure $Mg(NO_3)_2$ sample and it contains a total of 7.4×10^{30} electrons, find the number of moles of $Mg(NO_3)_2$ present in the sample.

Solution

Step 1:
Finding the total number of molecules of $Mg(NO_3)_2$
1 molecule of $Mg(NO_3)_2$ has (12 + (31) × 2) = 74 electrons

$$\text{Number of molecules of } Mg(NO_3)_2 \text{ is} = \frac{\textit{Total number of electrons}}{\textit{Number of electons in one molecule}}$$

$$= \frac{7.4 \times 10^{30}}{74} = 10^{29} \text{ molecules}$$

Step 2:

Conversion: Particles to mole

Number of molecules of $Mg(NO_3)_2 = \dfrac{10^{29}}{6.022 \times 10^{23}} = 1.661 \times 10^5\ mol$

Conversion: Particle within a particle

If we have a pure $Mg(NO_3)_2$ sample and it contains a total of 37 moles of electrons, find the number of moles of $Mg(NO_3)_2$ present in the sample.

Solution

1 molecule of $Mg(NO_3)_2$ has $(12 + (31) \times 2) = 74$ electrons

37 moles of $Mg(NO_3)_2$ will contain $= 37 \times 74 \times 6.022 \times 10^{23} = 1.649 \times 10^{27}$ electrons

NOTES

C H E M I S T R Y

MOLE CONCEPT

CONNECTING THE MACRO AND MICRO WORLDS

What you already know

- Moles
- Conversion: Mole-particle

What you will learn

- Conversions: Particle within particles using moles
- Atomic mass
- Units of atomic mass
- Molar mass
- Average atomic mass
- Molecular mass
- g-atomic mass, g-molecular mass, and g-ionic mass
- Conversion: Mass-mole, particle within particles using moles
- Atomic mass

BOARDS

Conversions: Particles within a particle

- When you know the number of molecules of a compound, you can find the number of atoms present in it or vice-versa.
- In the same way, you can find the total number of electrons, protons or neutrons when the number of molecules or atoms is given or vice versa.
- Example: Let's say a sample has 10 H_2O molecules. There would be 20 H atoms in the sample. 100 electrons would be present in the same sample.
- Using these relations, we can find the number of subatomic particles, molecules, or atoms in a given sample of compound.

Conversion: Number of atoms to number of electrons

If we have a **pure $Mg(NO_3)_2$** sample and it contains a **total of 30 O atoms**, find the **total number of electrons** present in the sample.

Solution

Step 1:

Find the number of electrons in one $Mg(NO_3)_2$ unit.

Total number of electrons in $Mg(NO_3)_2$

$= 12 + (7 + 3 \times 8) \times 2 = 74$

Step 2:

Find the number of $Mg(NO_3)_2$ units.

Given, total number of O atoms = 30

Number of O atoms for each $Mg(NO_3)_2$ unit = 6

So, the number of $Mg(NO_3)_2$ units

$$= \frac{\text{Total O atoms}}{6} = \frac{30}{6} = 5$$

Step 3:

Multiply the number of electrons in one unit with the number of units to get the total number of electrons.

Total number of electrons = $74 \times 5 = 370$

Conversion: Number of atoms to number of electrons

If we have a **pure $Mg(NO_3)_2$ sample** and it contains a **total of 3.011×10^{24} O atoms**, find the **total number of electrons** present in the sample.

Solution

Step 1:

From the previous question, we already know that,

Total number of electrons in $Mg(NO_3)_2$

$= 12 + (7 + 3 \times 8) \times 2 = 74$

number of O atoms for each $Mg(NO_3)_2$ unit = 6

Step 2:

Find the number of $Mg(NO_3)_2$ units.

Given, total number of O atoms = 3.011×10^{24}

So, number of $Mg(NO_3)_2$ units

$$= \frac{\text{Total O atoms}}{6}$$

$$= \frac{3.011 \times 10^{24}}{6}$$

Step 3:

Multiply the number of electrons in one unit with the number of units to get the total number of electrons.

Total number of electrons = $74 \times 3.011 \times \frac{10^{24}}{6} = 3.7 \times 10^{25}$

BOARDS

Conversions Using Moles

Mole is nothing but a huge (6.023×10^{23}) number. Why this number, we will see further in the chapter. When amounts of particles are given in moles, the number of particles is nothing but the number of moles multiplied by N_A.

Conversion: Moles of electrons to number of atoms

If we have a **pure $Mg(NO_3)_2$ sample** and it contains **a total of 3.7 moles of electrons**, find the total number of O atoms present in the sample.

Solution

Step 1:

Find the number of $Mg(NO_3)_2$ units.

Given, total number of electrons = 3.7 N_A
Number of electrons in one formula unit of $Mg(NO_3)_2$ = 74

So, number of $Mg(NO_3)_2$ units

$= 3.7 \times \frac{N_A}{74}$

Step 2:

Multiply the number of $Mg(NO_3)_2$ units with O atoms in each $Mg(NO_3)_2$ unit.

Number of O atoms for each $Mg(NO_3)_2$ unit = 6

Total number of O atoms

$= (3.7 \times \frac{N_A}{74}) \times 6 = 0.3\ N_A$

Conversion: Moles of atoms to moles of electrons

If we have a **pure $Mg(NO_3)_2$** sample and it contains a **total of 3.6 moles of O atoms**, find the **total number of electrons** present in the sample.

Solution

Step 1:

Find the number of particles .

Total number of O atoms = 3.6 *mol*

$= 3.6 \times N_A$

Step 2:

Find the number of $Mg(NO_3)_2$ units.

Number of O atoms per $Mg(NO_3)_2$ unit = 6

So, number of $Mg(NO_3)_2$ units

$= 3.6 \times \frac{N_A}{6}$

Step 3:

Multiply the number of electrons in one unit with the number of units to get the total number of electrons.

Number of electrons per $Mg(NO_3)_2$ unit = 74

Total number of electrons $= 74 \times 3.6 \times \frac{N_A}{6}$

$= 44.4\ N_A$

The conversion between moles of particles within particles is similar to the conversion between numbers of particles within particles. The only difference is that the conversion between moles has its units as moles only.

Conversion: Moles of atoms to moles of electrons

If we have a **pure $Mg(NO_3)_2$ sample** and it contains a total of **3.6 moles of O atoms**, Find the total **number of moles of electrons** present in the sample.

Solution

Step 1:

Find the number of moles of $Mg(NO_3)_2$ units.

Total number of electrons = 3.6 *mol*

Number of O atoms per one $Mg(NO_3)_2$ unit = 6

Number of moles of

$Mg(NO_3)_2$ units = $\frac{(3.6)}{6}$ *mol*

Step 2:

Multiply the number of electrons in one unit with the number of moles of units to get the total number of moles of electrons.

Number of electrons per $Mg(NO_3)_2$ = 74

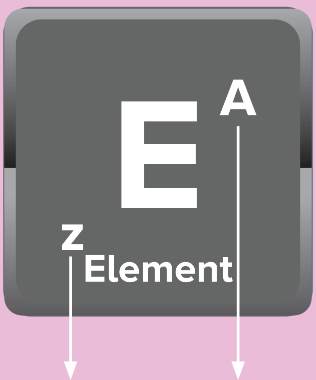
Total number of moles of electrons

$= \frac{(3.6)}{6} \times 74$

= 44.4 *mol*

BOARDS MAIN

Atomic Mass



Atomic number - number of electrons or protons

Mass number - Number of nucleons present or sum of number of neutrons and protons in an atom

Mass of an atom
Equal to summation of mass of **subatomic particles**

Neutron n^0

Electron e^-

Proton P^+

Mass (n^0)	:	Mass (P^+)	:	Mass (e^-)
$1.68 \times 10^{-24}\,g$	:	$1.67 \times 10^{-24}\,g$	:	$9.1 \times 10^{-28}\,g$
1	:	0.994	:	$\frac{1}{1837}$

- Generally, mass of electrons is considered to be negligible since the mass of protons/ neutrons is approximately 1837 times larger than the electrons.
- Mass of protons and neutrons is considered to be equal.

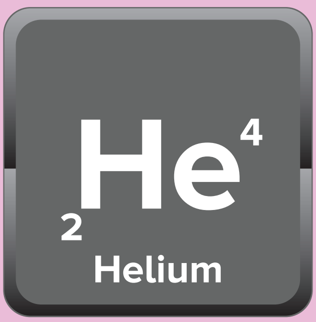
Atomic Mass = mass of neutrons (m_n) + mass of protons (m_p) + mass of electrons (m_e)

$= (n_n \times m_n) + (n_p \times m_p) + (n_e \times m_e)$

$= (n_n \times m_n) + (n_p \times m_p)$ {since $m_n, m_p >>>> m_e$}

$= (n_n + n_p) \times m_p = (A) \times (1.66 \times 10^{-24}\,g)$ {since $m_n = m_p$} and A (mass number) = $n_p + n_n$

Finding atomic mass



What is the mass of one atom of He in grams?

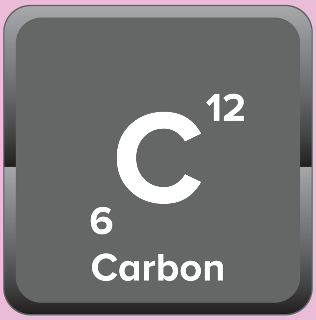
Solution

Number of nucleons in a He atom (A) = 4

Mass of each nucleon = $1.66 \times 10^{-24}\,g$

Mass of He atoms = $4 \times 1.66 \times 10^{-24} = 6.64 \times 10^{-24}\,g$

Finding atomic mass



Calculate the mass of a carbon atom in grams.

Solution

Number of nucleons in a C atom = 12

Mass of each nucleon = 1.66×10^{-24} *g*

Mass of C atoms = $12 \times 1.66 \times 10^{-24} = 20 \times 10^{-24}$ *g*

Finding atomic mass

What is the atomic mass of Mg in grams?

Solution

Number of nucleons in a Mg atom = 24

Mass of each nucleon = 1.66×10^{-24} *g*

Mass of Mg atoms = $24 \times 1.66 \times 10^{-24} = 40 \times 10^{-24}$ *g*

Units of Atomic Mass

- Atomic Mass = (Mass Number) × (1.66×10^{-24}) *g*
  = (A) × (1 *amu*)
  = (A) × (1 *u*)
  = (A) × (1 *Da*)
- The quantity 1/12 × (mass of an atom of C–12) is known as an atomic mass unit.
- 1 *amu* = 1 Dalton (Da) = 1 *u*
- The actual mass of one atom of C-12 = 1.9924×10^{-26} kg

Finding atomic mass in amu

Calculate the atomic mass of the following atoms in *amu*.

a. $^{4}_{2}\mathrm{He}$ Helium

b. $^{12}_{6}\mathrm{C}$ Carbon

c. $^{23}_{11}\mathrm{Na}$ Sodium

d. $^{40}_{20}\mathrm{Ca}$ Calcium

e. $^{27}_{13}\mathrm{Al}$ Aluminium

Solution

a) Atomic mass of He = 4 *amu*

b) Atomic mass of C = 12 *amu*

c) Atomic mass of Na = 23 *amu*

d) Atomic mass of Ca = 40 *amu*

e) Atomic mass of Al = 27 *amu*

Finding molar mass of an atom

- Molar mass of an atom is nothing but the mass of 6.022×10^{23} atoms.
- $1\ \text{amu} \times N_A = (1.66 \times 10^{-24})\ \text{g} \times 6.022 \times 10^{23} \cong 1\ g$
- So, the molar mass of an atom in grams is numerically equal to the mass number (A).

Find the molar mass of oxygen.

$^{16}_{8}\mathrm{O}$ Oxygen

Solution

Number of nucleons in an O atom = 16

Mass of each nucleon = $1.66 \times 10^{-24}\ g$

Molar mass of O

$= 16 \times 1.66 \times 10^{-24} \times 6.022 \times 10^{23}$

$= 16\ g$

Easier approach:

Molar mass in grams is numerically equal to the mass number.

Hence, the molar mass of O atoms = 16 *g*

BOARDS MAIN

Average Atomic Mass

Isotopes

Isotopes are those particles that have the same number of protons but different number of neutrons.

Examples:

$^{1}_{1}H$ Hydrogen	$^{2}_{1}H$ Hydrogen	$^{3}_{1}H$ Hydrogen
Protium	**Deuterium**	**Tritium**

$^{16}_{8}O$ Oxygen	$^{18}_{8}O$ Oxygen
O-16 Isotope	**O-18 Isotope**

- When % abundance and mass of each isotope is given, the average atomic mass can be calculated as follows:

$$\textit{Average atomic mass} = \frac{(\%\ abundance)_1\ Mass_1 + (\%\ abundance)_2\ Mass_2 + \ldots}{100}$$

- Example:
  Carbon-12 → 99%
  Carbon-13 → 1%
  Average atomic mass of carbon
  $= (12 \times 0.99) + (13 \times 0.01)\ g = 12.01\ amu$
- Average atomic mass is commonly reported as the atomic mass for an element.

Finding average atomic mass

What is the **average atomic mass** of chlorine, if it contains two types of atoms having **masses 35 *u* and 37 *u***? The relative abundance of these isotopes in nature is in the ratio **3 : 1**.

75%	**25%**
$^{35}_{17}Cl$ Cholrine	$^{37}_{17}Cl$ Cholrine

Solution

% abundances of Cl^{35} and Cl^{37} are 75% and 25%, respectively.
Hence, average atomic mass

$$= \frac{(75 \times 35 + 25 \times 37)}{100} = 35.5\ u$$

Finding % abundance from average atomic mass

An element X has three isotopes **X^{20}, X^{21}, and X^{22}**. The **percentage abundance of X^{20} is 90%** and **the average atomic mass of the element is 20.11**. The percentage abundance of X^{21} should be _______.

Solution

Given that % abundance of X^{20} is 90%.

Let % abundance of X^{21} be 'a'.

So, % abundance of $X^{22} = 100 - 90 - a = 10 - a$

∴ Average atomic mass of X

$$= \frac{\{90 \times 20 + a \times 21 + (10 - a) \times 22\}}{100}$$

Given, the average atomic mass of X = 20.11

$$\Rightarrow \frac{\{90 \times 20 + a \times 21 + (10 - a) \times 22\}}{100} = 20.11$$

Hence, % abundance of $X^{21} = a = 9\%$

Molecular Mass

- Molecular mass numerically indicates the mass of a molecule.
- Summation of masses of all the atoms that are contained in a molecule gives the molecular mass.

AMMONIA (NH_3)

1 N atom | **3 H atoms**

Molecular mass = $(1 \times 14.00 + 3 \times 1.008)$

= 17.024 u

Finding molecular mass in amu

Find the molecular mass of CO_2 in atomic mass unit.

Solution

Atomic masses of C and O are 12 *u* and 16 *u*, respectively.

Molecular mass of $CO_2 = 12 + 2 \times 16 = 44\ u$

Finding molecular mass in grams

Find the molecular mass of N_2O in grams.

Solution

Atomic masses of N and O are 14 *u* and 16 *u*, respectively.

Molecular mass of N_2O

$= (2 \times 14 + 16)\ u = 44\ u = 44 \times 1.66 \times 10^{-24}\ g$

$= 7.30 \times 10^{-23}\ g$

Finding molecular mass in amu

Find the formula mass of $Mg(NO_3)_2$ in *amu*.

Solution

Atomic masses of Mg, N, and O are 24 u, 14 *u* and 16 *u*, respectively.
Formula mass of $Mg(NO_3)_2$ = 24 + (14 + 16 × 3) × 2 = 148 *u*

Finding molecular mass in grams

Find the molecular/formula mass of the following compounds in grams.
Given: atomic mass of Cu = 63.5, S = 32, Mg = 24, Al = 27, P = 31, Cl = 35.5, Ca = 40, O = 16, and H = 1

a) $CuSO_4$ b) $Mg(OH)_2$ c) $Al_2(SO_4)_3$ d) PCl_5 e) $Ca_3(PO_4)_2$

Solution

a) Formula mass of $CuSO_4$
= $(63.5 + 32 + 16 \times 4) \times 1.66 \times 10^{-24}$
= $159.5 \times 1.66 \times 10^{-24}$
= 2.65×10^{-22} *g*

b) Formula mass of $Mg(OH)_2$
= $(24 + (1 + 16) \times 2) \times 1.66 \times 10^{-24}$
= $58 \times 1.66 \times 10^{-24}$
= 9.63×10^{-23} *g*

c) Formula mass of $Al_2(SO_4)_3$
= $(27 \times 2 + (32 + 4 \times 16) \times 3) \times 1.66 \times 10^{-24}$
= $342 \times 1.66 \times 10^{-24}$
= 5.68×10^{-22} *g*

d) Molecular mass of PCl_5
= $(31 + (35.5 \times 5) \times 1.66 \times 10^{-24}$
= $208.5 \times 1.66 \times 10^{-24}$
= 3.46×10^{-22} *g*

e) Formula mass of $Ca_3(PO_4)_2$
= $\{40 \times 3 + (31 + 4 \times 16) \times 2\} \times 1.66 \times 10^{-24}$
= $310 \times 1.66 \times 10^{-24}$
= 5.14×10^{-22} *g*

BOARDS MAIN

Molar Mass of Molecules

- Similar to how molecular mass is calculated, molar mass of a compound can be calculated by adding molar masses of all the constituent atoms of the molecules.
- For ionic compounds as well, molar mass of formula units can be calculated as the sum of molar masses of the constituent atoms.
- **g-atomic mass** : It is the mass of 1 mole of atoms of a type in grams.
- **g-molecular mass** : It is the mass of 1 mole of molecules of a type in grams.
- **g-ionic mass** : It is the mass of 1 mole of ions of a type in grams.

Finding molar mass of a molecule

Find the **molar mass** of the molecule $\mathbf{N_2}$.

Solution

Molar mass of N atoms = 14 *g*
Molar mass of N_2 = 2 × 14 = 28 *g*

Finding molar mass of a molecule

Find the **molar mass** of the molecule $\mathbf{N_2O_5}$.

Solution

Molar masses of N and O atoms are 14 *g* and 16 g, respectively
Molar mass of N_2O_5 = 2 × 14 + 16 × 5 = 108 *g*

Finding molar mass of a molecule

Calculate the **molar mass** of **hydrated copper sulphate salt ($\mathbf{CuSO_4.5H_2O}$)**.

Solution

Molar masses of Cu, S, O and H atoms are 63.5, 32, 16, and 1 *g*, respectively.

Molar mass of $CuSO_4.5H_2O$

= 63.5 + 32 + (4 × 16) + 5 × {(1 × 2) + 16}

= 249.5 *g*

Finding molar mass of a molecule

Let us consider **an atom 'X' having gram atomic weight 'y'**. Find out the **molar mass of the molecule $\mathbf{X_2O_3}$ in terms of y**.

Solution

Molar mass or gram atomic mass of X is 'y' *g*.
Molar mass of O atoms
= 16 *g*
Molar mass of X_2O_3
= y × 2 + 16 × 3 = (2y + 48) *g*

Mass-Moles Conversion

'Molar mass' g of the substance ⟶ 1 mole of the substance

Therefore,

"W" *g* of the substance contains ⟶ $\dfrac{\mathbf{W}}{\mathbf{Molar\ Mass}}$ moles of the substance

Hence,

$$\text{Moles} = \frac{\textbf{Mass}}{\textbf{Molar Mass}}$$

MOLE CONCEPT

UNDERSTANDING THE LAWS OF YESTERDAY FROM TODAY'S PERSPECTIVE

NOTES

What you already know

- Mole
- Subatomic particles
- Atomic mass
- Average Atomic mass
- Molar mass
- Molecular mass
- Interconversion: Mole - particle within particle

What you will learn

- Mole - mass conversion
- Relative atomic mass
- Relative molecular mass
- Average molecular mass
- Laws of chemical combinations
- Dalton's atomic theory
- Law of conservation of mass
- Law of definite proportions
- Law of multiple proportions

BOARDS

Mole-mass conversion

Mass of the substance present in 1 mol = M *g*
Mass of the substance present in 'n' mol is w = M × n *g* (unitary method)

Hence, the number of moles $n = \frac{w}{M}$ *mol*

Finding moles from the mass of the compound

Calculate the **number of moles** of **$CaCO_3$** in **50 *g*.**

Solution

Step 1:
Calculate the molar mass of $CaCO_3$.

The molar mass of $CaCO_3$ = ∑ Molar mass of elements
= 40 + 12 + (16 × 3) = 100 *g/mol*

Step 2:

Find the number of moles using the given mass and molar mass.

Given, 50 *g* of $CaCO_3$

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

So, moles of $CaCO_3 = \frac{50}{100} = 0.5\ mol$

Finding moles from the mass of an ion

Calculate the **number of moles** of the **anion in 60 *g*** of $\mathbf{CO_3^{2-}}$.

Solution

Step 1:

Calculate the molar mass of $\mathbf{CO_3^{2-}}$.

The molar mass of CO_3^{2-}

$= \sum$ Molar mass of elements

$= 12 + (16 \times 3) = 60\ g/mol$

Step 2:

Find the number of moles using the given mass and molar mass.

Given, 60 *g* of CO_3^{2-}

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

So, moles of $CO_3^{2-} = \frac{60}{60} = 1\ mol$

- While calculating the molar mass of an ion, we can ignore the charge on the ion since electrons have negligible mass.

Finding moles from the mass of an element

How many **moles of He atoms** are present in **8 *g* of He**?

Solution

Given mass of He = 8 *g*

Molar mass of He = 4 *g/mol*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

So, moles of He $= \frac{8}{4} = 2\ mol$

Finding moles from the mass of a compound

How many **moles of MgO** are present in **120 *g* of MgO?**

Solution

Step 1:

Calculate the molar mass of MgO.

Molar mass of MgO

= ∑ Molar mass of elements

= 24 + 16 = 40 *g/mol*

Step 2:

Find the number of moles using the given mass and molar mass.

Given mass of MgO = 120 *g*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

So, moles of MgO = $\frac{120}{4}$ = 3 *mol*

Finding mass from the moles of a compound

Calculate the **mass** of **5 moles** of **$CaCO_3$**.

Solution

Step 1:

Calculate the molar mass of $CaCO_3$.

Molar mass of $CaCO_3$

= ∑ Molar mass of elements

= 40 + 12 + (16 × 3) = 100 *g/mol*

Step 2:

Find the mass using the moles and the molar mass.

Given moles of $CaCO_3$ = 5 *mol*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Mass of $CaCO_3$ = 5 × 100 = 500 *g*

Finding mass from the moles of a compound

The **weight in grams of H_2SO_4 present** in **0.25 mole of H_2SO_4** is:

a. 0.245 b. 2.45 c. 24.5 d. 49.0

Solution

Step 1:

Calculate the molar mass of H_2SO_4.

Molar mass of H_2SO_4

= ∑ Molar mass of elements

= (2 × 1) + 32 + (4 × 16) = 98 *g/mol*

Step 2:

Find the mass using the moles and the molar mass.

Given moles of H_2SO_4 = 0.25

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Mass of H_2SO_4 = 0.25 × 98 = 24.5 *g*

Finding the number of particles from the mass of an element

How many **He atoms** are present in **4 *g*** of **He**?

Solution

Step 1:

Calculate the number of moles of He.

Given, mass of He = 4 *g*

Molar mass of He = 4 *g/mol*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

$$\text{Moles of He} = \frac{4}{4} = 1\ mol$$

Step 2:

Finding number of atoms using moles.

Given moles of He atoms = 1

$$Moles = \frac{Number\ of\ particles}{Avogadro\ number}$$

Number of particles (He atoms) = $1 \times N_A = N_A$

Finding molecules from the mass of a compound

How many **CO_2 molecules** are present in **88 *g*** of **CO_2**?

Solution

Step 1:

Calculate the molar mass of CO_2.

Molar mass of CO_2

$= \sum$ Molar mass of elements

$= 12 + (2 \times 16) = 44\ g/mol$

Step 2:

Find the number of moles.

Given mass of CO_2 = 88 *g*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

$$\text{Moles of } CO_2 = \frac{88}{44} = 2\ mol$$

Step 3:

Find number of molecules using moles.

$$Moles = \frac{Number\ of\ particles}{Avogadro\ number}$$

Number of particles (CO_2 molecules) = $2 \times N_A = 2N_A$

Finding the mass of the compound from the molecules

Calculate the **mass** of **3.011×10^{20} molecules** of **H_2O.**

Solution

Step 1:

Find the moles from the molecules.

Given, 3.011×10^{20} molecules of H_2O

$$Moles = \frac{Number\ of\ particles}{Avogadro\ number}$$

$$\text{Moles of } H_2O = \frac{3.011 \times 10^{20}}{6.022 \times 10^{23}}$$

$$= 0.5 \times 10^{-3}\ mol$$

Step 2:

Find the mass from the number of moles.

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Molar mass of H_2O is 18 *g/mol*

Mass of H_2O $= 0.5 \times 10^{-3} \times 18$

$= 9 \times 10^{-3}\ g$

Finding atoms from the given mass of a compound

Calculate the total **number** of **H-atoms** present in **147 *g*** of $\mathbf{H_2SO_4}$.

Solution

Step 1:

Calculate the molar mass of $\mathbf{H_2SO_4}$.

Molar mass of $H_2SO_4 = \sum$ Molar mass of element

$= (2 \times 1) + 32 + (4 \times 16) = 98\ g/mol$

Step 2:

Find the moles using the given mass and the molar mass.

Given, mass of $H_2SO_4 = 147\ g$

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Moles of $H_2SO_4 = \frac{147}{98} = 1.5\ mol$

Step 3:

Find atoms from moles.

$$Moles = \frac{Number\ of\ particles}{Avogadro\ number}$$

Number of particles (H_2SO_4 molecules) $= 1.5 \times N_A$

Number of H atoms $= 2 \times$ Number of H_2SO_4 molecule $= 2 \times 1.5 \times N_A = 3N_A$

Finding the mass of a compound from electrons

If **$\mathbf{1.325 \times 10^{25}}$ electrons** there are in a **$\mathbf{CO_2}$ sample,** then what is the **mass** of **$\mathbf{CO_2}$ present?**

Solution

Step 1:

Find moles of $\mathbf{CO_2}$ from the number of electrons.

Given,

Total number of electrons $= 1.325 \times 10^{25}$

1 molecule of CO_2 contains $6 + (2 \times 8) =$ 22 electrons

So, moles of CO_2

$= \frac{1}{22N_A} \times 1.325 \times 10^{25} = 1\ mol$

Step 2:

Find the mass of the compound from moles.

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Mass of CO_2 = moles × molar mass $= 1 \times 44 = 44\ g$

Finding the number of electrons from the mass of an element of a compound

If the total **mass of oxygen present in $\mathbf{CuSO_4.5H_2O}$** is **160 *g*,** then calculate the total **number of electrons present in the sample.**

Solution

Step 1:

Find the mass of oxygen in 1 mole of the compound.

1 mole of $CuSO_4.5H_2O$ contains 9 mole of oxygen atoms.

$Moles = \frac{Given\ mass}{Molar\ mass}$

Mass of O atoms in one mole of compound

= moles × molar mass

= 9 × 16 = 144 *g*

Step 2:

Find the moles of the compound from the mass of O atoms.

Given, 160 *g* of oxygen is present in the sample

144 *g* of O atoms are present in 1 mole of $CuSO_4.5H_2O$.

So, moles of $CuSO_4.5H_2O$

$= \frac{160}{144} = 1.11\ mol$

Step 3:

Find number of electrons from moles of compound.

1 molecule of $CuSO_4.5H_2O$ contains = 29 + 16 + (4 × 8) + (10 × 5) = 127 electrons

Electrons in 1.11 mole of $CuSO_4.5H_2O$ sample = 1.11 × 127 mol = 1.11 × 127 N_A = 141.11 N_A

Finding the number of atoms from the mass and the moles of compounds

Which of the following substances contain the **greatest number** of **chlorine atoms?**

a. 5 *g* Cl_2 b. 60 *g* $NaClO_3$ c. 0.1 mole of KCl d. 0.5 mole of Cl_2

Solution

a)

Given, mass of Cl_2 = 5 *g*

Molar mass of Cl_2 = 71 *g/mol*

$Moles\ of\ Cl_2 = \frac{5}{71} = 0.07$

Moles of Cl = 2 x Moles of Cl_2

= 2 × 0.07 = 0.14 *mol*

b)

Given, mass of $NaClO_3$ = 60 *g*

Molar mass of $NaClO_3$ = 23 + 35.5 + (16 × 3)

= 106.5 *g*

$Moles\ of\ NaClO_3 = \frac{60}{106.5} = 0.563\ mol$

Moles of Cl = Moles of $NaClO_3$ = 0.563 *mol*

c)

Given, 0.1 mole of KCl

Moles of Cl = Moles of KCl

Moles of Cl = 0.1 *mol*

d)

Given, 0.5 mole of Cl_2

Moles of Cl = 2 × Moles of Cl_2

Moles of Cl = 2 × 0.5 = 1 *mol*

So, option 'd' is correct.

BOARDS MAIN ADVANCED

Relative atomic mass

It is the ratio of the average mass of the atom to $1/12^{th}$ of the mass of a carbon-12 atom. Since it is a ratio, it is a dimensionless quantity.

Finding the relative atomic mass

What is the **relative atomic mass** of **Na atom?**

Solution

Atomic mass of Na = 23 *amu*

$$\text{Relative atomic mass} = \frac{\textit{Mass of an atom of the element}}{\textit{1/12 * [mass of an atom of C-12 isotope]}} = \frac{23\ amu}{1\ amu} = 23$$

MAIN ADVANCED

Relative molecular mass

It is the ratio of the average mass of the molecule to $1/12^{th}$ of the mass of a carbon-12 atom. Since it is a ratio, it is a dimensionless quantity.

Finding the relative molecular mass

What is the **relative molecular mass** of $\mathbf{H_2SO_4}$?

Solution

Molecular mass of H_2SO_4 = 98 *amu*

$$\text{Relative molecular mass} = \frac{\textit{Mass of an molecule of the element}}{1/12^{th}\ \textit{mass of an atom of C-12 isotope}} = \frac{98\ amu}{1\ amu} = 98$$

Finding the mass of 1 mole of a substance by taking a different reference

If we consider that **$1/8^{th}$ (in place of $1/12^{th}$) the mass of carbon atom** is taken to be the relative atomic mass unit, what will be **the difference in mass of one mole of a substance?**

Solution

Mass depends on the number of particles and the atomic mass of the element.
By changing the reference neither the number of particles nor the atomic mass of the element changes. Therefore, **there is no difference in mass of one mole of a substance.**

BOARDS MAIN ADVANCED

Average molecular mass

When a set of compounds is given along with the number of molecules of each compound, Average atomic mass can be calculated as follows:

$$\textbf{Average molecular mass} = \frac{\textit{(no of molecules)}_1 \times \textit{(Molecular mass)}_1 + \textit{(no of molecules)}_2 \times \textit{(Molecular mass)}_2 + \ldots}{\textit{Total number of particles}}$$

Finding the average molecular mass from the mass of the components

If a gas mixture contains **560 *g*** of N_2 **gas and 320 *g*** of **oxygen gas,** find the **average molecular mass of the mixture.**

Solution

Step 1:

Calculate moles of N_2 and O_2 from their given mass.

Given,

Mass of N_2 = 560 *g*, Mass of O_2 = 320 *g*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

$$\text{Moles of } N_2 = \frac{560}{28} = 20\ mol$$

$$\text{Moles of } O_2 = \frac{320}{32} = 10\ mol$$

Step 2:

Calculate average molecular mass (AMM) from mass and moles of components.

$$AMM = \frac{Total\ mass}{Total\ moles}$$

Total mass = mass of N_2 + mass of O_2

Total moles = moles of N_2 + moles of O_2

$$\text{AMM} = \frac{560 + 320}{20 + 10} = 29.33\ amu$$

Finding the average molecular mass from the mass composition of components

If the **mass composition** of air is given as **76.7 %** of **nitrogen gas and 23.3 %** of **oxygen gas,** find the **average molecular mass of the mixture.**

Solution

Step 1:

Calculate the moles of N_2 and O_2 from their mass composition.

Given,

mass composition of air = 76.7 % nitrogen gas, 23.3 % oxygen gas

Assume total mass of gas = 100 *g*

Mass of N_2 = 76.7 g

Mass of O_2 = 23.3 g

$$\text{Moles of } N_2 = \frac{76.7}{28} = 2.74\ mol$$

$$\text{Moles of } O_2 = \frac{23.3}{32} = 0.73\ mol$$

Step 2:

Calculate the average molecular mass (AMM) from the mass and the moles of the components.

Total mass = mass of N_2 + mass of O_2

Total moles = moles of N_2 + moles of O_2

$$AMM = \frac{Total\ mass}{Total\ moles} = \frac{76.7 + 23.3}{2.74 + 0.73}$$

$$= 28.83\ amu$$

Finding the average molecular mass from the molar composition of the components

If the **molar composition** of air is given as **79% of nitrogen gas and 21% of oxygen gas,** find the **average molecular mass of the mixture.**

Solution

Step 1:

Find moles of N_2 and O_2 assuming that total moles of mixture is 100.

Given molar composition of air is 79% nitrogen gas and 21% oxygen gas

Moles of N_2 = 79, Moles of O_2 = 21

Step 2:

Calculate mass of N_2 and O_2.

Mass of $N_2 = 79 \times 28 = 2212\ g$

Mass of $O_2 = 21 \times 32 = 672\ g$

Step 3:

Calculate average molecular mass (AMM) from mass and moles of components.

Total mass = mass of N_2 + mass of O_2

Total moles = moles of N_2 + moles of O_2

$$AMM = \frac{Total\ mass}{Total\ moles} = \frac{2212 + 672}{79 + 21}$$

$$= 28.84\ amu$$

BOARDS MAIN

Laws of Chemical Combinations

Dalton's Atomic Theory

Postulates

- Matter consists of **tiny particles** known as **atoms** that are indivisible.
- All the atoms of a given element have **identical properties** including **identical mass**.
- Atoms of different elements differ in mass and they combine in a **fixed** ratio to form compounds.
- Chemical reactions involve **rearrangement** of atoms.

Drawbacks

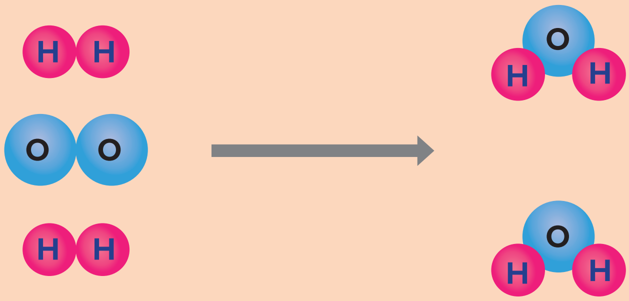
- According to the theory, an atom is indivisible but they can be **divided** into **electrons, neutrons,** and **protons**.
- Atoms of the same element can have **different masses** and atoms of different elements can have same mass as well, as in the case of **isotopes** and **isobars**.
- A complex reaction **does not** react with a simple whole number **ratio**.

BOARDS MAIN

Law of Conservation of Mass

It states that **matter can neither be created nor destroyed** in ordinary chemical and physical changes. In a chemical reaction, the **total mass of the reactants** is always equal to the **total mass of the products** formed.

Example



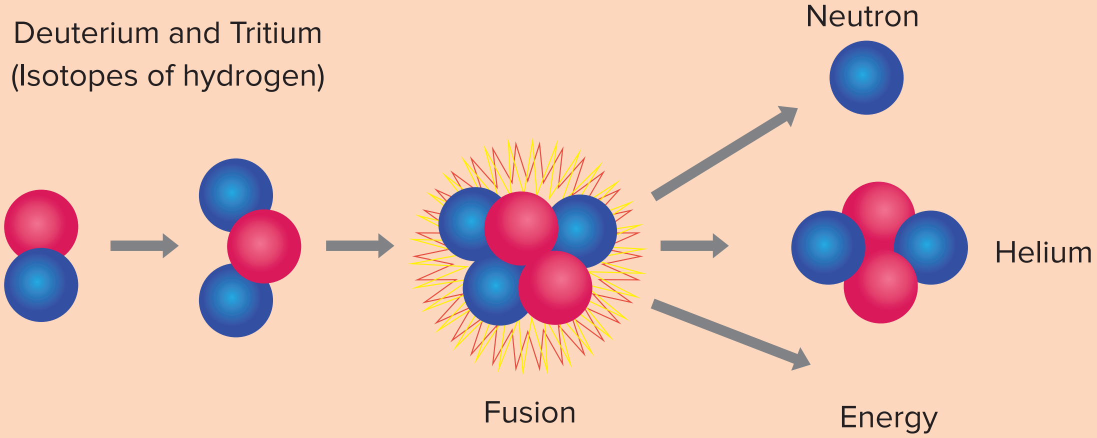
Reactants:
4 H + 2 O
atom

Products:
4 H + 2 O
atom

- The number of H and O atoms are the same on either side of the reaction.
- The number of molecules is not conserved. On the reactant side, there are three molecules and on the product side there are two molecules.

Exception

Nuclear reactions are an exception to the law of conservation of mass.

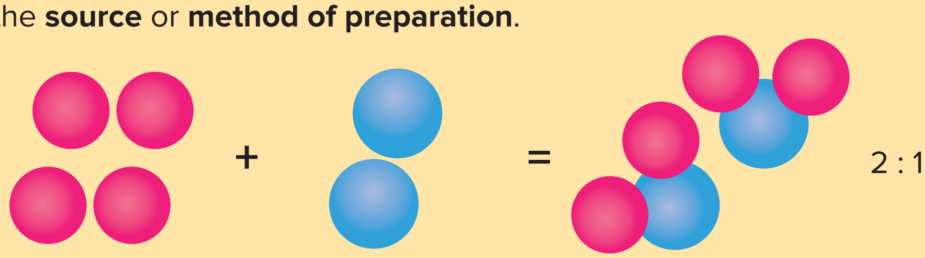


- In nuclear reactions, some of the mass is converted into energy and hence the mass is not conserved.

BOARDS MAIN

Law of Definite Proportions

A given compound always contains **exactly the same proportion** of elements by weight **irrespective** of the **source** or **method of preparation**.



4 parts
Hydrogen

2 parts
Oxygen

2 gaseous
water

2 H atom : 2 *amu*
+
1 O atom : 16 *amu*

Molecular weight of water = 18 *amu*
% Hydrogen by weight = 11.11 %
% Oxygen by weight = 88.89 %

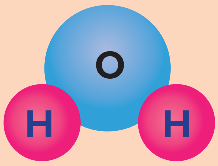
The ratio of mass of H to the oxygen in water is constant. This means irrespective of the source it is obtained, the mass % of H in water is 11.11% and the mass % of oxygen is 88.89%.

BOARDS MAIN

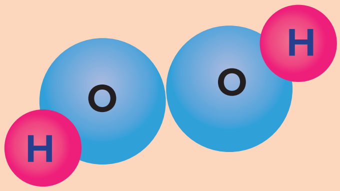
Law of Multiple Proportions

If **two elements can combine** to form more than one compound, for a **fixed mass of any one element** in both the compounds, the **ratio of masses of the other element** in the two compounds comes out to be in **small whole numbers.**

Example



$$\frac{H}{O} = \frac{2}{1}$$



$$\frac{H}{O} = \frac{1}{1}$$

In the above example,

Ratio of mass of H : O in water is 2 : 16

Ratio of mass of H : O in hydrogen peroxide is 2 : 32 = 1 : 16

Ratio of water and hydrogen peroxide for a fixed mass of oxygen is 2 : 1

When two elements H and O combine to form more than one compound, water and hydrogen peroxide, the masses of hydrogen that combine with the fixed mass of oxygen will also be in a whole number ratio.

MOLE CONCEPT

MORE INTO LAWS

What you already know

- Dalton's atomic theory
- Law of conservation of mass
- Law of definite proportions
- Relative atomic mass
- Relative molecular mass
- Law of multiple proportions

What you will learn

- Law of reciprocal proportions
- Gay Lussac's law
- Avogadro's law
- Relative density
- Specific gravity
- Vapour density
- Molar volumes at STP, NTP and SATP

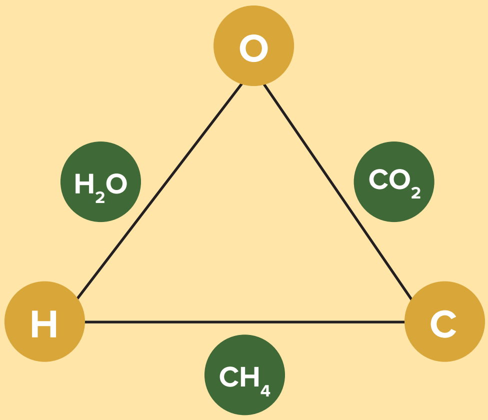
BOARDS

Law of Reciprocal Proportions

Statement

If an **element A** combines with an **element B** in the ratio of **x : y** and **A** combines with another **element C** in the ratio of **x : z**, then **B** and **C** combine in the ratio of **y:z** or a simple whole number multiple of it.

Example



Carbon combines with oxygen and forms carbon dioxide and it combines with hydrogen to form methane. Hydrogen and oxygen combine to form water.

Mass ratio of **C : H = 12 : 4 = 3 : 1;**

Mass ratio of **C : O = 12 : 32 = 3 : 8;**

Mass ratio of **H : O = 2 : 16 = 1 : 8**

So here, **C** combines with **H** in the ratio of **3 : 1** and C combines with O in the ratio of **3 : 8**. **H** and **O** combine in the ratio of **1 : 8**. The above example clearly illustrates the law of reciprocal proportions.

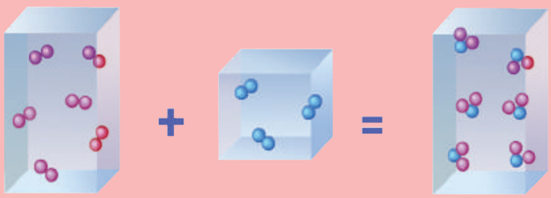
MAIN

Gay-Lussac's Law

Statement

When gases combine or are produced in a chemical reaction, they do so in a **simple ratio** by **volume** provided all gases are at the **same temperature and pressure**.

Example



$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)}$$

We already know the mole ratio of reactants and products in a chemical reaction is in the whole number ratio. According to Gay Lussac's law, the volume ratio of the reactants and products in a chemical reaction is also in the whole number ratio of their number of moles under the conditions of the same temperature and pressure.

$$3H_{2(g)} + N_{2(g)} \longrightarrow 2NH_{3(g)}$$

Here, 3 volumes of hydrogen combine with 1 volume of nitrogen to give 2 volumes of ammonia under the same conditions of temperature and pressure.

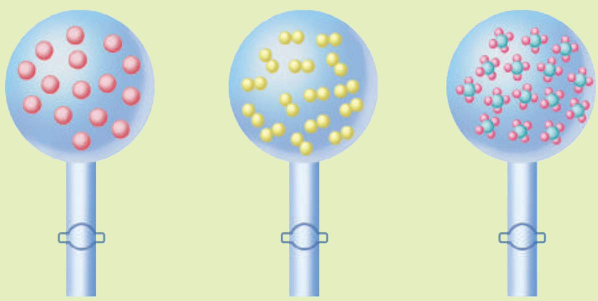
BOARDS MAIN

Avogadro's Law

Statement

Equal volumes of gases at the **same temperature and pressure** should contain **equal** numbers of **molecules/ particles**.

The number of particles in all the bulbs is 15 and the volumes are identical. This can also be stated as, the volume of the gas is directly proportional to the number of particles.



$$\textbf{V} \propto \textbf{number of particles} \Rightarrow \frac{V_1}{n_1} = \frac{V_2}{n_2}$$

BOARDS

Relative Density

It is the density of a substance **relative** to the density of another substance at the same T and P. It is a **dimensionless** quantity.

$$\textbf{R.D.} = \frac{\rho_{substance}}{\rho_{reference}}$$

Specific Gravity

- Specific gravity for **liquids** is measured with respect to **water** at 4 °C.

$$\textbf{S.G. (for liquid)} = \frac{\rho_{substance}}{\rho_{water} \text{ at } 4\ ^\circ C}$$

- It is numerically equal to the density of the liquid in g/mL.

- For **gases**, it is measured with respect to **air** at STP. It is a **dimensionless** quantity.

$$\textbf{S.G. (for gas)} = \frac{\rho_{gas} \text{ at STP}}{\rho_{air} \text{ at STP}}$$

Vapour Density

- Density of the gas with respect to **hydrogen** gas at the **same** temperature and pressure.

It indicates **how heavy a gas is** with respect to the **lightest gas.**
It is also a **dimensionless** quantity.

$$V.\ D. = \frac{\textit{Density of gas A at same Temperature and Pressure}}{\textit{Density of } H_2 \textit{ gas at same Temperature and Pressure}} = \frac{\text{molar mass of gas}}{2}$$

Relative density, specific gravity and vapour density are dimensionless quantities.

MAIN

Molar volumes

(It is the volume occupied by 1 mole of the gas)

Molar volume at STP (Standard Temperature and Pressure)

At STP condition: Temperature = **0°C** or **273 *K*;** Pressure = **1 *Bar*;** Molar volume at STP = **22.7 *L***

1 *atm* = 1.01 *Bar*

Calculation of Molar volume at STP

$\rho(H_2$ at **STP**$) = 0.089\ g/L$

Now, using Density $= \frac{M}{V_m} = 0.089\ g/L = \frac{2 g/mol}{V_m}$

Hence, molar volume $V_m = \frac{2}{0.089}\ L/mol$

$\mathbf{V_m = 22.4}$ ***L*** **at STP**

Initially 1 atm was taken as standard P and there V_m turns out to be 22.4 L.

Molar volume at NTP (Normal Temperature and Pressure)

At NTP condition: Temperature = **20°C or 293 *K***; Pressure = **1 *atm*;** Molar volume at NTP = **24 *L***

SATP (Standard Ambient Temperature and Pressure)

At SATP condition: Temperature = **25°C or 298 *K***; Pressure = **1 *atm*;** Molar volume at SATP = **24.8 *L***

1. Molar volume is the volume occupied by one mole of the gas.
2. Molar volume at STP is 22.7 *L*(1 bar) or 22.4 *L* (1 atm)
3. Molar volume at NTP is 24 *L*.
4. Molar volume at SATP is 24.8 *L*.

Finding the vapour density of methane.

Calculate the **vapour density** of 20 *g* of **methane gas.**

Solution

$$\textit{Vapour density} = \frac{\textit{molar mass}}{2}$$

Molar mass of methane (CH_4) = 12 + (1 × 4) = 16 *g*

Hence, vapour density of methane = $\frac{16}{2} = 8$

Vapour density does not depend upon the given mass.

Finding the vapour density of methane.

Calculate the **vapour density** of 30 *g* of **methane gas.**

Solution

As we know $\textit{Vapour density} = \frac{\textit{molar mass}}{2}$

Molar mass of methane (CH_4) = 12 + (1 × 4) = 16 *g*

Hence, vapour density of methane = $\frac{16}{2} = 8$

Vapour density of a gas.

The **relative density** of **gas A** with respect to another **gas B** is **2**. If the **vapour density** of **gas B** is **20**, the **vapour density of gas A** is:

Solution

Step 1:

Find the molar mass of gas B (M_B)

Given,

Vapour density of gas B = 20

$\frac{M_B}{2} = 20$ $\therefore M_B = 40\ g$

Step 3:

Find the vapour density of gas A

$Vapour\ density = \frac{molar\ mass}{2}$

$\therefore$ Vapour density of gas A = $\frac{80}{2} = 40$

Step 2:

Find the molar mass of gas A (M_A)

Relative density of gas A with respect to gas B = $\frac{\text{Density of gas A at same T and P}}{\text{Density of gas B at same T and P}}$

Relative density of gas A = $\frac{\text{Mass of gas A in 1 } ml \text{ at same T and P}}{\text{Mass of gas B in 1 } ml \text{ at same T and P}}$

$\therefore 2 = \frac{\text{N particles x (mass of one particle of gas A)}}{\text{N particles x (mass of one particle of gas B)}} = \frac{M_A\ amu}{M_B\ amu} = \frac{M_A}{40}$

$\therefore M_A = 80\ g$

Vapour Density

The density of ammonia is **0.77 *g/L***. Calculate its **vapour density at STP (1 atm, 273 K).**

Solution

Step 1:

Find the molar mass

Given density of ammonia = 0.77 *g/L*

$density = \frac{Molar\ mass}{Molar\ volume\ at\ STP}$

$0.77\ g/L = \frac{\text{Molar mass}}{22.4\ L}$

Molar mass = 0.77 × 22.4 *g* = 17.248 *g*

Step 2:

Find the vapour density

$Vapour\ density = \frac{molar\ mass}{2} = \frac{17.248}{2} = 8.624$

Finding Vapour Density from Molar mass.

N_2 has a **molecular mass** of **28 *amu***. What will be its **vapour density**?

Solution

Given,

Molecular mass = 28 *amu*

$\therefore Vapour\ density = \frac{molar\ mass}{2} = \frac{28}{2} = 14$

Finding vapour density of the mixture.

Calculate the **vapour density** of a mixture containing **44 *g* of CO_2** and **30 *g* of ethane**.

Solution

Step 1:

Find the mole of all component

Given

Mass of CO_2 = 44 *g*

Mass of ethane = 30 *g*

So,

$$\textit{Mole of } CO_2 = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{44}{44} = 1$$

$$\textit{Mole of ethane} = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{30}{30} = 1$$

Step 2:

Find Average Molar Mass

Total mass = (44 + 30) *g* = 74 *g*

Total moles = $n_{CO_2} + n_{C_2H_6} = 1 + 1 = 2\ mol$

$$\textit{Average Molar mass} = \frac{\textit{Total mass}}{\textit{Total moles}} = \frac{74}{2} = 37\ g$$

Step 3:

Find Vapour density

$$\therefore \textit{Vapour density} = \frac{\textit{molar mass}}{2} = \frac{37}{2} = 18.5$$

Finding the ratio of molar mass from the ratio of vapour density.

Vapour density of two gases is in the ratio of **1:3**. What will be the **ratio of their molar masses**?

Solution

Let the two gases be A and B.

Given :

$$\frac{\text{Vapour density of gas A}}{\text{Vapour density of gas B}} = \frac{1}{3}$$

We know, $\textit{Vapour density} = \dfrac{\textit{molar mass}}{2}$

$$\text{So, } \frac{(\text{molar mass of gas A}) \times 2}{(\text{molar mass of gas B}) \times 2} = \frac{1}{3}$$

$$\therefore \frac{\text{molar mass of gas A}}{\text{molar mass of gas B}} = \frac{1}{3}$$

Hence, the ratio of the molar masses of gas A and gas B is 1:3.

Finding the number of molecules from vapour density

Gas A has vapour density of 8.55. Find the number of molecules in 0.5 moles of the gas.

Solution

Given:

Number of moles of gas molecules = 0.5 *mol*

Number of moles = $\frac{\text{Number of Particles (given)}}{N_A}$ $\Rightarrow 0.5 = \frac{\text{Number of molecules}}{N_A}$

Hence, Number of molecules = 0.5 N_A (Vapour density of the gas was not required)

Vapour density from mass

A gas cylinder holds **85 *g* of gas X**. The **same** cylinder, when filled with **hydrogen**, holds **8.5 *g*** of **H_2** at the same temperature. Find its **vapour density**.

Solution

Let the volume of the cylinder be 'V' *mL*

Given:

Mass of gas X in 'V' *mL* = 85 *g*

Mass of gas H_2 in 'V' *mL* = 8.5 *g* (at the same temperature)

$\therefore$ Vapour density = $\frac{\text{mass of gas X in V mL at T and P}}{\text{mass of } H_2 \text{ gas in V mL at same T and P}} = \frac{85}{8.5} = 10$

Finding the vapour density of the mixture.

A gaseous mixture of **H_2** and **CO_2** gas contains **66% by mass of CO_2**. What is the **Vapour Density** of the mixture?

Solution

Step 1:

Find the mol of all component

Let the mass of the mixture = 100 *g*

Given:

Mass of CO_2 in the mixture = 66%

Mass of H_2 = (100 - 66) g = 34 *g*

Mol of CO_2 = $\frac{\text{Given mass}}{\text{Molar mass}} = \frac{66}{44} = 1.5\ mol$

Mol of H_2 = $\frac{\text{Given mass}}{\text{Molar mass}} = \frac{34}{2} = 17\ mol$

Step 2:

Find Average Molar Mass

Total mass = 100 *g*

Total moles = $n_{CO_2} + n_{H_2} = 1.5 + 17 = 18.5\ mol$

Average Molar mass = $\frac{\textit{Total mass}}{\textit{Total moles}} = \frac{100}{18.5} = 5.41\ g$

Step 3:

Find Vapour density

$\therefore$ *Vapour density* = $\frac{\textit{molar mass}}{2} = \frac{5.41}{2} = 2.7$

C H E M I S T R Y

MOLE CONCEPT

INTERCONVERSION FINALE AND PERCENTAGES

What you already know

- Mole-Mass Interconversion
- Relative density, Specific gravity and Vapour density
- Laws of Chemical Combinations
- STP and NTP Conditions
- Molar Volume at STP and NTP

What you will learn

- Interconversion of Mass and Number of Particles
- Interconversion of Particles-Volume
- Empirical and Molecular Formula
- Interconversion of Mole-Volume
- Interconversion of Mass-Volume
- Percentage Composition

BOARDS MAIN

Interconversion of Mole-Volume

$$\text{Volume at STP} \underset{\div\, 22.4\, L}{\overset{\times\, 22.4\, L}{\rightleftarrows}} \text{Mole}$$

- **22.4** ***L*** of any gas at **STP (1** ***atm*****, 0° *C*)** is the molar volume ($\mathbf{V_m}$, Volume of one mole of gas)
- **11.2** ***L*** of a gas at **STP** has 0.5 moles Or, 2 moles of a gas at **STP** contains 44.8 *L* volume.

$$\text{[illegible]ume at STP} = \frac{\text{Volume of gas (L, STP)}}{\text{Molar volume (22.4 L)}} = \frac{\text{Volume of gas (mL, STP)}}{22400\ mL}$$

Molar volume is only defined for gases and not for liquids or solids.

Mole - volume interconversion

At **STP, 11.2 *L*** of CO_2 gas contains:

Solution

Using the interconversion of mole-volume relationship,
The number of moles,

$$n = \frac{\text{Volume of gas (at STP)}}{22.4\ L}$$

$$n = \frac{11.2}{22.4}\ mol$$

n = 0.5 *mol*

Mole - volume interconversion

What volume would **0.735 moles** of O_2 gas occupy at **1 *atm*** and **0° C** ?

Solution

In the question, 1 *atm* and 273 *K* signifies STP condition.
So, **using the interconversion of mole-volume relationship**, we have,
$V\ (STP) = n \times 22.4\ L$
So,
$V = 22.4 \times 0.735 = 16.46\ L$

Mole - volume interconversion

What volume is occupied by **5 moles** of a gas at **STP**?

Solution

We can find the volume of gas at STP using the **interconversion of mole-volume relationship**.
$V\ (STP, L) = n \times 22.4$,
So,
$V = 22.4 \times 5 = 112\ L$

Mole - volume interconversion

The number of moles of H_2 in **0.24 *L*** of Hydrogen gas at **NTP** is:

Solution

Volume of gas occupied at NTP = 24 *L*
Using the interconversion of mole-volume relationship,
The number of moles,

$$n = \frac{\text{Volume of gas (at NTP)}}{24}$$

$$n = \frac{0.24}{24}\ mol$$

n = 0.01 *mol*

Mole - volume interconversion

How many moles are there in **29.4 *L*** of gaseous ethane at **NTP**?

Solution

Molar volume at NTP = 24 *L*
Using the **interconversion of mole-volume relationship,**
The number of moles,

$$n = \frac{\text{Volume of gas (at NTP)}}{24}$$

$$n = \frac{29.4}{24}\ mol$$

n = 1.225 *mol*

MAIN

Interconversion of Mass-Volume

- If the mass of a gas is given,

$$\text{Moles} = \frac{\text{mass}}{\text{atomic mass (or molecular mass)}}$$

Volume of gas at STP = Moles × Molar volume
= Moles × 22.4 L

- If the volume of gas is given,

$$\text{Moles} = \frac{\text{Volume of gas}}{\text{Molar volume}} = \frac{\text{Volume of gas at STP}}{22.4\ L}$$

Mass = Moles × Molecular mass

Mass - volume interconversion

What **volume (in *L*)** is occupied by **60 *g*** of **ethane** at **STP**?

Solution

Step 1: Finding the number of moles of ethane, using the formula

$$n = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$= \frac{60}{30} = 2\ mol$$

Step 2: Using the interconversion of mole-volume relationship,

$V(STP, L) = n \times 22.4,$

So, V = 22.4 × 2 = 44.8 *L*

Mass - volume interconversion

What **volume (in *mL*)** is occupied by **142 grams** of chlorine gas at **STP**?

Solution

Step 1: Finding the number of moles of chlorine gas, using the formula

$$n = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$= \frac{142}{71} = 2\ mol$$

Step 2: Using the interconversion of mole-volume relationship,

$V(STP, L) = n \times 22.4,$

So, V = 22.4 × 2 = 44.8 *L*

In the question, volume in *mL* is asked.

So, we have 1 *L* = 1000 *mL*

Therefore, V = 44800 *mL*

Mass - volume interconversion

What is the **mass in grams** of **5.6 *L*** of ethene gas at **STP**?

Solution

Step 1: Finding the number of moles of ethene gas

Using the interconversion of mole-volume relationship,

The number of moles,

$$n = \frac{\textit{Volume of gas (at STP)}}{22.4\ L}$$

$$n = \frac{5.6}{22.4}$$

n = ¼ *mol* = 0.25 *mol*

Step 2: Using the interconversion of mole-volume relationship,

$$n = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$0.25 = \frac{m}{28}$$

m = 0.25 × 28 = 7 *g*

Mass - mole - volume interconversion and molecular formula determination

16 *g* of an ideal gas **SO_x** occupies **5.6 *L*** at **STP**. The value of **x** is:

Solution

Step 1: Finding the number of moles of SO_x gas

Using the interconversion of mole-volume relationship,
We have, the number of moles,

$$n = \frac{\text{Volume of gas (STP, L)}}{22.4}$$

$$n = \frac{5.6}{22.4}\ mol$$

n = ¼ *mol* = 0.25 *mol*

Step 2: Finding the value of x

0.25 *mol* of gas contains 16 *g* of an ideal gas, SO_x

Then, 1 *mol* of gas will contain 64 *g* of SO_x

Thus, 64 *g* is the molar mass of SO_x
Molar mass = 32 + 16(x) = 64

x = 2

Molecular weight determination form molar volume at STP

The **ratio of the weight** of one litre of a gas to the weight of one litre oxygen gas both measured at **STP** is **2.22**. The molecular weight of the gas would be:

Solution

Given :

$$\frac{\text{mass of 1 litre of the gas at STP}}{\text{mass of 1 litre of oxygen at STP}} = 2.22$$

Mass of 1 litre of the gas at STP = 2.22 × mass of 1 litre of the oxygen at STP
Then,
Mass of 22.4 litres of the gas at STP = 2.22 × mass of 22.4 litres of the oxygen at STP

$$\frac{\text{(molar mass of gas X)}}{\text{molar mass of } O_2 \text{ (32 } g/mol)} = 2.22$$

molar mass of gas X = 32 × 2.22 = 71.04 *g/mol*

1 *kg* = 1000 *g* = 10^6 *mg* (milligrams)
1 *L* = 1000 *mL* = 1000 cm^3 = 10^{-3} m^3

Interconversion of Particles-Volume

- If the number of particles of a gas are given,

$$Moles = \frac{Number\ of\ particles}{N_A}$$

$$Volume\ of\ gas = Moles \times Molar\ volume$$
$$= Moles \times 22.4\ L$$

- If the volume of gas is given,

$$Moles = \frac{Volume\ of\ gas}{Molar\ volume} = \frac{Volume\ of\ gas}{22.4\ L}$$

$$Number\ of\ particles = Moles \times N_A$$

Particles - volume interconversions

Calculate the **number of atoms/molecules** present in the following at **STP** conditions.

(a) 22.4 *L* of CO_2

(b) 0.224 *L* of He

(c) 44.8 *L* of CH_4

(d) 112 *L* of He

Solution

(a) 22.4 *L* of CO_2

1 *mol* of CO_2 occupies 22.4 *L* at STP

1 *mol* of CO_2 contains N_A number of molecules

(b) 0.224 *L* of He

Using the interconversion of mole-volume relationship,

The number of moles,

$$n = \frac{Volume\ of\ gas\ (STP,\ L)}{22.4}$$

$$n = \frac{0.224}{22.4} = 0.01\ mol$$

Number of atoms $= 0.01 \times N_A$

$= 0.01 \times 6.023 \times 10^{23}$

$= 6.023 \times 10^{21}$

(c) 44.8 *L* of CH_4

Using the interconversion of mole-volume relationship,

The number of moles,

$$n = \frac{Volume\ of\ gas\ (at\ STP,\ L)}{22.4}$$

$$n = \frac{44.8}{22.4} = 2\ mol$$

Number of molecules = $2 \times N_A = 2 \times 6.023 \times 10^{23} = 1.2046 \times 10^{24}$

(d) 112 *L* of He

Using the interconversion of mole-volume relationship,
The number of moles,

$$n = \frac{\textit{Volume of gas (STP, L)}}{22.4}$$

$$n = \frac{112}{22.4} = 5\ mol$$

Number of atoms = $5 \times N_A = 5 \times 6.023 \times 10^{23} = 3.0115 \times 10^{24}$

Particles - volume interconversion

4.4 *g* of CO_2 and **2.24 *L* of H_2** at **STP** are mixed in a container. The **total number of molecules** present in the container will be:

Solution

Step 1: Finding the number of moles of CO_2
Using the interconversion of mole-mass relationship,
We have, the number of moles,

$$n = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$= \frac{4.4}{44} = 0.1\ mol$$

Step 2: Finding the number of moles of H_2
Using the interconversion of mole-volume relationship,

$$n = \frac{\textit{Volume of gas (STP, L)}}{22.4}$$

$$n = \frac{2.24}{22.4} = 0.1\ mol$$

Step 3: Total number of molecules
Total moles of gas ($CO_2 + H_2$) = 0.2 *mol*
The total number of atoms = $0.2 \times N_A = 12.046 \times 10^{22}$

Particles - volume interconversion

A sample of propane gas has a volume of **2.8 *L*** at **STP**. Find the **number of particles** in the sample.

Solution

Step 1: Finding the number of moles
Using the interconversion of mole-mass relationship,
We have, the number of moles,

$$n = \frac{\textit{Volume of gas (STP, L)}}{22.4}$$

$$n = \frac{2.8}{22.4} = 0.125\ mol$$

Step 2: Finding the number of molecules

Number of atoms = 0 .125 × N_A

= 0.125 × 6.023 × 10^{23}

= 7.528 × 10^{22}

Particles - volume interconversion

From **160 *g*** of a **SO_2 (*g*)** sample, **1.2046 × 10^{24}** molecules of **SO_2** are removed. Find out the volume of left over **SO_2 (*g*)** at **STP**.

Solution

Step 1: Finding the number of moles of SO_2

Using the interconversion of mole-mass relationship,

We have, the number of moles,

$$n = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$= \frac{160}{64} = 2.5\ mol$$

Step 2: Finding moles of SO_2 removed

Now, the number of moles of SO_2 removed

$$= \frac{1.2046 \times 10^{24}}{N_A}$$

$$= \frac{12.046}{6.023} \times 10^0 = 2\ mol$$

Step 3: Number of moles remaining

The number moles remaining = 2.5 - 2 = 0.5

Step 4:

Volume of SO_2 left = moles × 22.4

= 0.5 × 22.4 = 11.2 *L*

Particles - volume interconversion

Calculate the **number of electrons** in **44.8 *L*** of **H_2SO_4 (*g*)** at **STP**.

Solution

Step 1: Find the moles of H_2SO_4

Given:

Volume = 44.8 *L*

$$Moles = \frac{\textit{Volume of gas at STP}}{\textit{Molar volume at STP}}$$

$$\text{Moles} = \frac{44.8}{22.4}$$

Moles of H_2SO_4 = 2 *mol*

Step 2: Find the number of electrons of H_2SO_4

1 mole of H_2SO_4 contains

= (2 × 1) + 16 + (4 × 8)

= 50 mole of electrons

1 mole (N_A molecules of H_2SO_4) contains 50 N_A electrons.

2 mole (N_A molecules of H_2SO_4) contains 2 x 50 N_A electrons.

= 100 N_A

Particles - volume interconversion

A pure sample of $\mathbf{CO_2}$ contains a total of $\mathbf{2.2 \times 10^{10}}$ **electrons.** Find the volume occupied by $\mathbf{CO_2}$ at **STP.**

Solution

Step 1: Find the moles of CO_2

Given:

number of electrons = 2.2×10^{10}

$6 + (2 \times 8) = 22$ mole of electrons are present in 1 mole of CO_2 or 22 N_A of electrons are present in 1 mole of CO_2

2.2×10^{10} electrons are present in

$$= \frac{2.2 \times 10^{10}}{22\ N_A}$$

$$= \frac{10^9}{N_A} \text{ moles of } CO_2$$

Step 2: Find the volume of CO_2

Volume of 1 mole of CO_2 = 22.4 *L*

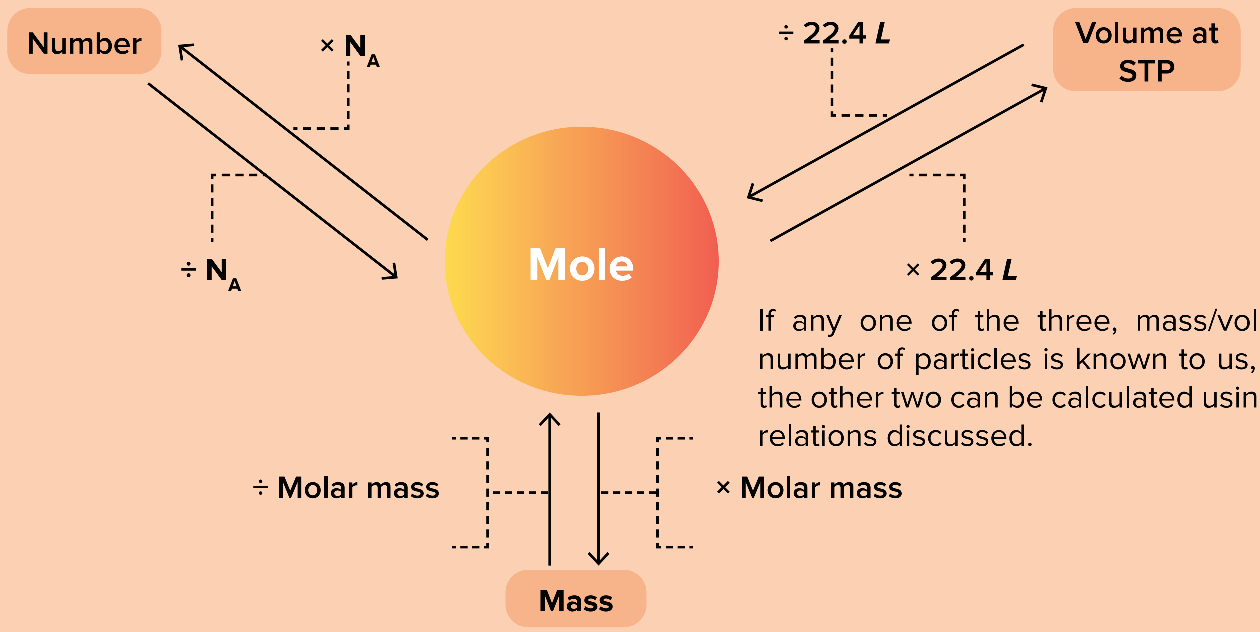
Volume of $\frac{10^9}{N_A}$ mole of CO_2

$$= \frac{10^9}{N_A} \times 22.4\ L$$

$$= 3.719 \times 10^{-14}\ L$$

Summary

MAIN



If any one of the three, mass/volume/number of particles is known to us, then the other two can be calculated using the relations discussed.

BOARDS MAIN

Percentage Composition

It is the relative mass of each of the constituent elements in 100 parts of the compound.

$$\textit{Mass \% of an element} = \frac{\textit{mass of the element in 1 mole of compound}}{\textit{Molar mass of compound}} \times 100\ \%$$

Percentage composition determination

Calculate the **mass percentage** of **sulphur** in **H_2SO_4**.

(a) 49.5 % | **(b) 50.1 %**
(c) 45.6 % | **(d) 32.6 %**

Solution

Molar mass of H_2SO_4 = 98 *g/mol*

Molar mass of S = 32 *g/mol*

$$\text{Mass \% of element} = \frac{\textit{mass of the element in 1 mole of compound}}{\textit{Molar mass of compound}} \times 100$$

$$\text{Mass \% of S} = \frac{32}{98} \times 100$$
$$= 32.65\%$$

Molar mass determination from percentage composition

A compound contains **8% sulphur by mass. The least molecular mass is:**

(a) 100 *g/mol* | **(b) 400 *g/mol***
(c) 155 *g/mol* | **(d) 256 *g/mol***

Solution

Step 1: Relating mass percentage to the molar mass of the compound

Given:

Mass % of S = 8

Molar mass of S = 32 *g/mol*

Let the molar mass of the compound be x *g/mol*.

Mass % of element =

$$\frac{\textit{mass of the element in 1 mole of compound}}{\textit{Molar mass of compound}} \times 100\ \%$$

Step 2 : Molar mass of the compound

Minimum molecular mass is possible when 1 atom of S is present in the compound.

$$\text{Mass \% of S} = 8 = \frac{32}{x} \times 100\ \%$$

x = 400 *g/mol*

So option 'b' is correct

Molar Mass vs Molecular Mass

Molar mass is the mass of 1 mole of substance. It gives the measure of number of molecules/atoms/compounds present in a mole of substance. Its SI unit is ***g/mol***.

Molecular mass is the mass of 1 molecule. Its unit is ***amu***.

Generally, both the masses have the same values and only units are different.

BOARDS MAIN

Empirical and Molecular Formula

- Molecular formula tells us the exact number of atoms of each element in a molecule. It is obtained by using the molar mass of each element.

Eg: 6, 12 & 6 atoms of C, H & O are present in a molecule of **Glucose ($C_6H_{12}O_6$)**.

- Empirical formula represents the simplest whole number ratio of various atoms present in a compound. It is determined by the mass percent of various elements present in the compound.

Eg: In Glucose, the ratio of number of atoms of C, H & O is respectively,

$$6 : 12 : 6 \Longrightarrow 1 : 2 : 1$$

Its empirical formula is **CH_2O**

Molecular formula and Empirical formula are related as,

$$\textbf{\textit{Molecular Formula = n × Empirical Formula}}$$

$$\Longrightarrow n = \frac{\textit{Molecular Formula}}{\textit{Empirical Formula}} = \frac{\textit{molecular formula mass}}{\textit{empirical formula mass}}$$

Where, 'n' is the factor by which empirical and molecular formulas differ.

Step by step: Getting to Empirical formula if the mass % of the constituent elements is given,

Mass % of elements	$\xrightarrow{\text{(assume 100 } g \text{ sample)}}$	Mass of elements	$\xrightarrow{\text{(divided by molar mass)}}$	Moles of elements	$\xrightarrow{\text{(simplest ratio)}}$	Empirical formula

Molecular formula determination from percentage composition

A compound on analysis gave the following results: **C = 39.99 %, H = 6.71 %** and **O = 53.28 %**. Determine the **molecular formula** of the compound if the **molar mass** is 180 *g/mol*.

Solution

Step 1: Find the moles of each element

Given:

mass % **of C = 39.99 %, H = 6.71 % and O = 53.28 %.**

Assuming mass of compound to be 100 *g*

$$Moles = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

Molar mass of C = 12 *g/mol*, H = 1 *g/mol* and O = 16 *g/mol*.

Mass of C = 39.99 *g*, moles of C $= \frac{39.99}{12} = 3.33$

Step 2 : Molar mass of the compound

For making the elements in a simple whole number ratio, we divide by the element having the least moles (by carbon)

Simple ratio of C $= \frac{3.33}{3.33} = 1$

Simple ratio of H $= \frac{6.71}{3.33} = 2$

Simple ratio of O $= \frac{3.33}{3.33} = 1$

Mass of H = 6.71, moles of H = $\frac{6.71}{1}$ = 6.71

Mass of O = 53.28 *g*,

moles of O = $\frac{53.28}{16}$ = 3.33

Empirical formula becomes = CH_2O

Step 3: Find the molecular mass

Empirical formula mass = 12 + (1 × 2) + 16 = 30

Given: molar mass = 180 *g/mol*.

$$n = \frac{\textit{Molecular formula mass}}{\textit{empirical formula mass}} = \frac{180}{30} = 6$$

Molecular formula (MF) = n × Empirical formula (EF)

= 6 × (CH_2O) = $C_6H_{12}O_6$

Mass ratio determination from atom ratio

Calculate the **weight ratio** of **Ca** and **Br** in a compound with **Ca** and **Br** in **1 : 2** atom ratio.

Solution

Given:

Atom Ratio of Ca : Br is 1 : 2.

Molar mass of Ca = 40 *g/mol*.

Molar mass of Br = 80 *g/mol*.

$$\textit{Moles} = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

Mass = moles × molar mass

Mass ratio of Ca : Br is (1 × 40) : (2 × 80)

40 : 160 = 1 : 4

Molecular formula determination from percentage composition

Calculate the **molecular formula** of a compound which contains **20 % Ca** and **80 % Br** (by wt.), if the **molecular weight** of the compound is **200**. (Atomic weight **Ca = 40, Br = 80**)

(a) $Ca_{1/2}Br$ (b) $CaBr_2$ (c) Ca_3Br (d) Ca_2Br

Solution

Step 1: Find the moles of each element

Given:

mass % **of Ca** = **20 %, Br** = **80** %

Assuming mass of compound to be 100 *g*

$$\text{Moles} = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

Molar mass of Ca = 40 *g/mol*,

Molar mass of Br = 80 *g/mol*.

Mass of Ca = 20 *g*, moles of Ca = $\frac{20}{40}$ = 0.5 *mol*

Mass of Br = 80 *g*, moles of Br = $\frac{80}{80}$ = 1 *mol*

Step 2 : Find the empirical formula
For making the elements in a simple whole number ratio, we divide by the element having least moles (by calcium).

Simple ratio of Ca = $\frac{0.5}{0.5} = 1$

Simple ratio of Br = $\frac{1}{0.5} = 2$

Empirical formula becomes = $CaBr_2$

Step 3: Find the molecular formula
Empirical formula mass = 40 + (80 × 2) = 200 *amu*
Given: Molecular mass = 200 *amu*

Using the formula,

molecular formula mass = n × empirical formula mass

200 = 200 × n
n = 1
Molecular formula = $CaBr_2$

Atom ratio determination from mass ratio

Calculate the **atom ratio** of **Ca** and **Br** in a compound with **Ca** and **Br** in **1 : 4** weight ratio.

Solution

Step 1: Converting weight ratio into percentage
Given:
Weight ratio of Ca : Br = 1 : 4.
Here, 1 : 4 ratio can be understood as 20 % of Ca and 80 % of Br.

Step 2: Find the moles of each element

Assuming mass of compound to be 100 *g*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Molar mass of Ca = 40 *g/mol*,
Molar mass of Br = 80 *g/mol*.
Mass of Ca = 20 *g*, moles of Ca = $\frac{20}{40}$ = 0.5 *mol*

Mass of Br = 80 *g*, moles of Br = $\frac{80}{80}$ = 1 *mol*

Hence atoms ratio will be 1 : 2

Step 3 : Find the empirical formula
For making the elements in a simple whole number ratio, we divide by the element having least moles (by calcium).

Simple ratio of Ca = $\frac{0.5}{0.5} = 1$

Simple ratio of Br = $\frac{1}{0.5} = 2$

Hence, the atom ratio of Ca and Br in the given compound is 1:2

Molecular formula determination from mass ratio

In an organic compound of molar mass **108 *g mol*$^{-1}$**, **C**, **H** and **N** atoms are present in **9 : 1 : 3.5** ratio by weight. The **molecular formula** can be:

(a) $C_6H_8N_2$ (b) $C_7H_{10}N$ (c) $C_5H_6N_3$ (d) $C_4H_{10}N_3$

Solution

Step 1: Convert the given ratio into simple whole numbers ratio

Given: The mass ratio of C : H : N

= 9 : 1 : 3.5 = 36 : 4 : 14

Step 2: Find the mole ratio and empirical formula

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

Dividing individual masses by respective molar masses.
Molar Ratio of C : H : N = 36/12 : 4/1 : 14/14 = 3 : 4 : 1

hence the empirical formula becomes C_3H_4N.

Step 3: Find molecular formula

For calculating Molecular Formula we need to use the formula,

Molecular formula mass = n x Empirical formula mass

$$n = \frac{Molecular\ mass}{Empirical\ Mass}$$

Given: Molecular mass = 108 *g/mol*

$$n = \frac{108}{(12*3 + 1*4 + 14*1)} = \frac{108}{54} = 2$$

Hence, the molecular Formula is $C_6H_8N_2$ (by multiplying 'n' by EF)

Empirical formula determination

Find the **empirical formula** of vanadium oxide if **2.73 *g*** of the oxide contains **1.53 *g*** of the metal. **(V = 51, O = 16)**

Solution

Step 1: Find the mass of O

Given:

Mass of metal oxide = 2.73 *g*

Mass of metal oxide = 1.53 *g*

Mass of metal oxide

= Mass of Metal + Mass of Oxygen,

2.73 = 1.53 + x

x = 1.2 *g* of Oxygen

Step 2: Find the molecular formula

Let the empirical formula be V_2O_x

$Moles = \dfrac{Given\ mass}{Molar\ mass}$

So, we can say that

(Mass of V / Mass of O)

$= \dfrac{1.53}{1.2} = \dfrac{(2 \times \text{Molar mass of Vanadium})}{(x \times \text{Molar mass of Oxygen})}$

Hence 'x' comes out to be approximately 5. So formula becomes V_2O_5.

Molecular formula determination from the mass composition and vapour density

A gaseous compound of nitrogen and hydrogen contains **12.5 %** (by mass) of hydrogen. The density of the compound relative to hydrogen is **16**. The **molecular formula** of the compound is:

(a) NH_2 (b) N_3H (c) NH_3 (d) N_2H_4

Solution

Step 1: Find the molar mass of the compound

Given:

Vapour density = 16.

We know that $Vapour\ Density = \dfrac{Molecular\ Mass}{2}$

Hence Molar Mass of compound comes out to be 32 *g*.

Step 2: Find mass ratio of elements

Percentage of Hydrogen = 12.5%.

Percentage of Nitrogen = (100 - 12.5)%

= 87.5%.

Weight ratio of N : H = 87.5 : 12.5 = 7 : 1

Step 3: Find the empirical formula

$Moles = \dfrac{Given\ mass}{Molar\ mass}$

Molar ratio is N : H = (7/14) : (1/1) = 1 : 2.

Hence, the empirical formula becomes NH_2.

Step 4: Find the molecular formula

Molecular formula (MF)

= n × Empirical formula (EF)

n = 32/16 = 2

Hence, the molecular Formula becomes N_2H_4

Empirical formula determination from mass percentage

The **simplest formula** of a compound containing **50 %** by mass of element **X** (at. wt. 10) and **50 %** by mass of element **Y** (atomic weight 20) is:

(a) XY_2 (b) X_2Y (c) X_2Y_2 (d) XY_3

Solution

Step 1: Finding the number of moles in 1 *g* of the sample

Given:

50% by mass of both X and Y

Molar mass of X = 10 *g/mol*.

Molar mass of Y = 20 *g/mol*.

Assuming that we have 1 *g* of compound.

So we have 0.5 *g* of X and 0.5 *g* of Y.

Now find the number of moles.

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

X = 0.5 / 10 = 0.05 and Y = 0.5 / 20 = 0.025.

Step 2 Find the simple whole number ratio

So, X : Y = 0.05 : 0.025 = 2 : 1

X = 2 and Y = 1.

Hence, the formula is X_2Y

Water of hydration determination

1.763 *g* of **hydrated $BaCl_2$** was heated to dryness. The **anhydrous salt** that remained was **1.505 *g***. What is the **formula of the hydrate**?

Solution

Step 1: Find the mass of water

Given :

Mass of hydrated $BaCl_2$ ($BaCl_2 \times H_2O$) = 1.763 *g*

Mass of anhydrous salt ($BaCl_2$) = 1.505 *g*

hence, mass of water = 1.763 - 1.505 = 0.258 *g*

On heating following reaction occurs:

$$BaCl_2 \times H_2O \longrightarrow BaCl_2 + xH_2O$$

We need to find 'x'.

Step 2: Find the molecular formula

Molar mass of $BaCl_2$ = 137 + 71 = 208 *g/mol*

Molar mass of H_2O = 18 *g/mol*

$$Moles = \frac{Given\ mass}{Molar\ mass}$$

$$\frac{\text{Mass of } BaCl_2}{\text{Mass of } H_2O} = \frac{1.505}{0.258}$$

$$= \frac{1 \times \text{molar mass of } BaCl_2}{x \times \text{molar mass of } H_2O} = \frac{1 \times (137+71)}{18x}$$

x = 2

Molecular formula of hydrate is $BaCl_2.2H_2O$

CHEMISTRY

MOLE CONCEPT

BEGINNING WITH STOICHIOMETRY

What you already know

- Interconversions: Mole, particle, mass, and volume
- Molar volume at STP

What you will learn

- Stoichiometric coefficients and ratios
- Introduction to stoichiometry
- Stoichiometry using moles, particles, mass, and volume interconversions
- Stoichiometry of sequential reactions
- Stoichiometry of parallel reaction

BOARDS

Balanced Chemical Equation

Definition:

A chemical equation in which the number of atoms of each and every element on both the product and the reactant sides is the same is a balanced chemical equation. Some of the examples are as follows:

$N_2 + 3H_2 \rightarrow 2NH_3$ $\quad$ $CO_2 + C \rightarrow 2CO$ $\quad$ $CaCO_3 \rightarrow CaO + CO_2$

Generalised Balanced Reaction Form

$aA + bB \rightarrow cC + dD$

A **reference to explain the general relations** between the reacting **species (A, B, C, D)**.

$N_2 + 3H_2 \rightarrow 2NH_3$

Applicable to all reactions. In this reaction, we can say that **we have no D**.

BOARDS MAIN

Stoichiometric Coefficients and Ratios

Stoichiometric coefficients

- Numbers before the reactants and the products in a balanced chemical reaction.
  For the reaction $\mathbf{N_2 + 3H_2 \rightarrow 2NH_3}$
  1, 3, and **2** are the respective stoichiometric coefficients of N_2, H_2, and NH_3.
- Stoichiometric coefficients are the mole ratio in which the reactants are consumed and the products get formed.
  For the reaction: $\mathbf{2H_2 + O_2 \rightarrow 2H_2O}$
  It means that when $\mathbf{H_2}$ and $\mathbf{O_2}$ combine to form $\mathbf{H_2O}$, it happens in a **1 : ½ : 1** ratio that can also be written as **2 : 1 : 2**.

Stoichiometric ratios

For the reaction: $\mathbf{N_2 + 3H_2 \rightarrow 2NH_3}$,

$$\frac{n_{N_2\ \text{consumed}}}{1} = \frac{n_{H_2\ \text{consumed}}}{3} = \frac{n_{NH_3\ \text{produced}}}{2}$$

$$n_{H_2\ \textbf{consumed}} = 3 \times n_{N_2\ \text{consumed}}$$

$$\frac{1}{3} = \frac{n_{N_2\ \text{consumed}}}{n_{H_2\ \text{consumed}}}$$

$$n_{NH_3\ \textbf{produced}} = 2 \times n_{N_2\ \text{consumed}}$$

$$\frac{1}{2} = \frac{n_{N_2\ \text{consumed}}}{n_{NH_3\ \text{produced}}}$$

$$2 \times n_{H_2\ \text{consumed}} = 3 \times n_{NH_3\ \text{produced}}$$

$$\frac{3}{2} = \frac{n_{H_2\ \text{consumed}}}{n_{NH_3\ \text{produced}}}$$

For the reaction: $\mathbf{N_2 + 3H_2 \rightarrow 2NH_3}$,

$1 + 3 \neq 2$:

Stoichiometric coefficients of reactants and products are NOT necessarily additive.

Numbers, not amount or molecules:
'1, 3, 2' for the mentioned reaction are NOT the $N_{molecules}$ or n_{moles}. Instead, they are the ratios of their number of moles (consumed for reactants and produced for products).

Stoichiometry

Calculations based on the **quantitative** relationship **(mole/mass/volume/number of molecules)** between the **reactants** and the **products** are referred to as **stoichiometry**.

Mass ⇌ moles ⇌ number of molecules

Or

Calculation of the consumed and the produced quantities of substances involved in the reaction.

Meaning of stoichiometry

It is about finding the amount of reactants consumed and products produced.

Alternate explanation

It gives the relation between the amount of substances that react together in a particular reaction and the amount of the products formed.

What we know about this reaction: $N_2 + 3H_2 \rightarrow 2NH_3$,

$$n_{H_2\ consumed} = 3 \times n_{N_2\ consumed}$$

$$\frac{n_{N_2\ consumed}}{1} = \frac{n_{H_2\ consumed}}{3}$$

$$n_{NH_3\ produced} = 2 \times n_{N_2\ consumed}$$

$$\frac{n_{NH_3\ produced}}{2} = \frac{n_{N_2\ consumed}}{1}$$

$$3 \times n_{NH_3\ produced} = 2 \times n_{H_2\ consumed}$$

$$\frac{n_{NH_3\ produced}}{2} = \frac{n_{H_2\ consumed}}{3}$$

$$\frac{n_{N_2\ consumed}}{1} = \frac{n_{H_2\ consumed}}{3} = \frac{n_{NH_3\ produced}}{2}$$

Similarly, for a general reaction: $aA + bB \rightarrow cC + dD$,

$$\frac{n_{A\ consumed}}{a} = \frac{n_{B\ consumed}}{b} = \frac{n_{C\ produced}}{c} = \frac{n_{D\ produced}}{d}$$

- Stoichiometric ratios represent the ratios of $n_{consumed}$ or $n_{produced}$ of the respective species and **NOT** of $n_{initial}$ or n_{final}.
- When the initial moles of the reactants are in stoichiometric ratio, they will be completely consumed, provided that the reaction goes to completion.

$N_2 + 3H_2 \rightarrow 2NH_3$,

If $\dfrac{n_{N_2\ initial}}{n_{H_2 initial}} = \dfrac{1}{3}$

Initial amounts of both N_2 and H_2 will be completely consumed.

So,

$n_{N_2\ initial} = n_{N_2\ consumed}$

$n_{H_2\ initial} = n_{H_2\ consumed}$

Which means,

$$\frac{n_{N_2\ initial}}{1} = \frac{n_{H_2\ initial}}{3} = \frac{n_{NH_3\ produced}}{2}$$

$aA + bB \rightarrow cC + dD$,

If $\dfrac{n_{A\ initial}}{n_{B\ initial}} = \dfrac{a}{b}$

Initial amounts of both A and B will be completely consumed.

$n_{A\ initial} = n_{A\ consumed}$

$n_{B\ initial} = n_{B\ consumed}$

$$\frac{n_{A\ inital}}{a} = \frac{n_{B\ initial}}{b} = \frac{n_{C\ produced}}{c} = \frac{n_{D\ produced}}{d}$$

Relating $n_{product}$ formed with $n_{reactants}$ consumed

When 7 moles of N_2 react with the required amount of H_2, what is the number of moles of NH_3 produced?

Solution

Step 1:
Balance the chemical equation.

$N_2 + 3H_2 \rightarrow 2NH_3$

Step 2:
Find the moles of N_2 consumed.

N_2 is getting consumed completely.

Therefore,

Moles of N_2 consumed = initial n_{N_2} = 7 *mol*

Step 3:
Find moles of NH_3 produced.

$$\frac{\text{Moles of } N_2 \text{ consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } NH_3 \text{ produced}}{\text{Stoic.Co-eff.}}$$

i.e., $\frac{7}{1} = \frac{\text{Moles of } NH_3 \text{ produced}}{2}$

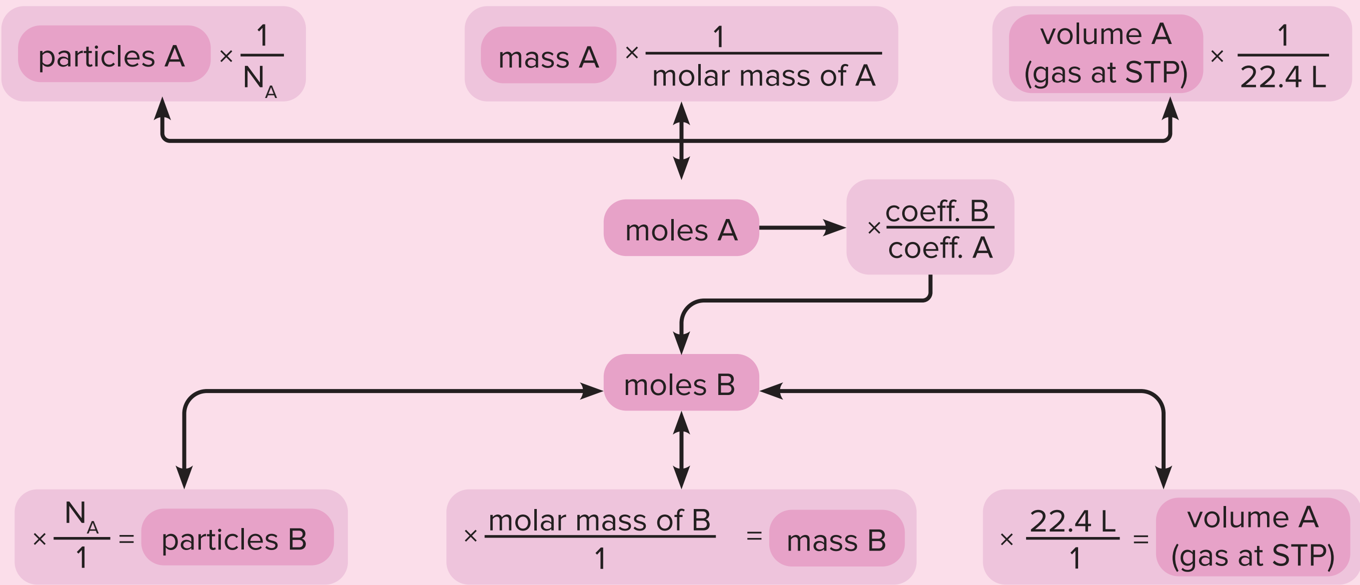
So, moles of NH_3 produced = 14 *mol.*

BOARDS MAIN ADVANCED

Interconversions + Stoichiometry

When the amount of reaction participants are known in terms of the mass or the volume, we should convert them into moles.

Or vice-versa, when the amounts are required to be found in terms of the mass or the volume.



Finding the moles of the reactant consumed

To produce 3 moles of NH_3, how many moles of N_2 should be consumed?

Solution

Step 1:

Balanced chemical equation.

$N_2 + 3H_2 \rightarrow 2NH_3$

Step 2:

Find moles of N_2 consumed.

$$\frac{\text{Moles of } N_2 \text{ consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } NH_3 \text{ produced}}{\text{Stoic.Co-eff.}}$$

$$\frac{\text{Moles of } N_2 \text{ consumed}}{1} = \frac{\text{Moles of } NH_3 \text{ produced}}{2}$$

$$\therefore \text{Moles of } N_2 \text{ consumed} = \frac{3}{2} \times 1$$

$$= 1.5\ mol$$

Finding the number of moles of the reactant consumed

On the reaction of 2 moles of N_2 with 6 moles of H_2, what is the mass of NH_3 produced?

Solution

Step 1:

Balance the chemical equation.

$N_2 + 3H_2 \rightarrow 2NH_3$

Step 2:

Find the moles of N_2 and H_2 consumed.

As the initial moles are in stoichiometric ratio,

Moles of N_2 consumed = initial n_{N_2} = 2 mol

Moles of H_2 consumed = initial n_{H_2} = 6 mol

Step 3:

Find the moles of NH_3 produced.

$$\frac{\text{Moles of } N_2 \text{ consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } H_2 \text{ consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } NH_3 \text{ produced}}{\text{Stoic.Co-eff.}}$$

i.e., $\frac{2}{1} = \frac{6}{3} = \frac{\text{Moles of } NH_3 \text{ produced}}{2}$

So, moles of NH_3 produced = 4 mol

Step 4:

Find mass of NH_3 produced.

Mass of NH_3 produced = n × (Molar Mass)

= 4 × 17 = 68 g

Finding the volume of product produced

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

On the complete decomposition of 245 g of $KClO_3$, the volume of O_2 gas produced at STP is as follows:

(Atomic mass of K = 39 and Cl = 35.5)

Solution

Step 1:

Balance the chemical equation

$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

Step 2:

Find the moles of $KClO_3$ Consumed

$$\text{Moles of } KClO_3 \text{ consumed} = \frac{\text{Given weight}}{\text{Molar mass}}$$

$$\text{Moles} = \frac{245\ g}{122.5\ g/mol} = 2\ mol$$

Step 3:

Find the moles of O_2 produced.

$$\frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoic. Coeff.}} = \frac{\text{Moles of } O_2 \text{ produced}}{\text{Stoic. Coeff.}}$$

$$\text{i.e., } \frac{\text{Moles of } KClO_3 \text{ Consumed}}{2} = \frac{\text{Moles of } O_2 \text{ produced}}{3}$$

$$\therefore \text{Moles of } O_2 \text{ produced} = \frac{2}{2} \times 3 = 3\ mol$$

Step 4:

Find the volume of O_2 produced at STP.

Volume of O_2 produced = n × 22.4 L = 3 × 22.4 L = 67.2 L

Finding the mass of the product produced

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

In this reaction, if 32 moles of O_2 had been produced, then the mass of KCl formed is as follows:

Solution

Step 1:

Balance the chemical Equation.

$2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

Step 2:

Find the moles of KCl produced.

$$\frac{\text{Moles of KCl produced}}{\text{Stoic.Coeff.}}$$

$$= \frac{\text{Moles of } O_2 \text{ produced}}{\text{Stoic. Coeff.}}$$

i.e., $\frac{\text{Moles of KCl formed}}{2} = \frac{\text{Moles of } O_2 \text{ produced}}{3}$

$\therefore$ Moles of KCl produced = $\frac{32}{3} \times 2 = 21.333\ mol$

Step 3:

Find the mass of KCl produced.

Mass of KCl Produced = n × (Molar mass)

Mass = 21.333 × 74.5 = 1589.3 g

Finding the mass of product formed using stoichiometry

If H_2, reacting with N_2 to give 85 g of NH_3, would rather have reacted with sufficient or excess amounts of O_2 to form H_2O, find the mass of H_2O produced.

Solution

Step 1:

Balance the Chemical Equations

$3H_2 + N_2 \rightarrow 2NH_3$

$2H_2 + O_2 \rightarrow 2H_2O$

Step 2:

Find Moles of NH_3 produced

Moles of NH_3 produced = $\frac{\text{Given weight}}{\text{Molar mass}}$

$= \frac{85\ g}{17\ g/mol} = 5\ mol$

Step 3:

Find the Moles of H_2 consumed

$\frac{\text{Moles of } H_2 \text{ consumed}}{\text{Stoic. Coeff.}} = \frac{\text{Moles of } NH_3 \text{ produced}}{\text{Stoic. Coeff}}$

i.e., $\frac{\text{Moles of } H_2 \text{ consumed}}{3} = \frac{\text{Moles of } NH_3 \text{ produced}}{2}$

Moles of H_2 consumed = $3 \times \frac{5}{2} = 7.5\ mol$

Step 4:

Find the Moles of H_2O produced

$\frac{\text{Moles of } H_2 \text{ consumed}}{\text{Stoic. Coeff.}} = \frac{\text{Moles of } H_2O \text{ produced}}{\text{Stoic. Coeff.}}$

i.e., $\frac{\text{Moles of } H_2 \text{ consumed}}{2} = \frac{\text{Moles of } H_2O \text{ produced}}{2}$

i.e., moles of H_2O produced = moles of H_2 consumed = 7.5

Hence, weight of H_2O produced = 7.5 × 18 = 135 g.

MAIN

Combination of Reactions - Sequential Reactions

ADVANCED

Sequential reactions consist of linked reactions in which the product of the first reaction becomes the substrate/reactant of the second reaction. For example,

$$A + 2B \rightarrow 3C + 4D$$

$$5C + 6E \rightarrow 7F + 8G$$

Alternative explanation 1

Sequential reactions consist of linked reactions in which the product of the first reaction becomes the substrate/reactant of the second reaction.

Alternative explanation 2

The amount of C available for the reaction with E in the second reaction must be produced from the first reaction.

Finding the number of moles in sequential reactions

$2A + 3B \rightarrow 4C + 5D$

$3C + 4E \rightarrow 5G + 6F$

Starting with **7 moles of A** and **excess of B and E,** find the **moles of G** formed.

Solution

Step 1:
Find the moles of C produced.

$$\frac{\text{Moles of C produced}}{\text{Stoic.Coeff.}} = \frac{\text{Moles of A consumed}}{\text{Stoic.Coeff.}}$$

i.e., $$\frac{\text{Moles of C produced}}{4} = \frac{\text{Moles of A consumed}}{2}$$

$$\text{Moles of C produced} = 4 \times \frac{7}{2} = 14\ mol$$

Step 2:
Find the moles of G produced.

$$\frac{\text{Moles of G produced}}{\text{Stoic.Coeff.}} = \frac{\text{Moles of C consumed}}{\text{Stoic.Coeff.}}$$

i.e., $$\frac{\text{Moles of G produced}}{5} = \frac{\text{Moles of C consumed}}{3}$$

$$\text{Moles of D produced} = 5 \times \frac{14}{3} = 23.3\ mol$$

Shortcut

$$\frac{\text{Moles of G produced}}{\text{Moles of C consumed}} = \frac{\text{Stoic.Coeff. G}}{\text{Stoic.Coeff. C}} = \frac{5}{3} \quad \text{(i)}$$

$$\frac{\text{Moles of C produced}}{\text{Moles of A consumed}} = \frac{\text{Stoic.Coeff. C}}{\text{Stoic.Coeff. A}} = \frac{4}{2} \quad \text{(ii)}$$

Multiply equation (i) and (ii)

$$\frac{\text{Moles of G produced}}{\text{Moles of C consumed}} \times \frac{\text{Moles of C produced}}{\text{Moles of A consumed}} = \frac{\text{Stoic.Coeff. G}}{\text{Stoic.Coeff. C}} \times \frac{\text{Stoic.Coeff. C}}{\text{Stoic.Coeff. A}}$$

$$\text{Moles of G produced} = 7 \times \frac{5}{3} \times \frac{4}{2} = 23.3\ mol$$

Finding the number of moles in sequential reactions

$2A + 3B \rightarrow 4C + 5D$ $\quad$ $7C + 8E \rightarrow 9F + 10G$ $\quad$ $5F + 6H \rightarrow 7I + 8J$

Starting with **3 moles of A** and **excess of B, E and H.** find the **moles of I formed.**

Solution

Step 1:

Find the number of moles of C produced.

$$\frac{\text{Moles of C produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of A consumed}}{\text{Stoic.Co-eff.}}$$

i.e., $\frac{\text{Moles of C produced}}{4}$

$= \frac{\text{Moles of A consumed}}{2}$

$$\text{Moles of C produced} = 4 \times \frac{3}{2} = 6\ mol$$

Step 2:

Find the number of moles of F produced.

$$\frac{\text{Moles of F produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of C consumed}}{\text{Stoic.Co-eff.}}$$

i.e., $\frac{\text{Moles of F produced}}{9}$

$= \frac{\text{Moles of C consumed}}{7}$

$$\text{Moles of F produced} = 9 \times \frac{6}{7} = 7.71\ mol$$

Step 3:

Find the moles of I produced.

$$\frac{\text{Moles of I produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of F consumed}}{\text{Stoic.Co-eff.}}$$

i.e., $\frac{\text{Moles of I produced}}{7} = \frac{\text{Moles of F consumed}}{5}$

$$\text{Moles of I produced} = 7 \times \frac{7.71}{5} = 10.8\ mol$$

Shortcut

$$\frac{\text{Moles of I produced}}{\text{Moles of F consumed}} \times \frac{\text{Moles of F produced}}{\text{Moles of C consumed}} \times \frac{\text{Moles of C produced}}{\text{Moles of A consumed}}$$

$$= \frac{\text{Stoic.Co-eff. I}}{\text{Stoic.Co-eff. F}} \times \frac{\text{Stoic.Co-eff. F}}{\text{Stoic.Co-eff. C}} \times \frac{\text{Stoic.Co-eff. C}}{\text{Stoic.Co-eff. A}}$$

$$\Rightarrow \frac{\text{Moles of I produced}}{\text{Moles of A consumed}} = \frac{7}{5} \quad \frac{9}{7} \times \frac{4}{2}$$

$$\text{Hence, moles of I produced} = 3 \times \frac{7}{5} \times \frac{9}{7} \times \frac{4}{2} = 10.8$$

Finding the number of moles in sequential reactions

$2A + 3B \rightarrow 4C + 5D$ $\quad$ $6E + 7C \rightarrow 8F + 9D$ $\quad$ $10F + 11G \rightarrow 12H + 13D$

Starting with **5 moles of A** and **excess of B, E, and G,** find the **moles of D** formed.

Solution

Step 1:

Find the moles of D produced in the reaction (i).

$2A + 3B \rightarrow 4C + 5D \quad \text{(i)}$

$$\frac{\text{Moles of D produced}}{\text{Stoic.Coeff.}} = \frac{\text{Moles of A consumed}}{\text{Stoic.Coeff.}}$$

i.e., $\dfrac{\text{Moles of D produced}}{5}$

$$= \frac{\text{Moles of A consumed}}{2}$$

$$\text{Moles of D produced} = 5 \times \frac{5}{2} = 12.50\ mol$$

Step 2:

Find the moles of D produced in the reaction (ii).

$6E + 7C \rightarrow 8F + 9D \quad \text{(ii)}$

$$\frac{\text{Moles of D produced in reaction (ii)}}{\text{Moles of C consumed in reaction (ii)}} \times \frac{\text{Moles of C produced in reaction (i)}}{\text{Moles of A consumed in reaction (i)}}$$

$$= \frac{\text{Stoic.Coeff. of D in reaction (ii)}}{\text{Stoic.Coeff. of C in reaction (ii)}} \times \frac{\text{Stoic.Coeff. of C in reaction (i)}}{\text{Stoic.Coeff. of A in reaction (i)}}$$

i.e., $\dfrac{\text{Moles of D produced in reaction (ii)}}{\text{Moles of A consumed in reaction (i)}}$

$$= \frac{9}{7} \times \frac{4}{2}$$

$$\text{Moles of D produced} = 5 \times \frac{9}{7} \times \frac{4}{2}$$

$$= 12.85\ mol$$

Step 3:

Find the moles of D produced in the reaction (iii).

$10F + 11G \rightarrow 12H + 13D \quad \text{(iii)}$

$$\frac{\text{Moles of D produced in reaction (iii)}}{\text{Moles of F consumed in reaction (iii)}} \times \frac{\text{Moles of F produced in reaction (ii)}}{\text{Moles of C consumed in reaction (ii)}}$$

$$\times \frac{\text{Moles of C produced in reaction (i)}}{\text{Moles of A consumed in reaction (i)}} = \frac{\text{Stoic.Coeff. of D in reaction (iii)}}{\text{Stoic.Coeff. of F in reaction (iii)}}$$

$$\times \frac{\text{Stoic.Coeff. of F in reaction (ii)}}{\text{Stoic.Coeff. of C in reaction (ii)}} \times \frac{\text{Stoic.Coeff. of C in reaction (i)}}{\text{Stoic.Coeff. of A in reaction (i)}}$$

$$\frac{\text{Moles of C produced in reaction (iii)}}{\text{Moles of A consumed in reaction (i)}} = \frac{13}{10} \times \frac{8}{7} \times \frac{4}{2}$$

$$\text{Moles of D produced} = 5 \times \frac{13}{10} \times \frac{8}{7} \times \frac{4}{2} = 14.86\ mol$$

Total Moles of D produced = 12.5 + 12.85 + 14.86 = 40.21 *mol.*

MAIN

Combination of Reactions - Parallel Reactions

ADVANCED

The reaction set in which a substance reacts or decomposes in more than one way are known as parallel reactions.

Example

$A \rightarrow 2B + 3C$

$4A \rightarrow 5D + 6E$

Alternative explanation

In these reactions, A is involved in two independent reactions at the same time. These are known as parallel reactions.

In parallel reactions, the ratio of $n_{\text{reactants consumed}}$ of the common element in each of the reactions is not equal to the ratio of its stoichiometric coefficients in the respective reactions.

Finding the number of moles in parallel reactions

$KClO_3$ decomposes by the following two parallel reactions:

(i) $2KClO_3 \rightarrow 2KCl + 3O_2$

(ii) $4KClO_3 \rightarrow 3KClO_4 + KCl$

If **3 moles of O_2** and **1 mol of $KClO_4$** is produced along with other products, determine **initial moles of $KClO_3$.**

Solution

Step 1:

Find the moles of $KClO_3$ consumed in the reaction.

$2KClO_3 \rightarrow 2KCl + 3O_2$

$$\frac{\text{Moles of } O_2 \text{ produced}}{\text{Stoic.Coeff.}} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoic.Coeff.}}$$

$$\text{i.e.,} \frac{\text{Moles of } O_2 \text{ produced}}{3} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{2}$$

Moles of $KClO_3$ consumed $= 2 \times \frac{3}{3} = 2\ mol$

Step 2:

Find the moles of $KClO_3$ consumed in the reaction.

$4KClO_3 \rightarrow 3KClO_4 + KCl$

$$\frac{\text{Moles of } KClO_4 \text{ produced}}{\text{Stoic.Coeff.}} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoic.Coeff.}}$$

i.e., $$\frac{\text{Moles of } KClO_4 \text{ produced}}{3} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{4}$$

Moles of $KClO_3$ consumed $= 4 \times \frac{1}{3}$

$= 1.3\ mol$

Total moles of $KClO_3$ consumed $= 2 + 1.3\ mol$

$= 3.3\ mol$

C H E M I S T R Y

MOLE CONCEPT

MORE INTO STOICHIOMETRY

What you already know

- Mole
- Avogadro's number
- Mole particle conversion
- Subatomic particles

What you will learn

- Percentage purity
- Percentage yield of single-step reaction
- Limiting reagent

MAIN

Percentage Purity

$$\% \text{ purity} = \frac{\text{Actual amount of desired species in the sample}}{\text{Total amount of the sample}} \times 100$$

- For example, 100 *g* of 80 % pure NaCl implies that the sample contains 80 *g* of NaCl and 20 *g* of unwanted substance also known as impurities.

Percentage purity determination from stoichiometry

A **10 *g* sample of $KClO_3$**, gave on complete decomposition, **2.24 *L* of oxygen** at STP. What is the **percentage purity** of the sample of **potassium chlorate**?

Solution

Step 1:
Balanced chemical equation

$2KClO_3 \rightarrow 2KCl + 3O_2$

Step 2:
Find moles of O_2 produced

$$\text{Moles of } O_2 \text{ produced} = \frac{\text{Volume Of } O_2 \text{ in litres}}{22.4\ L}$$

$$= \frac{2.24\ L}{22.4\ L}$$

$$= 0.1\ mol$$

Step 3:

Find moles of $KClO_3$ consumed

$$\frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoichiometric coefficient}} = \frac{\text{Moles of } O_2 \text{ produced}}{\text{Stoichiometric coefficient}}$$

i.e. $$\frac{\text{Moles of } KClO_3 \text{ consumed}}{2} = \frac{\text{Moles of } O_2 \text{ produced}}{3}$$

So, Moles of $KClO_3$ consumed = 0.0666 *mol*

Step 4:

Find mass of $KClO_3$ consumed

Mass of $KClO_3$ consumed

= Moles × Molar mass

= 0.0666 × 122.5

= 8.16 *g*

Step 5:

Find percentage purity of $KClO_3$

Percentage purity =

$$\frac{\textit{Actual amount of desired species in the sample}}{\textit{Total amount of the sample}} \times 100$$

$= \frac{8.16}{10} \times 100$

= 81.6%

Finding mass of product formed using percentage purity

Calculate the **weight of lime (CaO)** obtained by heating **200 *kg* of 95% pure limestone ($CaCO_3$)**.

Solution

Step 1:

Balanced chemical equation

$CaCO_3 \rightarrow CaO + CO_2$

Step 2:

Finding actual mass of $CaCO_3$

= 95/100 × 200 *kg* = 190 *kg*

Step 3:

Find mass of CaO produced

Using stoichiometry, we can say 100 *g* of $CaCO_3$ gives 56 *g* of CaO.

So, by unitary method 190 *kg* $CaCO_3$ gives

$\frac{56}{100} \times 190 \times 10^3$

$= 106.4 \times 10^3 g$

= 106.4 *kg* of CaO

BOARDS MAIN

Percentage yield in a single-step reaction

Percentage yield is the ratio of actual yield to the theoretical yield, where actual yield is the yield obtained by doing an experiment.

Formula for percentage yield is as follows:

$$\textit{Percentage yield} = \frac{\textit{Actual yield (or Experimental yield)}}{\textit{Theoretical yield}} \times 100$$

Finding percentage yield of reaction

If **11.5 *g* of Na** reacts with **9.0 *g* of H_2O, 0.40 *g* of H_2** is formed. What is the **percentage yield** of the reaction?

Solution

Step 1:

Balanced chemical equation

$Na + H_2O \rightarrow NaOH + \frac{1}{2} H_2$

Step 2:

Find theoretical mass of H_2 produced

Using stoichiometry,
we can say 23 *g* of Na reacts with 18 *g* of H_2O to give 1*g* of H_2.

So by unitary method
11.5 *g* of Na reacts with 9 *g* of H_2O to give

$\frac{1}{23} \times 11.5 = 0.5$ *g* of H_2.

Step 3:

Find percentage yield

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{0.4}{0.5} \times 100 = 80\%$$

Product yield determination from an impure sample

Calculate the **mass of CO_2** produced by heating **40 *g* of 20% pure $CaCO_3$**.

Solution

Step 1:

Balanced chemical equation

$CaCO_3 \rightarrow CaO + CO_2$

Step 2:

Find moles of $CaCO_3$ consumed

Weight of $CaCO_3$ = Weight of sample × Percentage purity = $40\ g \times \frac{20}{100} = 8\ g$

Moles of $CaCO_3$ = $\frac{\text{Given weight}}{\text{Molar mass}} = \frac{8\ g}{100\ g/mol} = 0.08\ mol$

Step 3:

Find moles of CO_2 produced

$$\frac{\text{Moles of } CaCO_3 \text{ consumed}}{\text{Stoichiometric coefficient}} = \frac{\text{Moles of } CO_2 \text{ produced}}{\text{Stoichiometric coefficient}} \Rightarrow \frac{0.08}{1} = \frac{\text{Moles of } CO_2 \text{ produced}}{1}$$

So, moles of CO_2 produced = 0.08 *mol*

Step 4:

Find mass of CO_2 produced

Mass of CO_2 produced = Moles × Molar mass = $0.08 \times 44 = 3.52\ g$

Percentage purity determination from stoichiometry

20 *g* sample of **magnesium carbonate** decomposes on heating to give **carbon dioxide** and **8.0 *g* magnesium oxide**. What will be the **percentage purity of magnesium carbonate** in the sample?

(Molar mass of Mg = 24 *g/mol*)

Solution

Step 1:

Balanced chemical equation

$MgCO_3 \rightarrow MgO + CO_2$

Step 2:

Find moles of MgO produced

Moles of MgO = $\frac{\text{Given weight}}{\text{molar mass}}$

$= \frac{8}{40} = 0.2\ mol$

Step 3:

Find moles of $MgCO_3$ consumed

$$\frac{\text{Moles of } MgCO_3 \text{ consumed}}{1} = \frac{0.2}{1}$$

Moles of $MgCO_3$ = 0.2 *mol*

Step 4:

Find mass of $MgCO_3$ consumed

Mass of $MgCO_3$ consumed

= Moles × Molar mass = 0.2 × 84 = 16.8 *g*

Step 5:

Find percentage purity of $MgCO_3$

$$\text{Percentage purity} = \frac{16.8}{20} \times 100 = 84\%$$

Product yield determination from percentage yield

The **percent yield** for the following reaction carried out in carbon tetrachloride (CCl_4) solution is 70%. **$Br_2 + Cl_2 \rightarrow 2BrCl$**

How many moles of **BrCl would be formed** from the reaction of **0.05 *mol* Br_2** and **0.05 *mol* Cl_2**?

Solution

Step 1:

Balanced chemical equation

$Br_2 + Cl_2 \rightarrow 2BrCl$

Step 2:

Find theoretical yield of BrCl

$$\frac{\text{Moles of } Br_2 \text{ consumed}}{\text{Stoich. Coeff}} = \frac{\text{Moles of } Cl_2 \text{ consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of BrCl produced}}{\text{Stoic.Co-eff.}}$$

$$\frac{0.05\ mol}{1} = \frac{0.05\ mol}{1} = \frac{\text{Moles of BrCl produced}}{2}$$

Moles of BrCl produced theoretically = 0.1 *mol*

Step 3:

Find actual yield of BrCl

$$\text{Moles of BrCl produced} = 0.1 \times \frac{70}{100} = 0.07\ mol$$

Percentage yield determination in single-step reaction

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

32 *g* of methane on combustion in an open atmosphere produces **72 *g* of CO_2**. Calculate the **percentage yield** of the reaction.

Solution

Step 1:
Balanced chemical equation

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Step 2:
Find moles of CH_4 consumed

$$\text{Moles of } CH_4 = \frac{\text{Given weight}}{\text{Molar mass}}$$

$$= \frac{32}{16}$$

$$= 2\ mol$$

Step 3:
Theoretical yield of CO_2

$$\frac{\text{Moles of } CO_2 \text{ produced}}{1} = \frac{2}{1}$$

$$\text{Moles of } CO_2 \text{ produced} = 2$$

Step 4:
Find mass of CO_2 produced

$$\text{Mass of } CO_2 \text{ produced} = \text{Moles} \times \text{Molar mass}$$
$$= 2 \times 44$$
$$= 88\ g$$

Step 5:
Find percentage yield of the reaction

$$\text{Percentage yield of } CO_2 \text{ produced} = 100 \times \frac{72}{88}$$
$$= 81.8\ \%$$

ADVANCED MAIN

Percentage yield in reactions that occur in more than one step

Reactions that occur in more than one step are known as multistep reactions. If a reaction is a sequential **multistep** process, then the overall **yield** is the product of the **yields** of each step. So for example, if a **synthesis** has two steps, each of **yield** 50% then the overall **yield** is 50% × 50% = 25%.

Experimental yield determination in a multistep reaction

$2A + 3B \rightarrow 4C + 5D$ yield = 70%

$3C + 4E \rightarrow 5G + 6F$ yield = 60%

Starting with **7 moles of A and excess of B and E**, find the **moles of G** formed.

Solution

Step 1:

Theoretical yield of C in reaction

$2A + 3B \rightarrow 4C + 5D$

$$\frac{\text{Moles of A consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of C produced}}{\text{Stoic.Co-eff.}}$$

$$\text{i.e. } \frac{7}{2} = \frac{\text{Moles of C produced}}{4}$$

$$\text{Moles of C produced} = 4 \times \frac{7}{2} = 14\ mol$$

Step 2:

Experimental yield of C in reaction

$2A + 3B \rightarrow 4C + 5D$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Actual yield of C} = 14 \times \frac{70}{100}$$

$$= 9.8\ mol$$

Step 3:

Theoretical yield of G in reaction

$3C + 4E \rightarrow 5G + 6F$

$$\frac{\text{Moles of C consumed}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of G produced}}{\text{Stoic.Co-eff.}}$$

$$\text{i.e. } \frac{9.8}{3} = \frac{\text{Moles of G produced}}{5}$$

$$\text{Moles of G produced} = 5 \times \frac{9.8}{3} = 16.33\ mol$$

Step 4:

Experimental yield of G in reaction

$3C + 4E \rightarrow 5G + 6F$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Actual yield of G} = 16.33 \times \frac{60}{100}$$

$$= 9.8\ mol$$

Actual yield determination in a multistep reaction

$2A + 3B \rightarrow 4C + 5D$ yield = 70%

$7C + 8E \rightarrow 9F + 10G$ yield = 50%

$5F + 6H \rightarrow 7I + 8J$ yield = 64%

Starting with 3 moles of A and excess of B, E, and H. Find the moles of I formed.

Solution

Step 1:
Find theoretical yield of C in reaction

$2A + 3B \rightarrow 4C + 5D$

$$\frac{\text{Moles of C produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of A consumed}}{\text{Stoic.Co-eff.}}$$

i.e. $$\frac{\text{Moles of C produced}}{4} = \frac{\text{Moles of A consumed}}{2}$$

$$\text{Moles of C produced} = 4 \times \frac{3}{2} = 6\ mol$$

Step 2:
Find actual yield of C in reaction

$2A + 3B \rightarrow 4C + 5D$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Actual yield of C} = 6 \times \frac{70}{100}$$

$$= 4.2\ mol$$

Step 3:
Find theoretical yield of F in reaction

$7C + 8E \rightarrow 9F + 10G$

$$\frac{\text{Moles of F produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of C consumed}}{\text{Stoic.Co-eff.}}$$

i.e. $$\frac{\text{Moles of F produced}}{9} = \frac{\text{Moles of C consumed}}{7}$$

$$\text{Moles of F produced} = 9 \times \frac{4.2}{7} = 5.4\ mol$$

Step 4:
Find actual yield of F in reaction

$7C + 8E \rightarrow 9F + 10G$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Actual yield of F} = 5.4 \times \frac{50}{100}$$

$$= 2.7\ mol$$

Step 5:
Find theoretical yield of I in reaction

$5F + 6H \rightarrow 7I + 8J$

$$\frac{\text{Moles of I produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of F consumed}}{\text{Stoic.Co-eff.}}$$

i.e. $$\frac{\text{Moles of I produced}}{7} = \frac{\text{Moles of F consumed}}{5}$$

$$\text{Moles of I produced} = 7 \times \frac{2.7}{5} = 3.78\ mol$$

Step 6:
Find actual yield of I in reaction

$5F + 6H \rightarrow 7I + 8J$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Actual yield of I} = 3.78 \times \frac{64}{100}$$

$$= 2.42\ mol$$

ADVANCED MAIN

Percentage yield in parallel reactions

The **reactions** in which a substance **reacts** or decomposes in more than one way are known as **parallel** or side **reactions**.

Determination of amount of reactant in parallel reaction using percentage yield

$KClO_3$ decomposes by two parallel reactions:

(i) $2KClO_3 \rightarrow 2KCl + 3O_2$ yield = 40%

(ii) $4KClO_3 \rightarrow 3KClO_4 + KCl$ yield = 50%

If **3 moles of O_2 and 1 mol of $KClO_4$** is produced along with other products, then determine **initial moles of $KClO_3$**.

Solution

Step 1:

Find theoretical yield of O_2 in reaction

$2KClO_3 \rightarrow 2KCl + 3O_2$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Theoretical yield of } O_2 = 3 \times \frac{100}{40}$$

$$= 7.5\ mol$$

Step 2:

Find moles of $KClO_3$ consumed in reaction

$2KClO_3 \rightarrow 2KCl + 3O_2$

$$\frac{\text{Moles of } O_2 \text{ produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoic.Co-eff.}}$$

i.e.

$$\frac{\text{Moles of } O_2 \text{ produced}}{3} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{2}$$

$$\text{Moles of } KClO_3 \text{ consumed} = 2 \times \frac{7.5}{3} = 5\ mol$$

Step 3:

Find theoretical yield of $KClO_4$ produced in reaction

$4KClO_3 \rightarrow 3KClO_4 + KCl$

$$\textit{Percentage yield} = \frac{\textit{Actual yield}}{\textit{Theoretical yield}} \times 100$$

$$\text{Theoretical yield of } KClO_4 = 1 \times \frac{100}{50}$$

$$= 2\ mol$$

Step 4:

Find moles of $KClO_3$ consumed in reaction

$4KClO_3 \rightarrow 3KClO_4 + KCl$

$$\frac{\text{Moles of } KClO_4 \text{ produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{\text{Stoic.Co-eff.}}$$

i.e.

$$\frac{\text{Moles of } KClO_4 \text{ produced}}{3} = \frac{\text{Moles of } KClO_3 \text{ consumed}}{4}$$

$$\text{Moles of } KClO_3 \text{ consumed} = 2 \times \frac{4}{3} = 2.6\ mol$$

Total moles of $KClO_3$ consumed is

$= 5 + 2.6 = 7.6\ mol$

BOARDS MAIN

Limiting Reagent and Excess Reagent

The **reactant which is consumed first** and **limits the amount of product formed** in the reaction is known as limiting reagent. If a reaction goes to completion, then limiting reagent gets consumed completely.

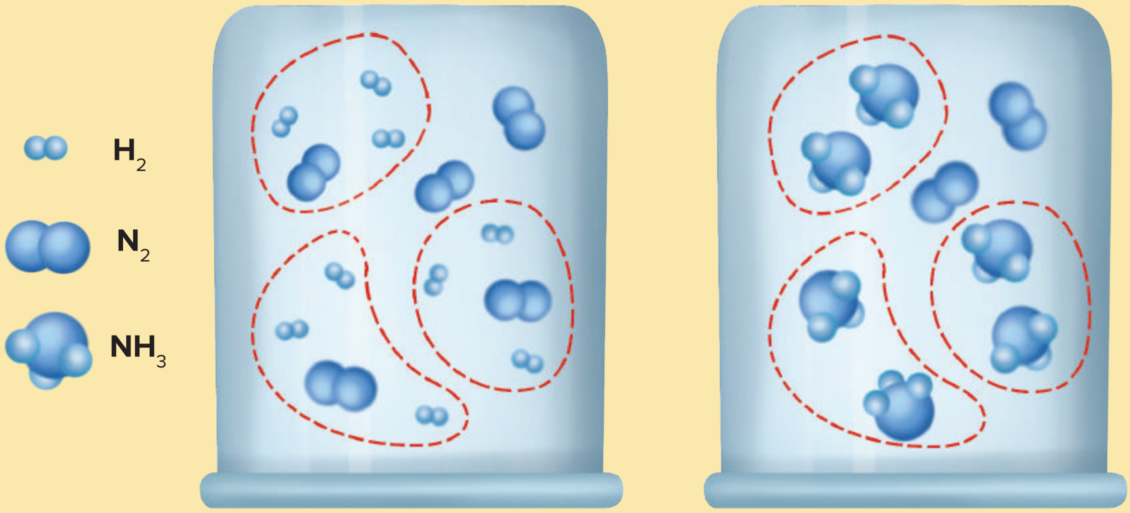
The **excess reagent** is one which **does not get consumed completely** even after the completion of reaction.

Let's understand with the following example.

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

From stoichiometry,

1 mole of N_2 reacts with 3 moles of H_2 to give 2 moles of NH_3.



Before the reaction **After the reaction**

- Let's consider 5 moles of nitrogen and 9 moles of hydrogen are present initially, 3 moles of hydrogen will react with 1 mole of nitrogen, and hence 9 moles of hydrogen will react with the 3 moles of nitrogen.
- 2 moles of nitrogen will remain unreacted. Here, the amount of ammonia formed is limited by the amount of hydrogen present.
- Hence, hydrogen is the limiting reagent and N_2 is the excess reagent.

How to determine limiting reagent (LR)

Steps:

1. Whatever data is given, weight, volume, or number of particles, first find number of moles.
2. Divide number of moles of each reagent by their respective coefficient.
3. Whichever gives smallest value in step 2 is LR and all other reagents are excess reagents.
4. In order to find moles of product and moles of excess reagent consumed, do mole analysis using moles of LR.

Limiting reagent determination

The reaction $2C + O_2 \rightarrow 2CO$ is carried out by taking **24 *g* of carbon** and **96 *g* O_2** . What are the limiting and excess reagent?

Solution

Step 1:
Find moles of C from given mass

$$\text{Moles of C} = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$= \frac{24}{12}$$

$$= 2\ mol$$

Step 2:
Find moles of O_2 from given mass

$$\text{Moles of } O_2 = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$= \frac{96}{32}$$

$$= 3\ mol$$

Step 3:
Find limiting and excess reagent

For C,

$$\frac{\text{Moles}}{\text{Stoic.Co-eff.}} = \frac{2}{2}$$

$$= 1\ mol$$

For O_2,

$$\frac{\text{Moles}}{\text{Stoic.Co-eff.}} = \frac{3}{1}$$

$$= 3\ mol$$

The ratio for carbon is less. Hence, Carbon is limiting reagent and O_2 is excess reagent.

Limiting reagent and product yield determination

For the reaction, $2P + Q \rightarrow R$, **8 mole of P** and **5 mol of Q** will produce ...... **mole of R**.

Step 1:
Find limiting reagent

For P,

$$\frac{\text{Moles}}{\text{Stoic.Co-eff.}}$$

$$= \frac{8}{2}$$

$$= 4$$

For Q,

$$\frac{\text{Moles}}{\text{Stoic.Co-eff.}}$$

$$= \frac{5}{1}$$

$$= 5$$

The ratio for P is less. Hence, P is limiting reagent.

Step 2:
Find moles of R produced

$$\frac{\text{Moles of R produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of P consumed}}{\text{Stoic.Co-eff.}}$$

i.e.

$$\frac{\text{Moles of R produced}}{1} = \frac{8}{2}$$

Moles of R produced = 4 *mol*

Limiting reagent and product yield determination

4 moles of $MgCO_3$ is reacted with **6 moles of HCl** solution. Find the volume of **CO_2 gas produced at STP**, the reaction is $MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$.

Solution

Step 1:
Finding limiting reagent

$$\text{For } MgCO_3 = \frac{\text{Moles}}{\text{Stoic.Co-eff.}} = \frac{4}{1} = 4$$

$$\text{For HCl} = \frac{6}{2} = 3$$

The ratio for HCl is less. Hence, HCl is limiting reagent .

Step 2:
Moles of CO_2 produced

$$\frac{\text{Moles of } CO_2 \text{ produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of HCl consumed}}{\text{Stoic.Co-eff.}}$$

$$\frac{\text{Moles of } CO_2 \text{ produced}}{1} = \frac{6}{2}$$

Moles of CO_2 produced = 3 *mol*

Step 3:
Finding the volume of CO_2 produced

Volume of 1 mole of gas at STP = 22.7 *L*
Volume of 3 moles of CO_2 at STP = 22.7 × 3 = 68.1 *L*

Limiting reagent and product yield determination

49 *g* of H_2SO_4 is mixed with **14.8 g of $Ca(OH)_2$**. What is the **mass of $CaSo_4$** formed?

Step 1:
Balanced chemical equation

$H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$

Step 2:
Find mole of H_2SO_4

$$\text{Moles of } H_2SO_4 = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{49}{98} = 0.5\ mol$$

Step 3:
Find mole of $Ca(OH)_2$

$$\text{Moles of } Ca(OH)_2 = \frac{\text{Given mass}}{\text{Molar mass}} = \frac{14.8}{74} = 0.2\ mol$$

Step 4:
Find limiting reagent

$$\text{For } H_2SO_4 = \frac{\text{Moles}}{\text{Stoic.Co-eff.}} = \frac{0.5}{1} = 0.5$$

$$\text{For } Ca(OH)_2 = \frac{\text{Moles}}{\text{Stoic.Co-eff.}} = \frac{0.2}{1} = 0.2$$

Here $Ca(OH)_2$ is limiting reagent.

Step 5:

Find moles of $CaSO_4$

$$\frac{\text{Moles of } CaSO_4 \text{ produced}}{\text{Stoic.Co-eff.}} = \frac{\text{Moles of } Ca(OH)_2 \text{ consumed}}{\text{Stoic.Co-eff.}}$$

i.e. $$\frac{\text{Moles of } CaSO_4 \text{ produced}}{1} = \frac{0.2}{1}$$

Moles of $CaSO_4$ produced = 0.2 *mol*

Step 6:

Find mass of $CaSO_4$

Mass of $CaSO_4$ consumed = Moles × Molar mass = 0.2 × 136 = 27.2 *g*

CHEMISTRY

MOLE CONCEPT

LIMITS OF A REACTION

What you already know

- % yield in case of single step reactions
- % yield in case of multiple step reactions
- Limiting & Excess reagent

What you will learn

- Applications of Limiting reagents in single step reactions
- Combining % yield with Limiting reagents
- Applications of Limiting reagents in multiple step reactions
- Determination of excess reagent

MAIN ADVANCED

Limiting reagent - Single step reaction

The reaction which occurs in one step is called a single step reaction.

- Limiting reagent can be determined by dividing the initial amount of the reactant (given in moles) by the stoichiometric coefficient of the reactant in a balanced chemical equation.

Example

In a given reaction:

$3A + 4B \rightarrow 5C + 6D$

The initial number of moles given are 5 moles of A and 6 moles of B.

So the limiting reagent can be determined as follows:

$$\frac{n_A}{\text{Stoic.Coeff. of A}} = \frac{5}{3} = 1.66 \qquad \frac{n_B}{\text{Stoic.Coeff. of B}} = \frac{6}{4} = 1.5$$

Where n_A and n_B are the initial number of moles of A and B.

In the above mentioned single step reaction, reactant **B will be the limiting reagent** and A will be the excess reagent because the **ratio calculated is less for B.**

1. Limiting reagent depends upon the initial amount of reactants.
2. Limiting reagent is not a property, it is just nomenclature.
3. Limiting reagent decides amount of excess reagents consumed and the amount of products formed.

Finding the excess reagent left

For the reaction
$2P + Q \longrightarrow R$,

If initially, **8 *mol* of P** and **5 *mol* of Q** are taken, find the **amount of excess reagent left.**

a) **4 *mol*** b) **3 *mol*** c) **2 *mol*** d) **1 *mol***

Solution

Step 1:
Finding limiting reagent and excess reagent

Given:
Initial moles of P = 8 *mol*

Initial moles of Q = 5 *mol*

$$\frac{\text{Initial moles of P}}{\text{Stoic.Coeff}} = \frac{n_{Pi}}{2} = \frac{8}{2} = 4$$

$$\frac{\text{Initial moles of Q}}{\text{Stoic.Coeff}} = \frac{n_{Qi}}{1} = \frac{5}{1} = 5$$

Ratio of P is less than Q. So,

P - Limiting reagent

Q - Excess reagent

Step 2:
Finding amount of excess reagent left

Using stoichiometry,

$$\frac{\text{Initial moles of P}}{2} = \frac{\text{moles of P consumed}}{2} = \frac{\text{moles of R produced}}{1} = \frac{\text{moles of Q consumed}}{1}$$

$$\frac{\text{Initial moles of P}}{2} = \frac{\text{moles of Q consumed}}{1} \Rightarrow \frac{8}{2} = \frac{\text{moles of Q consumed}}{1} = 4$$

Amount of excess reagent left = Initial moles of Q - moles of Q consumed
= 5 - 4 = 1 *mol*

Therefore, option d is correct.

Finding the volume of product formed and percentage of reactant reacted

27 *g* Al is heated with **49 *mL* of H_2SO_4 (sp. gravity = 2) and produces H_2 gas.** Calculate **the volume of H_2 gas at STP** and % **of Al reacted with H_2SO_4.**

$$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$$

Solution

Step 1:
Balanced chemical equation
$2Al + 3H_2SO_4 \longrightarrow Al_2(SO_4)_3 + 3H_2$

Step 2:
Finding initial moles of reactant

We know,

$$\text{Specific gravity} = \frac{\text{density of the substance (g/mL)}}{\text{density of water (g/mL)}}$$

$\therefore$ Density of H_2SO_4 = 2 × 1 = 2 *g/mL*.
Mass = density × volume = 2 × 49 = 98 *g*

Since values are given in grams,

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{Molar mass}}$$

$$\therefore \text{Number of moles of Al} = \frac{27\ g}{27\ g/mol} = 1\ mol$$

$$\text{Number of moles of } H_2SO_4 = \frac{98\ g}{98\ g/mol}$$

$$= 1\ mol$$

Step 3:
Finding limiting reagent and excess reagent

For Al:

$$\frac{\text{Initial moles of Al}}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of Al}}{2} = \frac{1}{2} = 0.5$$

For H_2SO_4:

$$\frac{\text{Initial moles of } H_2SO_4}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } H_2SO_4}{3} = \frac{1}{3} = 0.33$$

$\Rightarrow$ The ratio for H_2SO_4 is less than Al. H_2SO_4 is limiting reagent and Al is excess reagent

Step 4:
Finding volume of H_2 produced and moles of Al consumed
According to stoichiometric law:

$$\frac{\text{moles of } H_2SO_4 \text{ consumed}}{\text{Stoic.Coeff}} = \frac{\text{moles of } H_2 \text{ produced}}{\text{Stoic.Coeff}}$$

$$\frac{\text{moles of } H_2SO_4 \text{ consumed}}{3} = \frac{\text{moles of } H_2 \text{ produced}}{3}$$

Moles of H_2 produced = $\frac{1}{3} \times 3 = 1\ mol = 22.4\ L$ (at STP i.e. T = 0°C, P = 1 *atm*)

$$\frac{\text{moles of Al consumed}}{\text{Stoic.Coeff}} = \frac{\text{moles of } H_2SO_4 \text{ consumed}}{\text{Stoic.Coeff}}$$

$$\text{i.e.,} \frac{\text{moles of Al consumed}}{2} = \frac{\text{moles of } H_2SO_4 \text{ consumed}}{3}$$

$$\text{Moles of Al consumed} = \frac{1}{3} \times 2 = \frac{2}{3} = 0.667 \ mol$$

Step 5:
Percentage of Al reacted

$$\text{Percentage of Al reacted} = \frac{0.667}{1} \times 100 = 66.7\%$$

Finding amount (in litres) of product formed

How many litres of ammonia gas is produced at STP from 16 g H_2 and 70 g N_2.

a) 112 *L* **b) 244 *L*** **c) 122 *L*** **d) 326 *L***

Solution

Step 1:
Balanced chemical equation

$$N_2 + 3H_2 \longrightarrow 2NH_3$$

Step 2:
Finding initial moles of reactant

$$\textit{Number of moles} = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$\text{Number of moles of } N_2 = \frac{70 \ g}{28 \ g/mol} = 2.5 \ mol$$

$$\text{Number of moles of } H_2 = \frac{16 \ g}{2 \ g/mol} = 8 \ mol$$

Step 3:
Finding limiting reagent and excess reagent

For N_2

$$\frac{\text{Initial moles of } N_2}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } N_2}{1} = \frac{2.5}{1} = 2.50$$

For H_2:

$$\frac{\text{Initial moles of } H_2}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } H_2}{3} = \frac{8}{3} = 2.66$$

$\Rightarrow$ The ratio for N_2 is less than H_2. N_2 is limiting reagent and H_2 is excess reagent

Step 4:
Finding Number of NH_3 produced

According to stoichiometric law:

$$\frac{\text{moles of } NH_3 \text{ produced}}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } N_2}{\text{Stoic.Coeff}}$$

$$\frac{\text{moles of } NH_3 \text{ produced}}{2} = \frac{2.50}{1}$$

$$\text{Moles of } NH_3 \text{ produced} = 2.5 \times 2 = 5 \ mol$$

$$\text{Volume of } NH_3 \text{ produced} = 5 \times 22.4 \text{ (at STP)} = 112 \ L$$

Therefore, option a is correct.

Finding the amount of excess reagent

Find the amount of the **excess reagent left in grams if 16 *g* H_2** and **70 *g* N_2** combine to form **NH_3.**

Solution

Step 1:
Balanced chemical equation

$N_2 + 3H_2 \longrightarrow 2NH_3$

Step 2:
Finding initial moles of reactant

$$\textit{Number of moles} = \frac{\textit{Given mass}}{\textit{Molar mass}}$$

$$\text{Number of moles of } N_2 = \frac{70\ g}{28\ g/mol} = 2.5\ mol$$

$$\text{Number of moles of } H_2 = \frac{16\ g}{2\ g/mol} = 8\ mol$$

Step 3:
Finding limiting reagent and excess reagent

For N_2:

$$\frac{\text{Initial moles of } N_2}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } N_2}{1} = \frac{2.5}{1} = 2.50$$

For H_2 :

$$\frac{\text{Initial moles of } H_2}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of } H_2}{3} = \frac{8}{3} = 2.66$$

⇒ The ratio for N_2 is less than H_2. N_2 is limiting reagent and H_2 is excess reagent

Step 4: Finding the amount of excess reagent (H_2) consumed

3 × number of moles of L.R. = 3 × 2.5 = 7.5 *mol*

Step 5:
Finding the amount of excess reagent left

According to stoichiometric law:

$$\frac{\text{Initial moles of } N_2}{\text{Stoic.Coeff}} = \frac{\text{moles of } H_2 \text{ consumed}}{\text{Stoic.Coeff}}$$

$$\frac{\text{Initial moles of } N_2}{1} = \frac{\text{moles of } H_2 \text{ consumed}}{3}$$

But initial moles of H_2 = 8 *mol*

⇒ Excess reagent (H_2) left = 8 - 7.5 = 0.5 *mol*

Mass = Moles × molar mass = 0.5 × 2 = 1 *g*

MAIN ADVANCED

Limiting reagent
Multiple step reactions (Sequential)

Sequential reactions consist of linked reactions in which the product of one reaction becomes the reactant of another reaction.

Example of sequential reactions is:

$2A + 3B \longrightarrow 4C + 5D \longrightarrow 5E + 6F$

Finding the amount of product formed in sequential reactions

Reaction 1 : 2A + 3B $\longrightarrow$ 4C + 5D
Reaction 2 : 3C + 4E $\longrightarrow$ 5G + 6F
Starting with **7 moles of A, 6 moles of B** and **8 moles of E,** find the **moles of G formed.**

Solution

Step 1:
Finding limiting reagent in reaction
2A + 3B $\longrightarrow$ 4C + 5D ....(i)
For A :

$$\frac{\text{Initial moles of A}}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of A}}{2} = \frac{7}{2} = 3.5$$

For B :

$$\frac{\text{Initial moles of B}}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of B}}{3} = \frac{6}{3} = 2$$

⇒ The ratio for B is less than A. Hence B is limiting reagent and A is excess reagent

Step 2:
Finding Moles of C produced
According to stoichiometric law

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff}} = \frac{\text{moles of C consumed}}{\text{Stoic.Coeff}}$$

$$\frac{6}{3} = \frac{\text{moles of C produced}}{4}$$

so, moles of C formed = $\frac{6}{3} \times 4 = 8\ mol$

Step 3:
Finding limiting reagent in reaction
3C + 4E $\longrightarrow$ 5G + 6F
Since it is a sequential reaction, the amount of C produced in reaction (i) will be used for formation of products.

For C:

$$\frac{\text{Moles of C produced in reaction (i)}}{\text{Stoic.Coeff}}$$

$$\Rightarrow \frac{\text{Moles of C produced in reaction (i)}}{3} = \frac{8}{3} = 2.66$$

For E:

$$\frac{\text{Initial moles of E}}{\text{Stoic.Coeff}} = \frac{\text{Initial moles of E}}{4} = \frac{8}{4} = 2$$

⇒ The ratio for E is less than C. Hence E is limiting reagent and C is excess reagent

Step 4:
Finding moles of G produced
According to stoichiometric law:

$$\frac{\text{moles of E consumed}}{\text{Stoic.Coeff}} = \frac{\text{moles of G consumed}}{\text{Stoic.Coeff}}$$

$$\Rightarrow \frac{8}{4} = \frac{\text{moles of G consumed}}{5}$$

so, moles of G produced = $\frac{8}{4} \times 5 = 10\ mol$

Finding moles of product produced

2A + 3B $\longrightarrow$ 4C + 5D yield = 70%
3C + 4E $\longrightarrow$ 5G + 6F
Starting with **7 moles of A, 6 moles of B** and **8 moles of E,** find the **moles of G formed.**

Solution

Step 1:

Finding Limiting Reagent for 1st reaction

Given: initial moles of A = 7 *mol*

Initial moles of B = 6 *mol*

Yield of 1st reaction = 70%

For reaction, $2A + 3B \longrightarrow 4C + 5D$

$$\frac{\text{Initial moles of A}}{\text{Stoic.Coeff of A}} = \frac{n_A}{2} = \frac{7}{2} = 3.5$$

$$\frac{\text{Initial moles of B}}{\text{Stoic.Coeff of B}} = \frac{n_B}{3} = \frac{6}{3} = 2.0$$

Hence, B is limiting reagent and A is excess reagent.

Step 2:

Finding Moles of C formed

According to stoichiometry,

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff of B}} = \frac{\text{moles of C produced}}{\text{Stoic.Coeff of C}}$$

$$\frac{6}{3} = \frac{\text{moles of C produced}}{4}$$

so, moles of C formed $= \frac{6}{3} \times 4 = 8\ mol$

As, Yield of reaction is 70%

so, actual moles of C formed

$$= 8 \times \frac{70}{100} = 5.6\ mol$$

Step 3:

Finding Limiting Reagent for 2nd reaction

Given: initial moles of E = 8 *mol*

For reaction, $3C + 4E \longrightarrow 5G + 6F$

$$\frac{\text{Initial moles of C}}{\text{Stoic.Coeff of C}} = \frac{n_C}{3} = \frac{5.6}{3} = 1.87$$

$$\frac{\text{Initial moles of E}}{\text{Stoic.Coeff of E}} = \frac{n_E}{4} = \frac{8}{4} = 2.0$$

Hence, C is limiting reagent and E is excess reagent.

Step 4:

Finding Moles of G formed

According to stoichiometry,

$$\frac{\text{moles of C consumed}}{\text{Stoic.Coeff of C}} = \frac{\text{moles of G produced}}{\text{Stoic.Coeff of G}}$$

$$\frac{5.6}{3} = \frac{\text{moles of G produced}}{5}$$

so, moles of G formed $= \frac{5.6}{3} \times 5 = 9.33\ mol$

Finding number of moles of unreacted reactant

$2A + 3B \longrightarrow 4C + 5D$ $\qquad$ $3C + 4E \longrightarrow 5G + 6F$

Starting with **7 moles of A, 6 moles of B** and **8 moles of E,** find the **moles of A and C left unreacted** once the overall reaction has gone to completion.

(Assume 100% yield for both individual reactions.).

Solution

Step 1:

Finding Limiting Reagent for 1st reaction

Given: initial moles of A = 7 *mol*

Initial moles of B = 6 *mol*

For reaction, $2A + 3B \longrightarrow 4C + 5D$

$$\frac{\text{Initial moles of A}}{\text{Stoic.Coeff of A}} = \frac{n_A}{2} = \frac{7}{2} = 3.5$$

$$\frac{\text{Initial moles of B}}{\text{Stoic.Coeff of B}} = \frac{n_B}{3} = \frac{6}{3} = 2.0$$

Hence, B is limiting reagent and A is in excess

Step 2:

Finding Moles of A left unreacted

According to stoichiometry,

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff of B}} = \frac{\text{moles of A consumed}}{\text{Stoic.Coeff of A}}$$

$$\frac{6}{3} = \frac{\text{moles of A consumed}}{2}$$

so, moles of A consumed $= \frac{6}{3} \times 2 = 4\ mol$

i.e Moles of A left unreacted = 7 - 4 = 3 *mol*

Step 3:

Finding Moles of C formed

According to stoichiometry,

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff of B}} = \frac{\text{moles of C formed}}{\text{Stoic.Coeff of C}}$$

$$\frac{6}{3} = \frac{\text{moles of C formed}}{4}$$

so, moles of C formed $= \frac{6}{3} \times 4 = 8\ mol$

Step 4:

Finding Limiting Reagent for 2nd reaction

Given: initial moles of E = 8 *mol*

For reaction, $3C + 4E \longrightarrow 5G + 6F$

$$\frac{\text{Initial moles of C}}{\text{Stoic.Coeff of C}} = \frac{n_C}{3} = \frac{8}{3} = 2.6$$

$$\frac{\text{Initial moles of E}}{\text{Stoic.Coeff of E}} = \frac{n_E}{4} = \frac{8}{4} = 2.0$$

Hence, E is limiting reagent and C is excess reagent.

Step 5:

Finding moles of C left unreacted

According to stoichiometry,

$$\frac{\text{moles of E consumed}}{\text{Stoic.Coeff of E}} = \frac{\text{moles of C consumed}}{\text{Stoic.Coeff of C}} \Rightarrow \frac{8}{4} = \frac{\text{moles of C consumed}}{3}$$

So, moles of C consumed $= \frac{8}{4} \times 3 = 6\ mol$

i.e moles of C left unreacted = 8 - 6 = 2 *mol*

Finding moles of product produced

$2A + 3B \longrightarrow 4C + 5D$ **$7C + 8E \longrightarrow 9F + 10G$** **$5F + 6H \longrightarrow 7I + 8J$**

Starting with **3 moles of A, 4.4 moles of B, 2 moles E and 1 mole of H,** find the **moles of I formed.**

Solution

Step 1:
Finding Limiting Reagent for 1st reaction
Given: initial moles of A = 3 *mol*
Initial moles of B = 4.4 *mol*
For reaction, $2A + 3B \longrightarrow 4C + 5D$

$$\frac{\text{Initial moles of A}}{\text{Stoic.Coeff of A}} = \frac{n_A}{2} = \frac{3}{2} = 1.5$$

$$\frac{\text{Initial moles of B}}{\text{Stoic.Coeff of B}} = \frac{n_B}{3} = \frac{4.4}{3} = 1.47$$

Hence, B is limiting reagent and A is excess reagent.

Step 2:
Finding Moles of C formed
According to stoichiometry,

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff of B}} = \frac{\text{moles of C produced}}{\text{Stoic.Coeff of C}}$$

$$\frac{4.4}{3} = \frac{\text{moles of C consumed}}{4}$$

so, moles of C formed = $\frac{4.4}{3} \times 4 = 5.87$

Step 3:
Finding Limiting Reagent for 2nd reaction
Given: initial moles of E = 2 *mol*
For reaction, $7C + 8E \longrightarrow 9F + 10G$

$$\frac{\text{Initial moles of C}}{\text{Stoic.Coeff of C}} = \frac{n_C}{7} = \frac{5.87}{7} = 0.84$$

$$\frac{\text{Initial moles of E}}{\text{Stoic.Coeff of E}} = \frac{n_E}{8} = \frac{2}{8} = 0.25$$

Hence, E is limiting reagent and C is excess reagent.

Step 4:
Finding Moles of F formed
According to stoichiometry,

$$\frac{\text{moles of E consumed}}{\text{Stoic.Coeff of E}} = \frac{\text{moles of F produced}}{\text{Stoic.Coeff of F}}$$

$$\frac{2}{8} = \frac{\text{moles of F formed}}{9}$$

so, moles of F formed = $\frac{2}{8} \times 9 = 2.25$ *mol*

Step 5:
Finding Limiting Reagent for 3rd reaction
Given: initial moles of H = 1 *mol*
For reaction, $5F + 6H \longrightarrow 7I + 8J$

$$\frac{\text{Initial moles of F}}{\text{Stoic.Coeff of F}} = \frac{n_F}{5} = \frac{2.25}{5} = 0.45$$

$$\frac{\text{Initial moles of H}}{\text{Stoic.Coeff of H}} = \frac{n_H}{6} = \frac{1}{6} = 0.167$$

Hence, H is limiting reagent and F is excess reagent.

Step 6:
Finding moles of I formed
According to stoichiometry,

$$\frac{\text{moles of H consumed}}{\text{Stoic.Coeff of H}} = \frac{\text{moles of I formed}}{\text{Stoic.Coeff of I}}$$

$$\Rightarrow \frac{1}{6} = \frac{\text{moles of I formed}}{7}$$

So, moles of I formed = $\frac{1}{6} \times 7 = 1.167$ *mol*

Finding moles of product produced

$2A + 3B \longrightarrow 4C + 5D$ **$7C + 8E \longrightarrow 9F + 10G$ Yield = 20%**
$5F + 6H \longrightarrow 7I + 8J$
Starting with **3 moles of A, 4.4 moles of B, 2 moles E and 1 mole of H,** find the **moles of I formed.**

Solution

Step 1:

Finding Limiting Reagent for 1st reaction

Given: initial moles of A = 3 *mol*

Initial moles of B = 4.4 *mol*

For reaction, $2A + 3B \longrightarrow 4C + 5D$

$$\frac{\text{Initial moles of A}}{\text{Stoic.Coeff of A}} = \frac{n_A}{2} = \frac{3}{2} = 1.5$$

$$\frac{\text{Initial moles of B}}{\text{Stoic.Coeff of B}} = \frac{n_B}{3} = \frac{4.4}{3} = 1.47$$

Hence, B is limiting reagent and A is excess reagent.

Step 2:

Finding Moles of C formed

According to stoichiometry,

$$\frac{\text{moles of B consumed}}{\text{Stoic.Coeff of B}} = \frac{\text{moles of C produced}}{\text{Stoic.Coeff of C}}$$

$$\frac{4.4}{3} = \frac{\text{moles of C consumed}}{4}$$

so, moles of C formed $= \frac{4.4}{3} \times 4 = 5.87$

Step 3:

Finding Limiting Reagent for 2nd reaction

Given: initial moles of E = 2 *mol*

For reaction, $7C + 8E \longrightarrow 9F + 10G$

$$\frac{\text{Initial moles of C}}{\text{Stoic.Coeff of C}} = \frac{n_C}{7} = \frac{5.87}{7} = 0.84$$

$$\frac{\text{Initial moles of E}}{\text{Stoic.Coeff of E}} = \frac{n_E}{8} = \frac{2}{8} = 0.25$$

Hence, E is limiting reagent and C is excess reagent.

Step 4:

Finding Moles of F formed

According to stoichiometry,

$$\frac{\text{moles of E consumed}}{\text{Stoic.Coeff of E}} = \frac{\text{moles of F produced}}{\text{Stoic.Coeff of F}}$$

$$\frac{2}{8} = \frac{\text{moles of F formed}}{9}$$

so, moles of F formed $= \frac{2}{8} \times 9 = 2.25\ mol$

As yield = 20%

i.e., moles of F formed = $2.25 \times 0.20 = 0.45\ mol$

Step 5:

Finding Limiting Reagent for 3rd reaction

Given: initial moles of H = 1 *mol*

For reaction, $5F + 6H \longrightarrow 7I + 8J$

$$\frac{\text{Initial moles of F}}{\text{Stoic.Coeff of F}} = \frac{n_F}{5} = \frac{0.45}{5} = 0.09$$

$$\frac{\text{Initial moles of H}}{\text{Stoic.Coeff of H}} = \frac{n_H}{6} = \frac{1}{6} = 0.167$$

Hence, F is limiting reagent and H is excess reagent.

Step 6:

Finding moles of I formed

According to stoichiometry,

$$\frac{\text{moles of F consumed}}{\text{Stoic.Coeff of F}} = \frac{\text{moles of I formed}}{\text{Stoic.Coeff of I}}$$

$$\Rightarrow \frac{0.45}{5} = \frac{\text{moles of I formed}}{7}$$

So, moles of I formed $= \frac{0.45}{5} \times 7 = 0.63\ mol$

CHEMISTRY

MOLE CONCEPT

MOLE FINALE

What you already know

- Molarity
- Stoichiometry
- Molality
- Limiting Reagent

What you will learn

- Mole Fraction
- Percentage and fraction concentrations
- Strength of a solution
- Parts per million
- Parts per billion
- Interconversions between various Concentration terms
- Stoichiometry and concentration terms

Concentration Terms

The concentration of a solution is the amount of solute dissolved in a known amount of solvent or solution. Some of the concentration terms are explained below.

MAIN

Mole Fraction

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

- Mole fraction is denoted by 'χ'.

For a solution of A and B,

Mole fraction of A

$$= \frac{\text{Number of moles of A}}{\text{Number of moles of solution}}$$

$$\chi_A = \frac{n_A}{n_A + n_B} \quad \text{similarly}, \ \chi_B = \frac{n_B}{n_A + n_B}$$

Sum of mole fractions = $\chi_A + \chi_B$

$$= \frac{n_A}{n_A + n_B} + \frac{n_B}{n_A + n_B} = \frac{n_A + n_B}{n_A + n_B} = 1$$

For a solution containing 'i' number of components,

Mole fraction of 1st component

$$\chi_1 = \frac{n_1}{n_1 + n_2 + .. + n_i} = \frac{n_1}{\Sigma n_i}$$

$$\text{Similarly}, \ \chi_2 = \frac{n_2}{\Sigma n_i} \ \text{and} \ \chi_i = \frac{n_i}{\Sigma n_i}$$

Sum of mole fractions = $\chi_1 + \chi_2 + ... + \chi_i$

$$= \frac{n_1}{\Sigma n_i} + \frac{n_2}{\Sigma n_i} + ... + \frac{n_i}{\Sigma n_i} = \frac{n_1 + n_2 + ... + n_i}{\Sigma n_i} = \frac{\Sigma n_i}{\Sigma n_i} = 1$$

- Mole fraction is a dimensionless number.
- Mole fraction is independent of temperature.

Finding mole fraction from the given mass.

Calculate **mole fraction of each species** in a container having **44 g of CO_2, 64 g of O_2, and 28 g of N_2 gas** respectively.

Solution

Step 1:

Finding moles of CO_2, O_2, and N_2

$$n_{CO_2}\ (n_1) = \frac{Given\ weight}{Molar\ mass} = \frac{44}{44} = 1\ mol$$

$$n_{O_2}\ (n_2) = \frac{Given\ weight}{Molar\ mass} = \frac{64}{32} = 2\ mol$$

$$n_{N_2}\ (n_3) = \frac{Given\ weight}{Molar\ mass} = \frac{28}{28} = 1\ mol$$

Step 2:

Dividing the moles of individual species by total number of moles

$$\chi_{CO_2} = \frac{n_1}{n_1 + n_2 + n_3} = \frac{1}{1+2+1} = 0.25$$

$$\chi_{O_2} = \frac{n_1}{n_1 + n_2 + n_3} = \frac{2}{1+2+1} = 0.50$$

$$\chi_{N_2} = \frac{n_1}{n_1 + n_2 + n_3} = \frac{1}{1+2+1} = 0.25$$

Finding mass from the given mole fraction.

What is the **quantity of water (in g)** that should be added to **16 g of methane** to make the **mole fraction of methane as 0.25**?

Solution

Step 1:

Find moles of all components

Moles of methane (n_{CH_4}) =

$$\frac{weight\ of\ methane}{molar\ mass\ of\ methane} = \frac{16}{16} = 1\ mol$$

Let moles of water (n_{H_2O}) be n.

Step 2:

Find mole fraction

Mole fraction of CH_4

$$\chi_{CH_4} = \frac{n_{CH_4}}{n_{CH_4} + n_{H_2O}} = \frac{1}{1+n}$$

Step 3:

Equate obtained χ_{CH_4} to given value to find moles of H_2O

$$\frac{1}{1+n} = 0.25$$

$\therefore$ n = 3 *mol*

Step 4:

Multiply n with molar mass with to get mass of H_2O

Mass of H_2O = n × 18 = 3 × 18 = 54 *g*

Concentrations of a Solution (pph)

- It refers to the amount of the solute per 100 parts of a solution.
- it can also be called as parts per hundred (pph).

BOARDS

Different Percentage Concentrations

Weight by Weight Percentage (% w/w)

- It is given as the mass of solute present in per 100 *g* of solution.
- It is denoted by (% w/w).

$$(\%\ w/w) = \frac{\text{mass of solute in g}}{\text{mass of solution in g}} \times 100$$

Example: 10 (% w/w) ethanol solution consists of 10 *g* of ethanol for every 100 *g* of the solution.

Weight by Volume Percentage (% w/V)

- It is given as mass of solute present in per 100 *mL* of solution.
- It is denoted by (% w/V).

$$(\%\ w/V) = \frac{\text{mass of solute in g}}{\text{volume of solution in mL}} \times 100$$

Example: 10 (% w/V) Na_2CO_3 consists of 10 *g* of Na_2CO_3 for every 100 *mL* of solution.

Volume by Volume Percentage (% V/V)

- It is given as volume of solute present in per 100 *mL* of solution.
- It is denoted by (% V/V).

$$(\%\ V/V) = \frac{\text{volume of solute in mL}}{\text{volume of solution in mL}} \times 100$$

Example: 10 (% V/V) ethanol solution consists of 10 *mL* of ethanol for every 100 *mL* of the solution.

MAIN

Strength of a Solution

The term can be defined as the amount of solute mass (in gram) present in one-litre solution.

- The strength is expressed in *g/L*.

$$\text{Strength of solution} = \frac{\text{solute mass in g}}{\text{volume of solution in L}} = \frac{\text{solute mass (g)} \times 1000}{\text{volume of solution (mL)}}$$

- Unit for the strength of a solution is *g/L*.
- Strength of a solution depends on temperature.

Strength of a solution

If **10 g of glucose** is dissolved in **100 mL of glucose solution.** Calculate **the strength of the solution.**

Solution

Given mass of glucose in 100 mL of glucose solution = 10 g

$$\textit{Strength of solution} = \frac{\textit{solute mass (g)} \times 1000}{\textit{volume of solution (mL)}} = \frac{10 \times 1000}{100} = 100$$

∴ Strength of the solution = 100 *g/L*

MAIN

Parts Per Million (ppm)

ppm is defined as the number of parts of solute particles present in per million parts of the solution.
ppm is defined in terms of (w/w), (w/V) and (V/V) respectively as:

$$ppm\ (w/w) = \frac{\text{weight of solute (g)}}{\text{weight of solution (g)}} \times 10^6$$

$$ppm\ (w/V) = \frac{\text{weight of solute (g)}}{\text{volume of solution (mL)}} \times 10^6$$

$$ppm\ (V/V) = \frac{\text{volume of solute (mL)}}{\text{volume of solution (mL)}} \times 10^6$$

This method of measuring concentration is particularly important when solute is present in trace quantities.

Concentration of a solution in ppm

What is the **concentration of a solution, in parts per million,** if **0.02 grams of NaCl** is dissolved in **1 kg of the solution**?

Solution

Given

Amount of NaCl in solution = $2 \times 10^{-2}\ g$

Amount of water (solution) = $10^3\ g$

$$ppm\ (w/w) = \frac{\text{weight of solute (g)}}{\text{weight of solution (g)}} \times 10^6 = \frac{2 \times 10^{-2}}{10^3} \times 10^6 = 20\ ppm$$

Parts Per Billion (ppb)

ppb is defined as the number of parts of solute particles present in per billion parts of the solution.

$$ppb\ (\%\ w/w) = \frac{weight\ of\ solute\ (g)}{weight\ of\ solution\ (g)} \times 10^9$$

Finding molality from molarity

If you are given a **2 *M* NaOH solution** having density **1 *g/mL***, then find the **molality of the solution**.

Solution

Step 1:

Assume volume of solution to be 1000 *mL* and find weight of solution.
Given density of solution = 1 g/mL

So,
weight of 1000 mL of the solution = 1000 × 1 = 1000 g

Step 2:

Find the weight of solute and solvent using molarity.
Since the given molarity of solution is 2 M, 1000 mL of solution should contain 2 moles of NaOH.
So, weight of $NaOH$ = 2 × 40 = 80 g

Weight of solvent = 1000 − 80 = 920 g

Step 3:

Find molality of NaOH.
Molality of NaOH

$$= \frac{no.of\ moles}{weight\ of\ solvent} \times 1000 = \frac{2 \times 1000}{920} = 2.174\ m$$

Finding molarity from molality

Find the **molarity** of **5 *m* NaOH solution** having a **density of 1.5 *g/mL***.

Solution

Step 1:

Assume mass of solvent to be 1000 g and find mass of solution.

For 1000 g of solvent, 5 moles of solute i.e, NaOH is present.

weight of solute ($NaOH$) = 5 × 40 = 200 g
So, *weight of solution*
= *weight of solvent* + *weight of solute*
= 1000 + 200
= 1200 g

Step 2:

Divide weight of solution with its density to get volume.

Volume of solution

$$= \frac{\textit{weight of solution}}{\textit{density of solution}} = \frac{1200}{1.5} = 800\ mL$$

Step 3:

Find molarity of solution

$$\textit{Molarity of the solution} = \left(\frac{\textit{no.of moles}}{\textit{volume}}\right) \times 1000$$

$$= \left(\frac{5}{800}\right) \times 1000$$

$$= 6.25\ M$$

- When molarity is given, assume that the volume of solution is 1000 *mL* or 1 *L*.
- When molality is given, assume that the mass of solvent is 1000 *g* or 1 *kg*.

Finding mole fraction from molarity

Given a **2 *M* NaOH solution** having **density 1 *g/mL***. Find the **mole fraction of solute**.

Solution

Step 1:

Assume volume of solution to be 1000 *mL*. Find weight of solution.

Weight of solution = volume of solution × density of solution

$= 1000 \times 1 = 1000\ g$

Step 2:

Use molarity of solution to find weight of solute and solvent.

Since molarity is 2M,

Weight of solute (NaOH) = 2 × molar mass

$= 2 \times 40 = 80\ g$

Weight of solvent = weight of solution − weight of solute

$= 1000 - 80 = 920\ g$

Step 3:

Use mass of solvent with to get moles of solvent and χ_{NaOH}.

$$\text{Moles of solvent} = \frac{\textit{mass of solvent}}{\textit{molar mass of solvent}}$$

$$= \frac{920}{18} = 51.11\ mol$$

$$\text{Hence, mole fraction} = \frac{2}{(2 + 51.11)} = 0.0376$$

Finding mole fraction from molality

Given a **5 *m* NaOH solution** having **density 1.5 *g/mL***. Find the **mole fraction of solute**.

Solution

5 *m* solution will contain 5 moles of NaOH in 1 *kg* of solvent.

Moles of solvent = $\frac{1000}{18}$ = 55.56 *mol*

$$\text{Mole fraction of solute} = \frac{\text{moles of solute}}{\text{moles of solute + moles of solvent}}$$

$$= \frac{5}{5 + 55.56} = 0.0826$$

Finding molality from mole fraction

Mole fraction of A in **H_2O is 0.2**. The **molality of A** in H_2O is:

a) 13.9 *m* b) 15.5 *m* c) 14.5 *m* d) 16.8 *m*

Solution

Step 1:

Calculate the amount of water present for 0.2 *mol* of A.

$\chi_{H_2O} + \chi_A = 1$

$\chi_{H_2O} = 1 - 0.2 = 0.8$

For 0.2 *mol* of A we have 0.8 *mol* of H_20

Amount of water = 0.8 × 18 = 14.4 *g*

Step 2:

Placing the values of moles of solute and amount of solvent to get molality

$$\text{molality} = \frac{\text{moles of solute} \times 1000}{\text{amount of solvent (g)}}$$

$$= \frac{0.2 \times 1000}{14.4} = 13.9\ m$$

Relation between (% w/w) and molarity

Density for **2 *M* CH_3COOH solution** is **1.2 *g/mL***. Calculate **(% w/w) for the solution**.

Solution

Step 1:

Calculate the amount of solute and solution in 1 *L* of solution.

Number of moles of CH_3COOH in 1 *L* = 2 *mol*

Mass of CH_3COOH = 2 × Molar mass

= 2 × 60 = 120 *g*

Mass of the solution

= *Density of the solution* × *Volume of solution*

= 1.2 *g/mL* × 1000 *mL* = 1200 *g*

Step 2:

Place the values to get (% w/w) of the solution

$$(\%\ w/w) = \frac{\text{mass of solute (g)}}{\text{Mass of solution (g)}} \times 100$$

$$= \frac{120}{1200} \times 100 = 10\ \%$$

Relation between (% w/V) and molarity

Density for **2 M CH_3COOH** solution is **1.2 g/mL**. Calculate **(% w/v) for the solution**.

Solution

Step 1:

Calculate the amount of solute in 1 L of solution.

Number of moles of CH_3COOH in 1 L = 2 mol

Mass of CH_3COOH

= 2 × Molar mass

= 2 × 60 = 120 g

Step 2:

Place the values to get (% w/v) of the solution

$$(\%\ w/V) = \frac{\textit{mass of solute (g)}}{\textit{Volume of solution (mL)}} \times 100$$

$$= \frac{120}{1000} \times 100 = 12\ \%$$

Relation between molality, molarity and mole fraction

The **molality of HNO_3 solution is 2 m (specific gravity = 1.50)** then find the following:

a) Molarity of the solution **b) Mole fraction of the solute**

Solution

Step 1:

Calculate the density of the solution.

$$\text{Specific gravity} = \frac{\textit{Density of the solution}}{\textit{Density of water at } 4^{0}\ C}$$

Density of the solution = 1.5 × 1 g/mL

= 1.5 g/mL

Step 2:

Finding the volume of solution for fixed moles of solute

2 moles of solute present in 1 kg of solution

$$\text{Volume of solution} = \frac{\textit{mass of solution}}{\textit{density of solution}}$$

$$= \frac{1000\ g}{1.5\ g/mL} = 666.67\ mL$$

$$\text{a) Molarity} = \frac{\textit{Moles of solute}}{\textit{Volume of solution in L}}$$

$$= \frac{2}{0.667} = 3\ M$$

Step 3:

Finding moles of solvent and mole fraction of solute

$$\text{Moles of solvent} = \frac{1000}{18} = 55.56\ mol$$

b) Mole fraction of solute

$$= \frac{\textit{moles of solute}}{\textit{moles of solute + moles of solvent}}$$

$$= \frac{2}{2+55.56} = 0.035$$

Relation between (%w/w) and molarity

What is the **molarity of H_2SO_4 solution** that has **a density of 1.84 *g/cc*** and contains **98 % by mass of H_2SO_4**? (Given atomic mass of S = 32 *u*)

a) 4.18 *M* **b) 8.14 *M*** **c) 18.4 *M*** **d) 18.0 *M***

Solution

Step 1:

Assume 1 *L* of the solution is taken and find the mass of 1 *L* of the solution.

Mass of the solution = density of solution × volume of solution

$= 1.84 \times 1000 = 1840\ g$

Step 2:

Use (% w/w) to find mass of H_2SO_4

$$\text{Mass of } H_2SO_4 = 1840 \times \frac{98}{100} = 1803.2\ g$$

Step 3:

Use mass of H_2SO_4 to find moles of H_2SO_4 and molarity

$$\text{Moles of } H_2SO_4 = \frac{1803.2}{98} = 18.4\ mol$$

$$\text{Molarity} = \frac{\textit{No. of moles of solute}}{\textit{Volume of solution in L}}$$

$$= \frac{18.4\ mol}{1\ L} = 18.4\ M$$

Relation between (% w/w) and ppm

A solution has a concentration of **2550 μg of solute per kg of solution (% w/w)**. What is its **concentration in *ppm***?

Solution

Given

(% w/w) concentration = 2550 $\mu g\ kg^{-1}$

So, the mass of solute is $2.55 \times 10^{-3}\ g$ for $10^3\ g$ of solution,

Mass of solute for $10^6\ g$ or $10^3 \times 10^3\ g$ of solution = $2.55 \times 10^{-3} \times 10^3 = 2.55\ g$

Hence, concentration of solution = 2.55 *ppm*

Stoichiometry and Concentration Terms

- Stoichiometry gives the relation between moles consumed and/or moles produced.
- So, when amounts of reactants or products are given in terms of concentration, we need to convert these amounts into moles.

Concentration terms and stoichiometry

32.65 *g* of a Zn metal reacts with **350 *mL* of 0.5 *M* HCl**. Calculate **the weight of $ZnCl_2$ formed**. (Take molar mass of Zn = 65.3 and Cl = 35.5 *g/mol*)

Solution

Step 1:

Find moles of Zn and HCl

$$\text{Moles of Zn} = \frac{\textit{Given weight}}{\textit{Molar mass}}$$

$$= \frac{32.65}{65.30} = 0.5\ mol$$

Moles of HCl

= *Molarity of HCl × Volume of HCl solution*

= 0.5 × 0.350 = 0.175 *mol*

Step 2:

Find the limiting reagent

$$\frac{0.175}{2} < \frac{0.5}{1}$$

∴ HCl is limiting reagent

Step 3:

Use stoichiometry to find moles of $ZnCl_2$

Zn	+	*2HCl*	$\longrightarrow$	$ZnCl_2$	+	H_2
0.5 *mol*		**0.175 *mol***		**0**		**0**
0.4125 *mol*		**0**		**0.0875 *mol***		**0.0875 *mol***

Mass of $ZnCl_2$ = moles of $ZnCl_2$ × molar mass of $ZnCl_2$

= 0.0875 × 136.3 = 11.93 *g*

Concentration terms and stoichiometry

For the reaction,

CsOH	+	**HI**	$\longrightarrow$	**CsI**	+	**H_2O**
37.5 *g* in 500 *mL*		**500 *mL* of 0.8 M**				

Select the correct option(s):

a) Acid is the limiting reagent

b) Molarity of H^+ in resulting solution is 0.2 *M*

c) Molarity of anion in resulting solution is 0.4 *M*

d) Base is the limiting reagent

Solution

Step 1:

Finding the millimoles of CsOH and HI

$$\text{Moles of CsOH} = \frac{\textit{Given weight}}{\textit{Molar mass of CsOH}}$$

$$= \frac{37.5}{150} = 0.25\ mol$$

$$\text{Molarity} = \frac{\textit{number of moles of solute}}{\textit{volume of solution in L}}$$

Number of moles HI = 0.8 × 0.5 = 0.4 *mol*

Step 2:

Identify the limiting reagent and perform the stoichiometric calculations.

$$\frac{0.40}{1} > \frac{0.25}{1}$$

∴ CsOH (base) is the limiting reagent.

Final volume of the solution = *Volume of acid* + *Volume of Base* = 1000 *mL*

$$\text{Molarity of anion} = \frac{\textit{number of moles of anion}}{\textit{volume of solution}} = \frac{0.4}{1} = 0.4\ M$$

Concentration terms and stoichiometry

The **amount of $BaSO_4$** formed upon mixing **100 *mL* of 20.8 % $BaCl_2$ solution** with **50 *mL* of 9.8 % H_2SO_4 solution will be:**

Atomic mass (*g/mol*) (Ba = 137, Cl = 35.5, S = 32, H = 1 and O = 16)

a) 17.60 *g*   **b) 11.65 *g***   **c) 30.60 *g***   **d) 33.20 *g***

Solution

Step 1:

Finding the moles of $BaCl_2$ and H_2SO_4

$$(\%\ w/V) = \frac{\textit{weight of solute in g}}{\textit{volume of solution in mL}} \times 100$$

$$\text{Weight of } BaCl_2\ (g) = \frac{20.8 \times 100}{100} = 20.8\ g$$

$$\text{Moles of } BaCl_2 = \frac{20.8}{(137+2\times 35.5)} = 0.1\ mol$$

$$\text{Weight of } H_2SO_4\ (g) = \frac{9.8 \times 50}{100} = 4.9\ g$$

$$\text{Moles of } H_2SO_4 = \frac{4.9}{98} = 0.05\ mol$$

Step 2:

Write the balanced chemical equation and perform the stoichiometric calculations.

$BaCl_2$	+	H_2SO_4	$\longrightarrow$	$BaSO_4$	+	$2\ HCl$
0.1 *mol*		0.05 *mol*				
0.05 *mol*		0		0.05 *mol*		0.1 *mol*

H_2SO_4 is limiting reagent.

Amount of $BaSO_4$ formed = *Moles of $BaSO_4$* × *Molar mass of $BaSO_4$* = 0.05 × 233 = 11.65 *g*