ULTIMATE KCET

CRASH COURSE 2026

Chemistry | Lecture - 01

Some basic concept of chemistry

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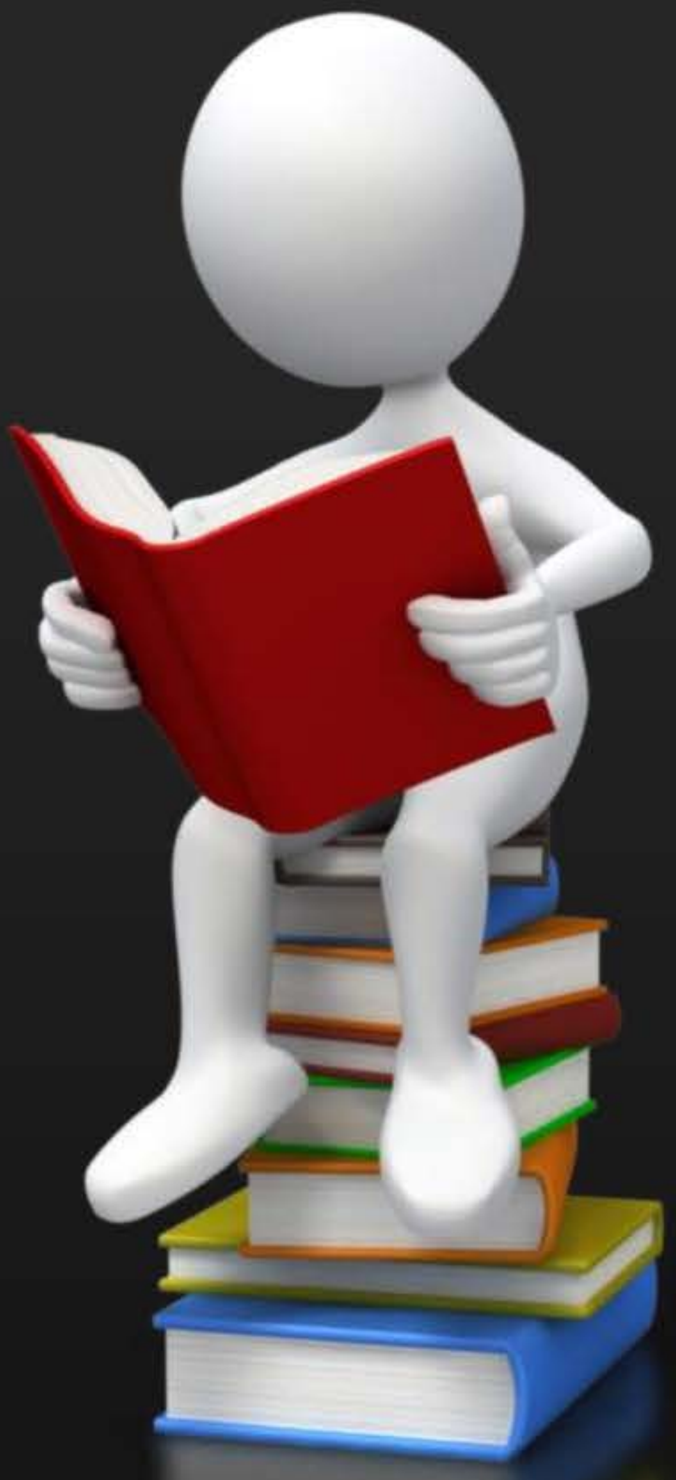
Topics *to be covered*

1. Introduction to matter
2. Mole concepts
3. Stochiometric calculation
4. Concertation calculation



SYNOPSIS

Importance of chemistry

- The branch of science that studies the preparation, properties, structure and reactions of material substances is called chemistry.

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Development of chemistry

Chemistry - came up as a result of search for two interesting things:

(i) Philosopher's stone (Paras) which would convert all baser metals e.g., iron and copper into gold.

(ii) 'Elexir of life' which would grant immortality.

Development of chemistry

- Sushruta Samhita explains the importance of Alkalies.

- The Charaka Samhita mentions ancient indians who knew how to prepare sulphuric acid, nitric acid and oxides of copper, tin and zinc; the sulphates of copper, zinc and iron and the carbonates of lead and iron.

- Rasopanishada describes the preparation of gunpowder mixture.

- Tamil texts also describe the preparation of fireworks using sulphur, charcoal, saltpetre (i.e., potassium nitrate), mercury, camphor, etc.

QUESTION

According to Charaka Samhita, extreme reduction in metal size was part of:

(A) Atomic fusion

(B) Ayurvedic cooling

(C) Nanotechnology

(D) Alkaline testing

Development of chemistry

- Chakrapani discovered mercury sulphide. The credit for inventing soap also goes to him. He used mustard oil and some alkalies as ingredients for making soap.
- Oil of Eranda and seeds of Mahua plant and calcium carbonate were used for making soap.

QUESTION

The preparation of soaps in 18th century India included:

(A) $CaCO_3$ only

(B) Ash and acid

(C) Oil of Eranda and seeds of Mahua // & $CaCO_3$

(D) Lemon and soda

Development of chemistry

VV IMP

- Chemistry provides methods for the isolation of lifesaving drugs from natural sources and makes possible synthesis of such drugs.
- Some of these drugs are cisplatin and taxol, which are effective in cancer therapy.
- The drug AZT (Azidothymidine) is used for helping AIDS patients.
- Safer alternatives to environmentally hazardous refrigerants, like CFCs (chlorofluorocarbons), responsible for ozone depletion in the stratosphere, have been successfully synthesized.
- Preservatives such as sodium benzoate and many others are extensively used in food industry

Matter and its nature:

- Any thing that has mass and occupies space is called matter.
- Depending on temperature and pressure a substance can exist in any one of the three forms of matter.

$$solid \underset{-Energy(freezing)}{\overset{+Energy(melt)}{\rightleftharpoons}} liquid$$

$$liquid \underset{-Energy(condense)}{\overset{+Energy(boil)}{\rightleftharpoons}} gas$$

$$gas \underset{+Energy(sublimation)}{\overset{-Energy(solidification)}{\rightleftharpoons}} solid$$

Base Physical Quantities and Their Units

Base physical quantity	Symbol for quantity	Name of SI Unit	Symbol for SI unit
Length	l	Meter	m
Mass	m	Kilogram	Kg
Time	t	Second	s
Electric Current	I	Ampere	A
Thermodynamic Temperature	T	Kelvin	K
Amount of substance	n	Mole	mol
Luminous Intensity	I_v	Candela	Cd

Multiplex	Prefix	Symbol
10^{-24}	yocto	y
10^{-21}	zepto	z
10^{-18}	atto	a
10^{-15}	femto	f
10^{-12}	pico	p
10^{-9}	nano	n
10^{-6}	micro	μ
10^{-3}	milli	m
10^{-2}	centi	c
10^{-1}	deci	d
10	deca	da
10^{2}	hecto	h
10^{3}	kilo	k
10^{6}	mega	M
10^{9}	giga	G
10^{12}	tera	T
10^{15}	peta	P
10^{18}	exa	E
10^{21}	zeta	Z
10^{24}	Yotta	Y

Some Old Units Still in Use

1. Mass: –

$1\text{Kg} = 10^{3}\text{g} = 1\text{mg}(\text{ milligram }) = 10^{-3}\text{ g} = 1\mu\text{ g}(\text{ micro gram }) = 10^{-6}\text{ g}$

$1\text{amu}(\text{U}) = 1.66 \times 10^{-24}\text{ g} = 1.66 \times 10^{-27}\text{ kg}$

$$1\text{amu} = \frac{1}{12}\text{ mass of isotope of } C^{12}$$

$1\text{amu}(\text{U}) = 1.66 \times 10^{-24}\text{ g} = 1.66 \times 10^{-27}\text{ kg}$

2. Volume:-

$1\text{litre} = 1000\ cm^3 = 1000\text{ml} = 1dm^3 = 10^{-3}\ m^3$

$1\ m^3 = 10^6\ cm^3$

$1m = 100\,cm$

$(1m)^3 = (10^2\,cm)^3$

$1m^3 = 10^6\,cm^3$

$1L \rightarrow 1000\,cm^3$

$\rightarrow 1000\,ml$

deci $= 10^{-1}$

$1L \rightarrow 1dm^3$

$0.1 \times 0.1 \times 0.1 = 10^{-3} m^3$

$10^{-1} \times 10^{-1} \times 10^{-1} = 10^{-3}$

3. Pressure: –

$$1\ \text{atm} = 760\ \text{torr} = 760\ \text{mmHg} = 1.013 \times 10^{5}\ \text{Pa} = 1.01325\text{bar} = 101325\text{Nm}^{-2}$$

$W = -P\Delta V$

- Precision refers to the closeness of various measurements for the same quantity.
- However, accuracy is the agreement of a particular value to the true value of the result.

- For example, if the true value for a result is 2.00 g and student 'A' takes two measurements and reports the results as 1.95 g and 1.93 g. These values are precise as they are close to each other but are not accurate.
- Another student 'B' repeats the experiment and obtains 1.94 g and 2.05 g as the results for two measurements. These observations are neither precise nor accurate.
- When the third student 'C' repeats these measurements and reports 2.01 g and 1.99 g as the result, these values are both precise and accurate

Uncertainity in Measurement

Scientific notation: $N \times 10^n$

where N = a number with a single non-zero digit [1.1 ⟷ 9.9]

eg. $0.0000187 = 1.87 \times 10^{-5}$

$120789 = 1.20789 \times 10^5$

Significant Figures

Significant figures are meaningful digits in a number including the last digit whose value is uncertain.

381

4.8256

Rules to Determine Significant Figures

i. All non-zero digits are significant:

eg: 42.3 has three significant figures.

243.4 has four significant figures.

ii. A zero becomes significant figure if it appears between two non-zero digits:

eg: 5.03 has three significant figures.

5.604 has four significant figures.

iii. Leading zeros or the zeros placed to the left of the numbers are never significant.

eg: 0.543 has three significant figures.

0.045 has two significant figures.

iv. The zeros placed to the right of the decimal are significant.

eg: 433.0 has four significant figures.

433.00 has five significant figures.

v. Counting numbers of objects, for example 2 balls or 20 eggs have infinite significant figures as these are exact numbers and can be represented by writing infinite number of zeros after placing a decimal

i.e $2 = 2.000000$

$20 = 20.000000$

QUESTION

How many significant figures are in the number 0.008941?

A) 2

B) 3

C) 4

D) 5

10

1.0×10^1 → 2SF

1.00×10^1 → 3SF

1.00000×10^1 → 6SF

1.0000000×10^1 → 8SF

$1 \times 10 = 10$

$1.00 \times 10 = 10.0$

➢ The non-significant figures in the measurements are rounded off.

a) If the figure following the last number to be retained is less than 5, all the unwanted figures are discarded and the last number is left unchanged,

eg.: 5.6724 is 5.67 to three significant figures.

b) If the figure following the last number to be retained is greater than 5, the last figure to be retained is increased by 1 unit and the unwanted figures are discarded,

eg. 8.6526 is 8.653 to four significant figures.

c) If the figure following the last number to be retained is 5, the last figure is increased by 1 only in case it happens to be odd. In case of even number, the last figure remains unchanged.

2.3524 is 2.4 to two significant figures.

7.4511 is 7.4 to two significant figures.

94.3% → 94%

94.6% → 95%

even no ← 94.5 → 94% (2 SF)

91.5% → 92%

91 + 1 = 92%

Calculations Involving Significant Figures

➢ Addition and subtractions: → Decimal place

The result of an addition or subtraction in the numbers having different precisions should be reported to the same number of decimal places as are present in the number having the least number of decimal places. The rule is illustrated by the following.

eg: a) 33.3 ← (has only one decimal place)

3.11

0.313

33.3
3.11
0.313
36.723 → 36.7

Sum: 36.723 ← (answer should be reported to one decimal place)

Correct answer = 36.7

b) 62.831 ← (has 3 decimal places)

24.5492

Diff: 38.2818

38.282

(answer should be reported to 3 decimal places after rounding off)

Correct answer = 38.282

Multiplication and division

➢ The answer to a multiplication or division is rounded off to the same number of significant figures as is possessed by the least precise term used in the calculation.

eg: a. 142.06

× 0.23 ← (two significant figures)

32.6738 ← (answer should have two significant figures)

Correct answer = 33

b. $\dfrac{0.90}{4.26} = \mathbf{0.2112676}$

Correct answer = 0.21

Question

Which set of figures will be obtained after rounding up the following up to three significant figures?

34.216, 0.04597, 10.4107

(A) 34.3, 0.0461, 10.4

(B) 34.2, 0.0460, 10.4

(C) 4.20, 0.460, 10.40

(D) 34.21, 4.597, 1.04

34.2

0.0459

0.0459
+1

0.0460

10.4107

Question

How many significant figures are present in 0.010100×10^3?

(A) 7

(B) 3

(C) 10

(D) 5

Question

Two students performed the same experiment separately and each one of them recorded two readings of mass which are given below. Correct reading of mass is 3.0 g. On the basis of given data, mark the correct option out of the following statements.

Student	Readings	
	(i)	(ii)
A	3.01	2.99
B	3.05	2.95

A. Results of both the students are neither accurate nor precise.

B. Results of student A are both precise and accurate.

C. Results of student B are neither precise nor accurate.

D. Results of student B are both precise and accurate.

precision
Accuracy
± 0.23

A → 3.01 2.99

B → 3.05 2.95

→ Accurate

→ Not precise.

Question

If the density of a solution is 3.12 g mL^{-1}, the mass of 1.5 mL solution in significant figures is ______.

(A) 4.7g

(B) 4680×10^{-3}g

(C) 4.680g

(D) 46.80g

$d = \frac{mass}{volume}$ $d = 3.12 \, g\, mL^{-1}$

$V = 1.5 \, mL$

$\therefore$ Mass = d × volume

= 3.12 × 1.5 (3SF) (2SF)

= 1560
312+
4.680

4.7

4. Temperature: –

$$K = {}^0C + 273.15$$

$$^0C = \frac{5}{9} \times ({}^0F - 32)$$

$$F^0 = \frac{9}{5} \times ({}^0C + 32)$$

Dalton's Atomic Theory

$^1H, ^2D, ^3T$

John Dalton in 1808 put forward a theory known as Dalton's atomic theory.

1. Matter is made up of small indivisible particles called atoms.
2. Atoms of the same element are identical in all respects, i.e., size, shape and mass. (isotope)
3. Atoms of different elements have different masses, sizes and also possess different chemical properties. (isobar ^{40}Ca ^{40}Ar)
4. Atoms of the same or different element combine together to form compound [redacted]

H_2O

law of constant proportion.

(fixed ratio by mass)

Law of multiple proportion

H_2O 2:1 CO_2 1:2 CO 1:1

H_2O → 2:1

H_2O_2 → 2:2

5. When atoms combine with one another to form compound atoms ,they do so in simple whole number ratios, such as 1:1, 2: 1, 2:3 etc.

6. Atoms of two elements may combine in different ratios to form more than one compound. For example, sulphur combines with oxygen to form sulphur dioxide and sulphur trioxide, the combining ratios being 1:2 and 1: 3 respectively.

7. An atom can neither be created nor destroyed.

→ Law of conservation of mass.

Limitations of dalton's Atomic Theory

- Theory did not explain why do atoms combine
- It could not explain law of gaseous volume

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The Law of Conservation of Mass

- This law was proposed by Lavoisier ✓
- The law states that "matter can neither be created nor destroyed during a chemical change(or) the total mass of the products formed during a chemical change is exactly equal to the total mass of the reactants".

Exception

→ Nuclear chemical reaction does not obey this law

$$E = mc^2$$

Law of Definite Proportions

- Proposed by Proust
- It is also known as Law of constant proportions.
- A given compound always contains the same elements combined in a fixed proportions by weight.

$H_2O \longrightarrow 2H + 1O$

$2 \times 1g \longrightarrow 16g$

^{1}H

$\overset{8p+8N}{O} \longrightarrow {}^{16}O$

2 : 16

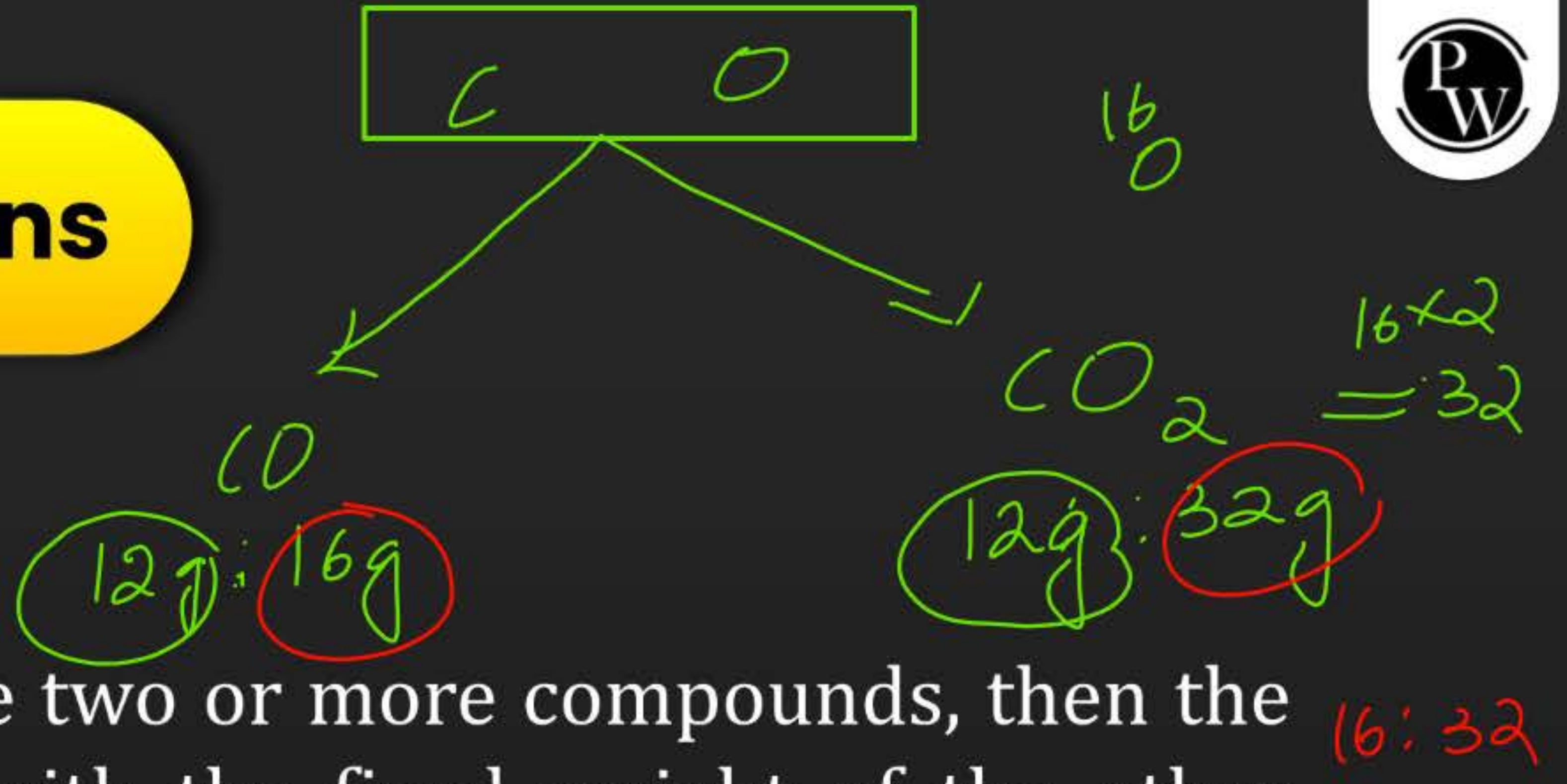
1g : 8g

Exception

└→ isotopes if considered this law can not be justified.

2g : 16g

Law of Multiple Proportions



Proposed by Dalton

- If two elements chemically combine to give two or more compounds, then the weight of one element which combines with the fixed weight of the other element in those compound bear a simple multiple ratio to one another.

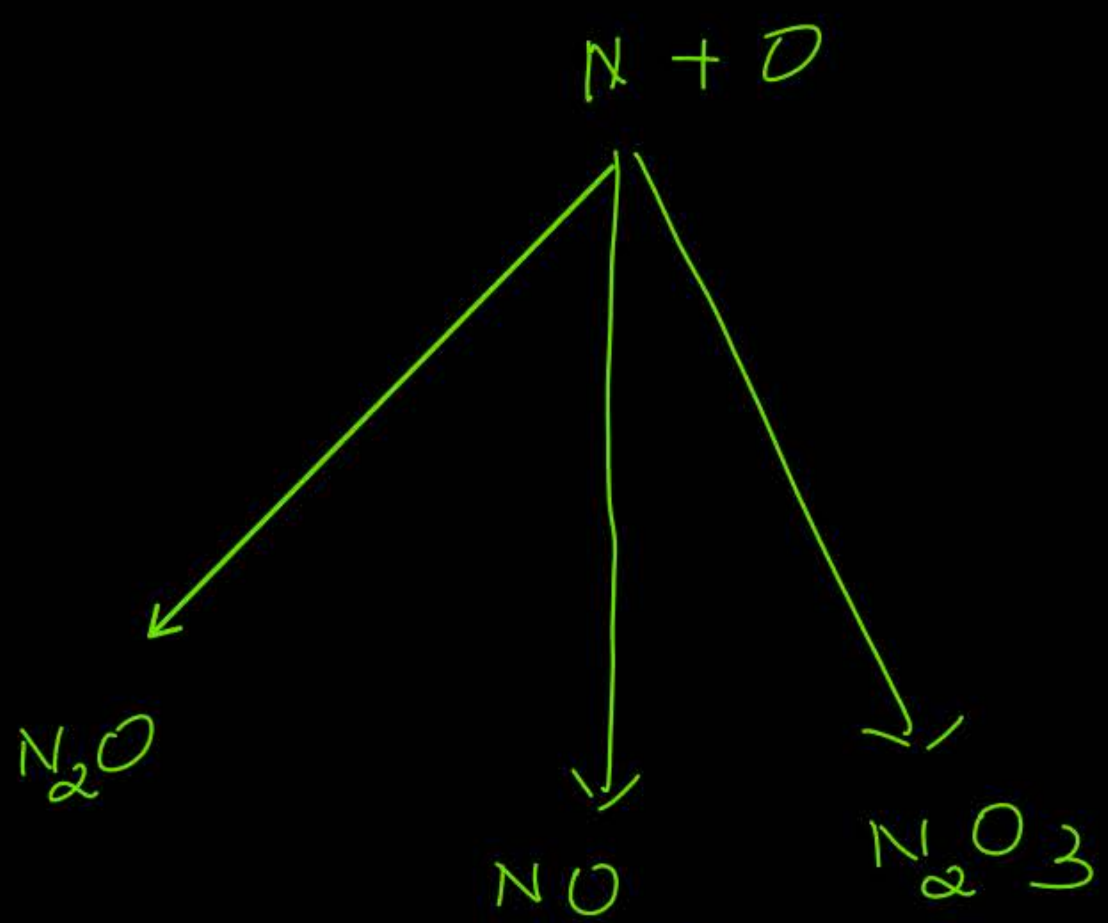
eg: Nitrogen forms the oxides; N_2O, NO, N_2O_3, NO_2, N_2O_5

- In these compounds 28 gm of Nitrogen combines with 16, 32, 48, 64, 80 gm of oxygen respectively. The weight of oxygen in these compounds are in the ratio 16:32:48:64:80 or 1:2 : 3:4:5 a simple multiple ratio.

^{16}O

^{14}N

^{14}N



N_2O: 28 : 16

NO: 14 : 16 (×2, ×2) → 28 : 32

N_2O_3: 28 : 48

16 : 32 : 48

1 : 2 : 3

Gay-Lussac's Law of Combining Volumes

- "Gases combine in the simple whole number ratio of their volumes under similar conditions of temperature and pressure". If products are also gases, the simple whole number ratio also extends to the products.

eg: $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Under similar conditions, 2 lts of Hydrogen combines with 1 lit of oxygen to give 2 lts of water vapour.

It is applicable only to gaseous reaction.

Avogadro's Law

- "Equal volumes of all gases under same conditions of temperature and pressure contain equal number of moles (or) molecules".

Applications

a. Atomicity: Number of atoms present in a molecule is called atomicity

i) Noble gases are monoatomic

eg. He, Ne, Ar, Kr, Xe and Rn

ii) Elementary gases are diatomic

eg. H_2, N_2, O_2, and Cl_2

iii) Ozone, O_3 is triatomic. White phosphorus P_4 is tetra atomic.

iv) Rhombic sulphur S_8 is octa atomic.

b. Relation between molecular mass and vapour density:

molecular mass = 2 × vapour density → VIMP

c. molar volume:

One mole of any gas at STP occupies 22.4 L

{22.4 dm^3, 22.4 L} T = 273 K or 0 °C, P = 1 atm

↓

{22,400 mL, 22,400 cm^3}

Atomic Mass

- 1 a.m.u (U) = 1.66×10^{-24} g
- Atomic weight is also called atomic mass and has no units, since it is relative.
- The absolute mass of an atom of Helium

  $= 4 \times 1.67 \times 10^{-24}$ g $= 6.68 \times 10^{-24}$ g

Question

The following data are obtained when dinitrogen and dioxygen react together to form different Compounds:

Mass of dinitrogen	Mass of dioxygen
14 g 28	16 g 32 1
14 g 28	32 g 64 2
28 g	32 g 32 1
28 g	96 g 96 3

Which law of chemical combination is obeyed by the above experimental data?

A Law of conservation of mass

B Law of definite proportions

C Law of multiple proportions

D Avogadro's Law

Question

1 : 3 : 2

1 L of N_2 combines with 3 L of H_2 to form 2 L of NH_3 under the same conditions. This illustrates the

$1L\,N_2 + 3L\,H_2 \longrightarrow 2L\,NH_3$

(A) law of constant composition

(B) law of multiple proportions

(C) law of conservation of mass

(D) Gay Lussac's law of gaseous volumes.

Question

Dalton's atomic theory could explain the law of

A. conservation of mass

B. definite proportions

C. multiple proportions

D. all of these

Question

One of the statements of Dalton's atomic theory is given below:
"Compounds are formed when atoms of different elements combine in a fixed ratio"
Which of the following laws is not related to this statement?

(A) Law of conservation of mass

(B) Law of definite proportions

(C) Law of multiple proportions

(D) Avogadro law

2 ans

Question

Which of the following statements is correct about the reaction given below:
$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(g)$

A) Total mass of iron and oxygen in reactants = total mass of iron and oxygen in product therefore it follows law of conservation of mass.

B) Total mass of reactants = total mass of product; therefore, law of multiple proportions is followed.

C) Amount of Fe_2O_3 can be increased by taking any one of the reactants (iron or oxygen) in excess.

D) Amount of Fe_2O_3 produced will decrease if the amount of any one of the reactants (iron or oxygen) is taken in excess.

Question

Which of the following reactions is not correct according to the law of conservation of mass.

(A) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

(B) $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$

(C) $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$

(D) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

Question

Which of the following statements indicates that law of multiple proportion is being followed.

(A) Sample of carbon dioxide taken from any source will always have carbon and oxygen in the ratio 1:2.

(B) Carbon forms two oxides namely CO_2 and CO, where masses of oxygen which combine with fixed mass of carbon are in the simple ratio 2:1.

(C) When magnesium burns in oxygen, the amount of magnesium taken for the reaction is equal to the amount of magnesium in magnesium oxide formed.

(D) At constant temperature and pressure 200 mL of hydrogen will combine with 100 mL oxygen to produce 200 mL of water vapour.

Absolute atomic mass – a.m.u / Da / u

$^{4}_{2}He$

4 amu
4 Da
4 u

Absolute atomic mass.

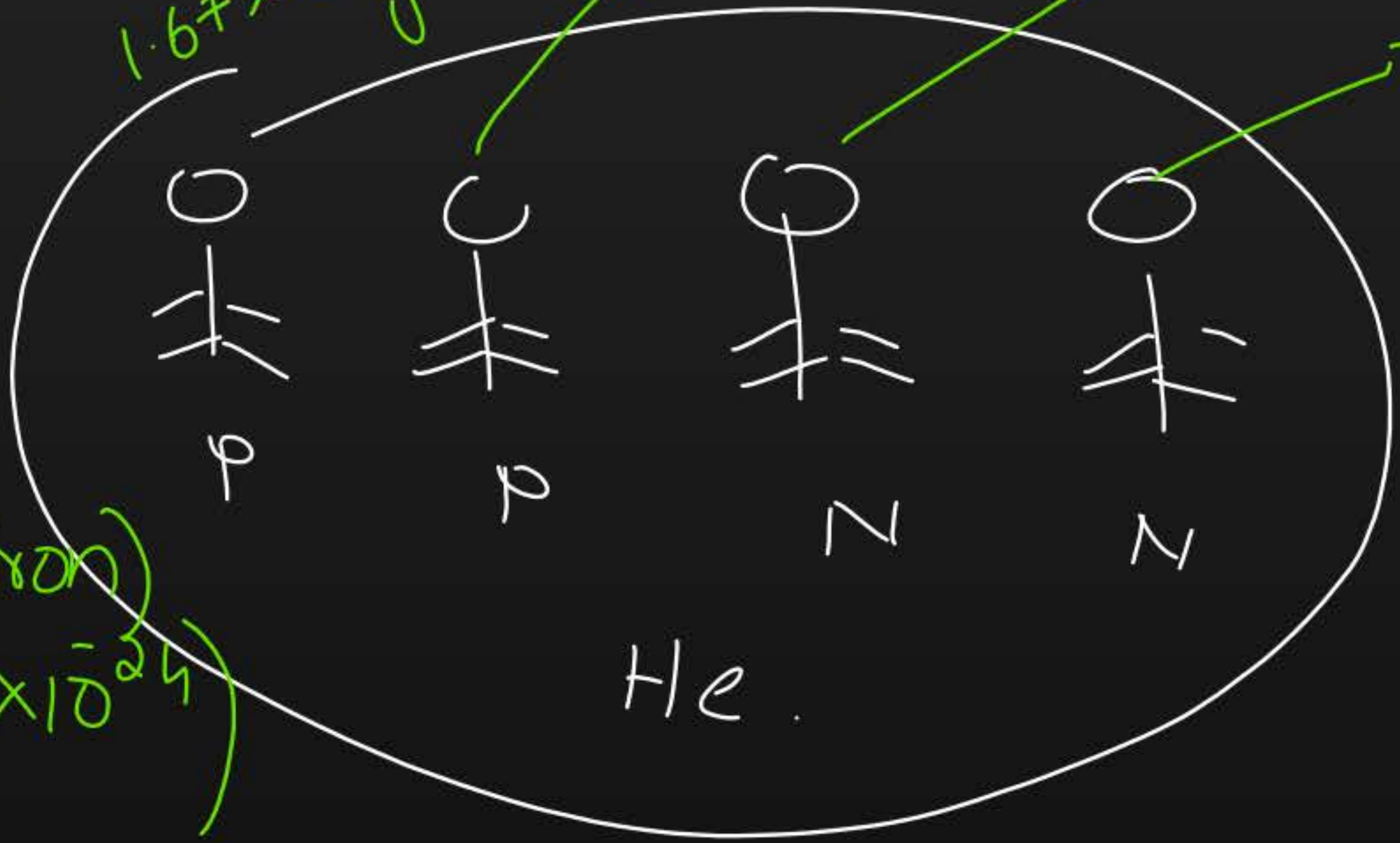
→ (He) → $^{4(2p+2N)}_{2}He$ →

1.67×10^{-24} g

1.67×10^{-24} g

1.67×10^{-24} g

1.67×10^{-24} g



= (No. of protons × mass of proton) + (No. of Neutrons × mass of Neutron)

$(2\times1.67\times10^{-24}) + (2\times1.67\times10^{-24})$

$4\times1.67\times10^{-24}$

4 × amu

a.m.u ⟶ $\frac{1}{12}$th the mass of 1 atom of ^{12}C isotope.

^{12}C ⟶ abundance
⟶ stability

a.m.u $= 1.67 \times 10^{-24}$ g
$= 1.67 \times 10^{-27}$ kg

A.A.m

^{16}O ⟶ 16 amu

16 D

16 u

$16 \times 1.67 \times 10^{-24}$ g

$16 \times 1.67 \times 10^{-27}$ kg

Relative atomic mass

→ No units

$$^{16}O = \frac{\text{Mass of 1 oxygen atom}}{1/12^{th} \text{ the mass of 1 atom of } ^{12}C}$$

$^{16}O \rightarrow$ 16 amu

$$= \frac{16 \text{ amu}}{1 \text{ amu}} = 16$$

Gram atomic mass (Mole).

$1 amu = 1.67 \times 10^{-24} g$.
$1.67 \times 10^{-27} kg$.

Atomic mass expressed in grams.

Gram atomic mass of Oxygen = 16g

→ 1 mole → Molar mass.

Molecular mass 18 D, 18 U, 18 x

$H_2O \rightarrow 2+16 = 18$ amu → absolute molecular mass

= 18 → Relative molecular mass

18g → Gram molecular mass (1 Mole)

Average Atomic Mass :

98.89 g

- Carbon occurs on earth as a mixture of ^{12}C (98.89%) and ^{13}C(1.11%) abundance.
- Isotopic masses of ^{12}C and ^{13}C are 12 amu and 13.0034 amu respectively. The average isotopic mass i.e atomic weight is 12.011 amu.

$$= \frac{\sum \% \text{ abundance} \times \text{Atomic mass}}{100} = \frac{(98.89g \times 12) + (1.11 \times 13)}{100}$$

Average atomic mass = 12.00

Formula weight (or) Formula mass

- Formula weight of a substance is commonly taken as molecular weight. The term formula weight is used for ionic compounds.

NaCl ⟶ 23g + 35.5 g = 58.5 g ⟶ 1 mole

58.5 g ⟶ Gram formula mass

^{23}Na

$^{35.5}Cl$

58.5 amu ⟶ absolute formula mass

58.5 ⟶ Relative formula mass

Mole

- The amount of a substance containing the same number of elementary chemical units as the number of atoms in exactly 12 grams of C-12 isotope is called mole (or)

$$\frac{\text{Given no of atoms / molecules}}{N_A \ (6.022 \times 10^{23})} = \text{Mole}$$

$$\frac{\text{Given volume}}{22.4\,L \ (STP)} \rightarrow \text{Mole}$$

1 mole

1 Gram atom

1 Gram molecule

$$\frac{\text{Given Mass}}{\text{Molar mass}} = \text{Mole}$$

Calculation of Empirical and Molecular formulae

- The formula which represents the simple ratio of the atoms of different elements in one molecule of the substance is called empirical formula

- The formula which represents the exact no. of atoms of each element present in one molecule of the substance is called molecular formula

- Molecular formula = $(\text{Empirical formula})_n$

$$\text{where } n = \frac{\text{Mol. wt.}}{\text{Emp. formula wt.}}$$

Compound	Molecular formula	Emperical formula
Glucose	$C_6H_{12}O_6$	CH_2O
Benzene	C_6H_6	CH
Diborane	B_2H_6	BH_3
Acetic acid	CH_3COOH	CH_2O

$C_2H_4O_2$ $\quad$ $C_1H_2O_1$

Molecular formula and empirical formula will be same for some of the compounds.

Compound	Molecular Formula	Emperical Formula
Methane	CH_4	CH_4
Sulphuric acid	H_2SO_4	H_2SO_4

Percentage yield : $\% \text{ yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$

Stochiometric calculations

$^{16}O \; ^{16}O = 32$

[redacted]

i. Calculations based on weight – weight relationships:

$C + O_2 \longrightarrow CO_2$

12g 32g 44g

ii. Calculations based on weight – volume relationships:

$C + O_2 \longrightarrow CO_2$

12g 32g → 22.4L

iii. Calculations based on volume-volume relationships:

↑ 1mole (only for gas)

$C_{(s)} + O_2 \longrightarrow CO_2$

↓ 22.4L ↓ 22.4L

iv. Calculations based on weight - volume- energy relationships:

v. Calculations based on mole-mole and mole-mass relationship

$C + O_2 \longrightarrow CO_2$

$12g : 32g \longrightarrow$ 1 mole of CO_2

1 mole 1 mole of oxygen → 44g.

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57. Which of the following methods of expressing concentration are unitless?

A (1) Mole fraction and Mass percent (W/W)
(2) Molality and Mole fraction
(3) Mass percent (W/W) and Molality
(4) Molality and Molarity

Mass % = Mass of solute (g) / Mass of solution (g)

58. Select the INCORRECT statement/s from the following:

B (a) 22 books have infinite significant figures   22.000000 . . . ∞

(b) In the answer of calculation 2.5×1.25 has four significant figures ✗

(c) Zero's preceding to first non-zero digit are significant   0.00125 ✗

(d) In the answer of calculation $12.11 + 18.0 + 1.012$ has three significant figures ✓

(1) b, c and d
(2) b and c only
(3) b and d only
(4) a and b only

12.11
18.0
1.012
31.122 → 31.1

Question:

Temperature of 25 °C in Fahrenheit and Kelvin scale respectively are

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A 77 °F and 298.15 K

B 17 °F and 298.15 K

C 45 °F and 260.15 K

D 47 °F and 312.15 K

$K = 273 + °C$

$°F = \frac{9}{5} \times °C + 32$

$K = 273 + 25$

$298 K$

$°F = \frac{9}{5} \times 25 + 32$

$°F = 45 + 32 = 77 °F$

Question

The reference standard used for defining atomic mass is

(A) H-1

(B) C-12

(C) C-13

(D) C-14

Question

The mass of 4 g-atom of S is

The mass of 4 mole of S.

(A) 32 g

(B) 128 g

(C) 64 g

(D) 256 g

$^{32}_{16}S$ (16p + 16N)

G-atom = mole

Gram atom = mole.

4 mole of S.

$$\text{Mole} = \frac{\text{Given Mass}}{\text{Molar mass}}$$

$$4 = \frac{\text{Mass}}{32}$$

$$32 \times 4$$

128 g

Question

The mass of one molecule of water is approximately

(A) 3×10^{-23} g

(B) 18 g → 1 mole → 6.022×10^{23} molecules of H_2O

(C) 1.5×10^{-23} g

(D) 4.5×10^{-23} g

Mass of 1 mole of H_2O = 18 g

1 mole = 6.022×10^{23} molecules

18g → 6.022×10^{23} molecules

→ 1 molecule.

$$3 \times 10^{-23}\,g = \frac{18g \times 1\,molecule}{6.022 \times 10^{23}\,molecule}$$

Question

How many atoms in total are present in 1 kg of sugar?

(A) 7.92×10^{25} atoms

(B) 6×10^{23} atoms

(C) 6.022×10^{25} atoms

(D) 1000 atoms

$C_{12}H_{22}O_{11} \longrightarrow 12\times12 + 1\times22 + 16\times11 = 342g$

$C_{12}H_{22}O_{11}$

1 molecule = 12+22+11 = 45 atoms

18×10^{23} = ? atoms

$\frac{18\times45\times10^{23}}{}$

7.92×10^{25}

1 kg → 1000g

Mole → $\frac{1000}{C_{12}H_{22}O_{11}} = \frac{\text{Given mass}}{\text{Molar mass}}$

$= \frac{1000}{342}$

= 2.92 mole

1 mole = 6.022×10^{23} molecule

2.92 mole = ?

$= 6.022\times10^{23}\times2.92$

$6\times10^{23}\times\times3$

$= 18\times10^{23}$ molecules

Question

Which is heaviest?

(A) 25 g mercury

(B) 2 mol carbon dioxide = 88g

(C) 2 mol water = 36g

(D) 4 g atom of oxygen gas = 128g

g atom → mol.

$O_2 = {}^{16}O + {}^{16}O$
$= 16 \times 2 = 32$

$CO_2 \rightarrow 1 \text{ mole} = 44g$.
$\rightarrow 2 \text{ mole} = ?$

88g.

$12 + 16 \times 2$
$12 + 32 = 44$

H_2O
$\rightarrow 2 + 16$

$\text{Mole} = \frac{\text{Mass}}{\text{Molar mass}}$

$\text{Mole} = \frac{\text{Mass}}{\text{Molar mass}}$

$2 \text{ mol} = \frac{\text{Mass}}{44 \text{ mol g}^{-1}}$

$= 88g$

$2 = \frac{\text{Mass}}{18}$

$= 36g$

$4 = \frac{\text{Mass}}{32}$

32×4
128

Question

The volume occupied by 1 mole atom of a diatomic gas at STP is

(A) 22.4 L

(B) 11.2 L

(C) 5.6 L

(D) 44.8 L

0.5 mole

STP

1 mole = 22.4 L

$1N_2 + 3H_2 \longrightarrow 2NH_3$

1 mole → 22.4 L

3 mole → 3×22.4

2 mole → 2×22.4

Question

An organic compound contains 24% carbon, 4% hydrogen and remaining chlorine. Its empirical formula is

(A) CHCI

(B) CH_2Cl

(C) $CHCl_2$

(D) CH3Cl

$24+4 = 28g$ $\quad 100g - 28 = 72g$ Cl

	Given mass	atomic mass	Mole	simple ratio
C	24	12	$\frac{24}{12} = 2$	1
H	4g	1	$\frac{4}{1} = 4$	2
Cl	72	35.5	$\frac{72}{35.5} = 2$	1

$C_1H_2Cl_1$

Question

Empirical formula of a compound is CH_2O and its molecular mass is 90, the molecular formula of the compound is

(A) $C_3H_6O_3$

(B) $C_6H_{12}O_6$

(C) $C_2H_4O_2$

(D) CH_2O

$$M.F = (E.F)_n$$

$$n = \frac{M.F\ Mass}{E.F\ Mass} = \frac{90}{CH_2O} = \frac{90}{12+2+16} = \frac{90}{30} = 3$$

$n = 3$

Mol. Formula $= (CH_2O)_3$

$= C_3H_6O_3$

Question

1 g → 0.75 g

0.25 g

$^{24}_{12}Mg$ $^{16}_{8}O$ $Mg^{2+} O^{2-}$

1 g of Mg is burnt in a closed vessel containing 0.5 g of O_2. Which reactant is limiting reagent and how much of the excess reactant will be left?

A. O_2 is a limiting reagent and Mg is in excess by 0.25 g.

B. Mg is a limiting reagent and is in excess by 0.5 g.

C. O_2 is a limiting reagent and is in excess by 0.25 g.

D. O_2 is a limiting reagent and Mg is in excess by 0.75 g.

$2Mg + O_2 \longrightarrow 2MgO$

2×24 32

48 g 32 g

1 g ?

$\frac{32}{48}$ = 0.66 g oxygen

$\frac{\frac{1}{2} \times 48}{\frac{48}{2 \times 32}\ 32}$

32 g → 48 g Mg

($\frac{1}{2}$) 0.5 g → ? 0.75 g

$$\underset{2\times24}{2Mg} + \underset{2\times16}{O_2} \longrightarrow \underset{2(24+16)}{2MgO}$$

$48\,g$ of Mg requires $32\,g$ of O_2

$1\,g$ of Mg requires $\frac{32}{48} = 0.66\,g$ of O_2

Oxygen available $= 0.5\,g$

Hence, O_2 is limiting reagent,

$32\,g$ of O_2 reacts with $48\,g$ of Mg

$0.5\,g$ of O_2 will react with $\frac{48}{32} \times 0.5 = 0.75\,g$ of Mg

Excess of $Mg = (1.0 - 0.75) = 0.25\,g$

Question

$CaCO_3$

$40 + 12 + 16 \times 3 = 100g$

$\overset{40}{Ca}$ $\overset{12}{C}$ $\overset{16}{O}$ $HCl \longrightarrow \overset{1}{H} + \overset{35.5}{Cl} = 36.5$

If 40 g of $CaCO_3$ is treated with 40 g of HCl, which of the reactants will act as limiting reagent?

→ excess reagent

(A) $CaCO_3$

(B) HCl

(C) Both (A) and (B)

(D) None of these

$CaCO_3 + 2HCl \longrightarrow CaCl_2 + H_2O + CO_2$ → Equation

100g $36.5 \times 2 = 73g$ → Balance

→ Standard Comparision

100g 73g

40g ? (HCl) $\frac{40g \times 73}{100} = 29.2g$ → Given data analysis

29.2 (HCl)

Solution:

$$\underset{1\text{mole}(100\text{g})}{CaCO_3} + \underset{2\text{moles}(73\text{g})}{HCl} \rightarrow CaCl_2 + H_2O + CO_2$$

$100\,g$ of $CaCO_3$ reacts with $73\,g$ of HCl

$40\,g$ of $CaCO_3$ will react with $\frac{73}{100} \times 40 = 29.2\,g$ of HCl

Since $CaCO_3$ is completely consumed and some amount $(40 - 29.2 = 10.8\,g)$ of HCl remains unreacted and hence, $CaCO_3$ is limiting reagent.

Thank

You