IIT PRAGATI CENTRE

Chapter : Some basic concept of Chemistry

SOLUTIONS

CLASS : XI STATE

Q.1 Select and write the most appropriate answers from given alternatives:

1) Ans. (b)

2) Ans. (d)

3) Ans. (a)

4) Ans. (a)

5) Ans. (c)

6) Ans. (c)

7) Ans. (c)

8) Ans. (c)

9) Ans. (c)

10) Ans. (c)

11) Ans. (b)

12) Ans. (a)

13) Ans. (b)

14) Ans. (d)

15) Ans. (d)

16) Ans. (d)

17) Ans. (a)

18) Ans. (d)

19) Ans. (a)

20) Ans. (d)

21) Ans. (d)

22) Ans. (b)

23) Ans. (a)

24) Ans. (c)

25) Ans. (b)

26) Ans. (b)

27) Ans. (d)

28) Ans. (b)

29) Ans. (c)

30) Ans. (d)

31) Ans. (a)

32) Ans. (a)

33) Ans. (c)

34) Ans. (b)

35) Ans. (c)

36) Ans. (a)

37) Ans. (b)

38) Ans. (a)

39) Ans. ©

Q.2 Answer the following very short questions:

1) Ans. The SI unit for the amount of a substance is mole (mol).

2) Ans. The number of particles in one mole is 6.0221367×10^{23}.

3) Ans. Volume is the amount of space occupied by a three dimensional object. Litre is the common unit of volume.

4) Ans. The basic unit of mass in the SI system is kilogram.

5) Ans. The SI unit of density is kg/m^3.

6) Ans. There are three types of elements as metals, non-metals and metalloids.

7) Ans. Molecular mass of a substance is the mass of one molecule of that substance relative to the mass of one carbon-12 atom. OR Molecular mass of a substance is the sum of average atomic masses of the atoms of an element which constitute the molecule.

8) Ans. ‘Compounds are formed when atoms of different elements combined in a fixed ratio’ is one of the main point of Dalton's atomic theory.

9) Ans. Atomic mass is the mass of an atom. Every element has a characteristic atomic mass. The atomic masses are expressed in amu.

10) Ans. The atomic weight of hydrogen is 1.00797 u and the atomic weight of oxygen is 15.9994 u. So the molecular mass of H_2O is (2 x 1.00797) + 15.9994= 18.01534 u i.e. 18 u.

11) Ans. The density of a substance is calculated by dividing mass by volume.

12) Ans. The law of conservation of mass states that matter cannot be created or destroyed in a chemical reaction.

13) Ans. John Dalton proposed law of multiple proportion.

14) Ans. It is the international system of units and it has seven base units.

15) Ans. Air is a homogeneous mixture of the gaseous substances like nitrogen, oxygen and smaller amount of other substances.

16) Ans. The volume of one mole of a gas at standard temperature and pressure is 22.4L, and here half mole of oxygen will occupy 11.2 liter.

17) Ans. Number of moles of a gas (n) = Volume of the gas at STP/ Molar volume of gas.

18) Ans. If 10 volumes of dihydrogen gas react with 5 volumes of dioxygen gas, then 10 volumes of water vapour would be produced.

19) Ans. The five main branches of chemistry are organic, inorganic, physical, analytical and biochemistry.

20) Ans. Physical chemistry is the study of principles underlying chemistry. It deals with the studies of properties of matter.

21) Ans. Elements are pure substances which cannot be broken down into simpler substances by ordinary chemical changes.

22) Ans. Non-metals are poor conductors of heat and electricity, they cannot be hammered into sheets or drawn into wire because they are brittle and they don't have luster

23) Ans. The mass of a hydrogen atom is 1.6736 x 10^{-24} g.

24) Ans. The volume occupied by one mole of a gas at standard temperature (0°C) and pressure (1 atm) (STP) is called as molar volume of a gas. The molar volume of a gas at STP is 22.4 dm^3.

25) Ans. 12 g of carbon is the molar mass of carbon while 12 u of carbon is the mass of one carbon atom.

26) Ans. One gross means 144 items.

27) Ans. Properties of matter such as mass, length, area, pressure, volume, time etc. are quantitative in nature.

28) Ans. When hydrogen reacts with nitrogen, ammonia is formed.

29) Ans. Avogadro law states that equal volumes of all gases at the same temperature and pressure contain equal number of molecules.

30) Ans. The mass of one hydrogen atom is 1.6736×10^{-24} gram or 1.0 u.

31) Ans. The formula mass of a substance is the sum of atomic masses of the atoms present in the formula.

32) Ans. The mass of one mole a substance in grams is called its molar mass.

33) Ans. The molecular mass of carbon dioxide

= 1(average atomic mass of C) + 2 (average atomic mass of O)

= 1(12.0 u) +2 (16.0 u)

= 44.0 u

34) Ans. The average atomic mass of an element is the weighted average of atomic masses of its isotopes (taking into account the atomic masses of isotopes and their relative abundance i.e. percent occurrence).

35) Ans. Physical properties are those which can be measured and observed without changing the chemical composition of the substance. Examples are colour, odour, melting point, boiling point, density etc.

36) Ans. One dozen means 12 items.

37) Ans. Arsenic, silicon and germanium are some examples of metalloids.

38) Ans. A matter may be n element like Cu, Ag, etc. Or may be a compound like NaCL, H_2O etc.

39) Ans. The law states that, "When two elements A and B form more than one compounds the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers".

Q.3 Answer the following Questions:

1) Ans. i. One mole of a substance is defined as the amount of a substance that contains the number of particles, atoms, molecules, ions or electrons equal to the number of carbon atoms, etc. present in 0.012kg of Carbon-12 i.e., 6.0221367×10^{23} particles.

ii. One mole of a substance contains 6.022×10^{23} molecules while one mole (or one gram atom) of an element contains 6.0221367×10^{23} atoms.

2) Ans. Molar Mass: The mass of one mole of a substance (element/compound) in grams is called its molar mass. The molar mass of any element in grams is numerically equal to atomic mass of that element in u.

Explanation:

Element	Atomic mass(u)	Molar mass ($g\ mol^{-1}$)
H	1.0 u	1.0 $g\ mol^{-1}$
C	12.0 u	12.0 $g\ mol^{-1}$
O	16.0 u	16.0 $g\ mol^{-1}$

Similarly molar mass of any substance, existing as polyatomic molecule, in grams is numerically equal to its molecular mass or formula mass in u.

Polyatomic substance	Molecular formula mass (u)	Molar mass ($g\ mol^{-1}$)
O_2	32.0 u	32.0 $g\ mol^{-1}$
H_2O	18.0 u	18.0 $g\ mol^{-1}$
NaCl	58.5 u	58.5 $g\ mol^{-1}$

3) Ans. 1. The average atomic mass for an element is calculated by summing the masses of element isotopes, each multiplied by its natural abundance on earth.

2. It is needed in order to take the average mass of an element that exist in different isotopic forms .

4) Ans. i. Mass of MgO = 80g

ii. Mass of MgO = 10 g

We can find:

Number of moles of MgO as given below :

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

Molecular mass of MgO = (1 x Average atomic mass of Mg) + (1 x Average atomic mass of O)

= (1 x24u) + (1 x 16 u)

= 40u

Molar mass of MgO = 40 g mol^{-1}

Mass of MgO = 80g

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

$$= \frac{80\ g}{40\ g\ mol^{-1}}$$

= 2 mol

ii. We have Mass of MgO = 10 g

Molar mass of MgO = 40 g mol^{-1}

So, we can find number of moles as:

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

$$= \frac{80\ g}{40\ g\ mol^{-1}}$$

= 0.25 mol

The number of moles in 80 g of magnesium oxide, MgO = 2 mol

The number of moles in 10 g of magnesium oxide, MgO = 0.25 mol

5) Ans. The formula mass of a substance is the sum of atomic masses of the atoms present in the formula.
Formula mass of NaCL= average atomic mass of Na + average atomic mass of CL
= 23.0 + 35.5 u = 58.5 u

6) Ans. We can find Average atomic mass of Boron (B) by:

$$\frac{(\text{At.mass of } {}^{10}B \times \%\text{ Abundance}) + (\text{At.mass of } {}^{21}B \times \%\text{ Abundance}) + (\text{At.mass of } {}^{11}B \times \%\text{ Abundance})}{100}$$

$$= \frac{(10.13\text{ u} \times 19.60) + (11.009\text{ u} \times 80.40)}{100} = 10.84\text{ u}$$

Average atomic mass of boron = 10.84 u.

H. Solution:
It is given:
Temperature in degree Celsius = 40 °C
We know, Temperature in degree Fahrenheit

$°F = \frac{9}{5}(°C) + 32$

Using 40°C in the formula, we get

$°F = \frac{9}{5}(°C) + 32$

$= \frac{9}{5}(40) + 32$

$= 72 + 32$

$= 104\ °F$

Similarly, Temperature in degree Celsius = 30 °C
So, Temperature in degree Fahrenheit

$°F = \frac{9}{5}(°C) + 32$

$= \frac{9}{5}(30) + 32$

$= 54 + 32$

$= 86\ °F$

i. The temperature 40 °C Correspondence to 104 °F
ii. The temperature 30 °C Correspondence to 86 °F

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i. The temperature 40 °C Correspondence to 104 °F
ii. The temperature 30 °C Correspondence to 86 °F

7) Ans. The volume of one mole of a gas is called molar volume. At STP condition (0°C and 1 atm), the molar volume of any gas is 22.4 dm^3.

Q.4 Solve the following:

1) Ans.

2) Ans. We have,

Mass of urea NH_2CONH_2 = 5.6 g

Number of atoms of hydrogen, nitrogen, carbon and oxygen=? (to be solved)

We have:

Molecular formula of urea: $CO(NH_2)_2$

So, Molecular mass of urea: 60 g mol^{-1}

Therefore,

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Molar mass of a substance}} = \frac{5.6\ g}{60\ g\,mol^{-1}} = 00.0933\ mol$$

So, the Moles of urea is found to be = 0.0933 mol

We also know,

Number of atoms N = Number of moles × Avogadro's constant

Now, it has been seen that 1 molecule of urea has 8 atoms, out of which 4 atoms are of Hydrogen, 2 atoms are Nitrogen, 1 of Carbon and 1 of Oxygen.

We can find no. Of atoms of different elements as:

$\therefore$ Number of H atoms in 5.6 g of urea = (4×0.0933) mol $\times 6.022 \times 10^{23}$ atoms/mol

$= 2.247 \times 10^{23}$ **atoms of hydrogen**

$\therefore$ Number of N atoms in 5.6 g of urea = (2×0.0933) mol $\times 6.022 \times 10^{23}$ atoms/mol

$= 1.124 \times 10^{23}$ **atoms of nitrogen**

$\therefore$ Number of C atoms in 5.6 g of urea = (1×0.0933) mol $\times 6.022 \times 10^{23}$ atoms/mol

$= 0.562 \times 10^{23}$ **atoms of carbon**

$\therefore$ Number of O atoms in 5.6 g of urea = (1×0.0933) mol $\times 6.022 \times 10^{23}$ atoms/mol

$= 0.562 \times 10^{23}$ **atoms of oxygen**

So, it has been calculated that:

- 5.6 g of urea contain **2.247** $\times 10^{23}$ **atoms of H,**
- **1.124** $\times 10^{23}$ **atoms of N**
- **0.562** $\times 10^{23}$ **atoms of C**
- **0.562** $\times 10^{23}$ **atoms of O.**

3) Ans. The molecular formula of potassium chlorate is $KClO_3$

The chemical reaction involved is:

$2KClO_3 \rightarrow 2KCl \quad + 3O_2 \uparrow$

[2 moles] [3 moles]

The mass of 2 moles of $KClO_3 = 2 \times 122.5 = 245$ g

And

3 moles of O_2 at STP occupies volume $= (3 \times 22.4\ dm^3) = 67.2\ dm^3$

Hence, 245 g of potassium chlorate will liberate 67.2 dm^3 of oxygen gas.

Let 'x' gram of $KClO_3$ liberate 6.72 dm^3 of oxygen gas at S.T.P.

So,

$$x = \frac{245 \times 6.72}{67.2} = 24.5 \text{ g}$$

Mass of potassium chlorate is found to be = 24.5 g

4) Ans. We know,

Molecular formula of acetaldehyde is given by C_2H_4O

The Moles of acetaldehyde given = 2 mol

So,

- Number of moles of carbon atoms = Moles of acetaldehyde x Number of carbon atoms
  = 2 x 2
  = 4 moles of carbon atoms
- Number of moles of hydrogen atoms = Moles of x Number of hydrogen atoms
  = 2 x 4
  = 8 moles of hydrogen atoms
- . Number of moles of oxygen atoms = Moles of acetaldehyde x Number of hydrogen atoms
  = 2 x 1
  = 2 moles of oxygen atoms
- Number of molecules of acetaldehyde = Moles of acetaldehyde x Avogardo number
  $= 2 \text{ mol} \times 6.022 \times 10^{23}$ molecules of acetaldehyde

5) Ans. **The following information is given that is:**

- Number of moles of Carbon-dioxide = 5 mol
- Number of moles of $CO_2 = 0.5$ mol
- Volume of STP =?(to be solved)

We know,

$$\text{Number of moles (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of gas}}$$

Molar volume of gas = 22.4 $dm^3\ mol^{-1}$ atSTP

$$\text{Number of moles of gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of gas}}$$

i. Volume of the gas at STP = Number of moles of a gas (n) x Molar volume of a gas
$= 5 \text{ mol} \times\ dm^3\ mol^{-1} = 112\ dm^3$

ii. Volume of the gas at STP = Number of moles of a gas (n) x Molar volume of a gas
$= 5 \text{ mol} \times\ dm^3\ mol^{-1} = 11.2\ dm^3$

Volume occupied by 5 mol of CO_2 = 112 dm^3

Volume occupied by 0.5 mol of CO_2 = 11.2 dm^3

6) Ans. Formula mass of $CaSO_4$ = average atomic mass of Ca + average atomic mass of S + 4 x average atomic mass of O
= 40.1 u + 32.1 u + 4 x 16.0 u.
= 136.2 u
Formula mass of $CaSO_4$ = 136.2 u.

Q.5 Answer the following Questions: 24

1) Ans. i. In the year 1811, Avogadro made a distinction between atoms and molecules and thereby proposed Avogadro's law.
ii. Avogadro's proposed that, “Equal volumes of all gases at the same temperature and pressure contain equal number of molecules".
E.g. Hydrogen gas combines with oxygen gas to produce water vapour as follows:

Hydrogen $_{(g)}$	+ Oxygen $_{(g)}$	$\longrightarrow$ Water $_{(g)}$
[100 mL]	[50 mL]	[100 mL]
[2 vol]	[1 vol]	[2 vol]

According to Avogadro's law, if 1 volume contains n molecules, then 2n molecules of hydrogen combining with n molecules of oxygen to give 2n molecules of water, i.e., 2 molecules of hydrogen gas combine with molecule of oxygen to give 2 molecules of water vapour as represented below:

Hydrogen $_{(g)}$	+ Oxygen $_{(g)}$	$\longrightarrow$ Water $_{(g)}$
[2n molecules]	[n molecules]	[2n molecules]
[2 molecules]	[1 molecule]	[2 molecules]

2) Ans.
i. Formula mass of NaCl
= average atomic mass of Na
+ average atomic mass of Cl
= 23.0u + 35.5u = 58.5u

ii. Formula mass of $Cu(NO_3)_2$
= average atomic mass of Cu + 2 x (average atomic mass of nitrogen + average atomic mass of three oxygen)
= (63.5) + 2(14 + 3 x 16) = 187.5 u

3) Ans. 1. It is a smallest particle of an element which may as may not have independent existence.
2. They are very small in size.
3. Their size is measured in unit nm (nanometer) .
4. 1nm – 10^{-9} m.
Molecule
1. It is smallest particle of an element or compound which is able to exist independently.
2. Example : carbon-dioxide molecule, hydrogen gas molecule etc .
The order of magnitude of mass of an atom is 10^{-3} Kg

Isotopes

1. The isotopes of an element are those that have the same number of protons but different number of neutrons in their nuclei.
2. Example :Isotopes of Hydrogen is –Protium, deuterium and Tritium

4) Ans. Law of conservation of mass

According to this law "matter can't be created nor be destroyed in a chemical reaction." That is it always remains constant.

Like, in all chemical reactions the total mass of products is equal to total mass of reactants.

For example:-

$2H_2 + O_2 \longrightarrow 2\ H_2O$

(Reactant) (Product)

=> 4 + 2 x 16.32 → 2 x 18

=> 36 g → 36 g.

This example shows that the total mass of reactants mixture is equal to total mass of products formed which is in accordance with the law.

5) Ans. **6.022×10^{23} are the particles present in 1 mole**

That is:

1 mole of NH_3 contain 6.022×10^{23} molecule of ammonia

Similarly,

1 mole of HNO_3 contains 6.022×10^{23} molecule of Nitric acid

If we look for their ratio, then it comes out to be:

$$\text{Ratio}=\frac{6.022 \times 10^{23}}{6.022 \times 10^{23}}=1:1$$

So, the answer is found to be 1:1(molecules ratio)

6) Ans. (i) French chemist Joseph Proust performed experiments on two samples of cupric carbonate.

(ii) One of the samples was natural in origin and the other sample was synthetic one.

(iii) He found that the composition of the elements present in both the samples were the same.

(iv)

Curpic carbonate	% of copper	% of oxygen	% of carbon
Natural sample	51.35	9.74	38.91
Synthetic sample	51.35	9.74	38.91

(v) This led Joseph to state that "A given compound always contain exactly the same proportions of elements by weight".

(vi) This law has been confirmed by various other experiments. This law is called as law of definite proportions.

7) Ans. Measurement of properties of matter:

(i) Different kinds of matter have characteristic properties called as physical properties and chemical properties.

(ii) Physical properties are those which can be measured or observed without changing the chemical composition of the substance.

(iii) Chemical properties are properties where a substance undergo a chemical change and thereby exhibit a change in the chemical composition.

(iv) Many properties of matter are quantitative in nature and are measurable by comparing with a standard quantity which is unchanging and reproducible.
(v) The arbitrarily decided and universally accepted standards are called as units.
(vi) There are several systems in which units are measured- CGS system, FPS system and MKS system.
(vii) International system of units called as SI system was introduced in the year 1960.

Base Physical quantity	Symbol for the quantity	Name of the SI unit	Symbol for the SI unit
Length	L	metre	M
Mass	M	kilogram	Kg
Time	T	seconds	S
Temperature	T	kelvin	K
Amount of Substance	N	mole	Mol
Luminous Intensity	I	candela	Cd
Electric current	I	ampere	A

8) Ans. (i) This law was proposed by Dalton.
(ii) It was observed that two or more elements may form more than one compound.
(iii)"When two elements A and B form more than one compounds, the masses of element B that combine with a given mass of A are always in the ratio of small whole numbers."
(iv) The validity of this law has been confirmed by many experiments:
(v) Hydrogen + Oxygen → Water
2g 16g 18g
Hydrogen + Oxygen → hydrogen peroxide
2g 32g 34g
(vi) Nitrogen + oxygen → Nitric oxide
14g 16g 30g
Nitrogen + oxygen → Nitrogen dioxide
14g 32g 46g

Q.6 Solve the following:

1) Ans. We are provided with following information i.e.
Mass of 1 atom of hydrogen = 1.008 u
So, Mass of 18 atoms of hydrogen comes out to be
= 18 × 1.008 u = 18.144 u
So, answer is = 18.144 u(i.e. Mass of 18 atoms of hydrogen)

2) Ans. i. 0.4 mole of nitrogen (N)

We can find number of atoms as :

Number of atoms N = Number of moles × Avogadro's number

$= 0.4\ mol \times 6.022 \times 10^{23}\ atoms/mol$

$= 2.4088 \times 10^{23}$ atoms of N

ii. 1.6 g of Sulphur (S)

We know:

Molar mass of sulphur = 32 g mol^{-1}

So,

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Molar mass of a substance}} = \frac{1.6\ g}{32\ g\ mol^{-1}} = 0.05\ mol$$

Therefore,

Number of atoms of S = Number of moles × Avogadro's constant

$= 0.05\ mol \times 6.022 \times 10^{23}\ atoms/mol$

$= 0.3011 \times 10^{23}$ atoms

$= 3.011 \times 10^{22}$ atoms of S

i. Number of nitrogen atoms in 0.4 mole = 2.4088 x 10^{23} atoms of N

ii. Number of sulphur atoms in 1.6 g = 3.011 x 10^{22} atoms of S

3) Ans. We have

Mass of C (Reactant) = 24 g

Mass of CO (product) = 88g

Mass of O (reactant) =? (To be solved)

We have,

12 g of carbon combine with 32g oxygen to form 44 g of carbon dioxide as follows:

Carbon	+	Oxygen	→	Carbon monoxide
12 g		32g		44g

So, (2 x 12 = **24 g**) of carbon will combine with (2 x 32 = 64 g) of oxygen to give (2× 44 = 88 g) CO

Therefore, Mass of oxygen used in reaction is found to be = 64 g

4) Ans. The metal oxide is obtained by two different methods that is:

- Reactions of metal with oxygen
- Reaction of metal with water vapour (steam)

In first case that is reaction of metal with oxygen,

The mass of oxygen in metal oxide = 3.2 – 2.0 = 1.2 g

We can find the percentage by:

$$\%\text{of oxide} = \frac{1.2}{3.2} \times 100 = 37.5\ \%$$

$$\%\text{of metal} = \frac{2.0}{3.2} \times 100 = 62.5\ \%$$

In second case that is reaction with steam

The mass of oxygen in metal oxide = 2.27 – 1.42 = 0.85 g

We can find the percentage by:

$$\%\text{of oxide} = \frac{0.85}{2.27} \times 100 = 37.44 \approx 37.5\ \%$$

$$\%\text{of metal} = \frac{1.42}{2.27} \times 100 = 62.56 \approx 62.5\ \%$$

Therefore it is proved that:

Irrespective of the source, the given compound contains same elements in the same proportion. This is in accordance with Law of definite proportion which states that:

"A given compound always contains exactly the same proportion of elements by weight ". Hence, the law of definite proportions is verified by these data.

5) Ans. We have:

Mass of acetic acid = 22 g

So, number of moles and molecules of acetic acid =?

i. We know: Number of moles = $\dfrac{\text{Mass of substance}}{\text{Molar mass of a substance}}$

ii. We also know that:

Number of molecules = Number of moles x Avogadro's constant

We have,

Mass of acetic = 22 g

Molecular mass of acetic acid, CH_3COOH

= (2 × Average atomic mass of c) + (4 × Average atomic mass of H) + (2 ×Average atomic mass of O)

= (2 × 12 u) + (4 ×1 u) + (2 × 16 u) = 60 u

So, we can find molar mass by:

Molar mass of acetic acid = $\dfrac{\text{Mass of substance}}{\text{Molar mass of a substance}}$

$$\frac{\text{Mass of substance}}{\text{Molar mass of a substance}} = \frac{22\text{ g}}{60\text{ g mol}^{-1}} = 0.367\text{mol}$$

Now, Number of molecules of acetic acid= Number of moles x Avogadro's constant

$= 0.367\text{ mol} \times 6.022 \times 10^{23}$ molecules/mol

$= 2.210 \times 10^{23}$ molecules

Number of moles = 0.367 mol

Number of molecules of acetic acid = 2.210 x 10^{23} molecules

6) Ans. **We are given, Volume** of hydrogen at STP = 0.448 L

We can find Number of moles of hydrogen by:

$$\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

We also know,

Molar volume of a gas (n) = 22.4 $dm^3 mol^{-1}$ = 22.4 L at STP

$$\text{Number of moles of a gas (n)} = \frac{\text{Volume of a gas at STP}}{\text{Molar volume of a gas}}$$

$$= \frac{0.448 \text{ L}}{22.4 \text{ L mol}^{-1}}$$

$= \mathbf{0.02\ mol}$

So, answer is 0.02 mol (that is No. of moles of H)

7) Ans. Volume of NH_3 at STP = 67.2 dm^3

molar volume of a gas = 22.4 $dm^3 mol^{-1}$

Number of moles (n)

$$\frac{\text{Volume of the gas at STP}}{\text{MolarVolumeofgas}}$$

$$\text{Number of moles of } NH_3 = \frac{67.2\,dm^3}{22.4\,dm^3\,mol^{-1}}$$

=3.0 mol

Number of molecules = Number of moles x

6.022 x 10^{23} molecules mol^{-1}

Number of molecules of NH_3 = 3.0 mol x

6.022 x 10^{23} molecules mol^{-1}

=18.066 x 10^{23} molecules

8) Ans.

i. 52 moles of Argon

1 mole Argon atoms = 6.022×10^{23} atoms of Ar

$\therefore$ 52 moles of Ar

$$= 52\ \text{moles} \times \frac{6.022 \times 10^{23}}{1\,\text{mol}}\ \text{atoms}$$

$= 313.144 \times 10^{23}$ atoms of Argon

ii. 52 u of Helium

Atomic mass of He = mass of 1 atom of He = 4.0u

$\therefore$ 4.0 u = 1 He $\therefore$ 52 u = ?

i. 52 moles of Argon

1 mole Argon atoms = 6.022×10^{23} atoms of Ar

$\therefore$ 52 moles of Ar

$$= 52\ \text{moles} \times \frac{6.022 \times 10^{23}}{1\,\text{mol}}\ \text{atoms}$$

$= 313.144 \times 10^{23}$ atoms of Argon

ii. 52 u of Helium

Atomic mass of He = mass of 1 atom of He = 4.0u

$\therefore$ 4.0 u = 1 He $\therefore$ 52 u = ?

$$= 52\ \text{u} \times \frac{1\ \text{atom}}{4.0\ \text{u}} = 13$$

Atoms of He

iii. 52 g of He

Mass of 1 mole of He = 4.0 g

Number of moles of He

$$= \frac{\text{mass of He}}{\text{mass of 1 mole of He}}$$

$$= \frac{52\ \text{g}}{4.0\ \text{g mol}^{-1}} = 13\ \text{mol}$$

Number of atoms of He

= Number of moles x 6.022×10^{23}

= 13 mol x 6.022×10^{23} atoms/mol

= 78.286×10^{23} atoms of He.

9) Ans.i. We know,

5 mg carbon = 5 $\times 10^{-3}$ g carbon

Also,

Atomic mass of carbon = 12 u

Molar mass of carbon = 12 g mol^{-1}

We can fine number of moles as given below:

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Molar mass of a substance}} \times \frac{5 \times 10^{-3}\,g}{12\,g\,mol^{-1}} = 4.167 \times 10^{-4}$$

ii. We know 12 mg carbon = 12 $\times 10^{-3}$ g carbon

Also, atomic mass of carbon is = 12 g

Therefore,

$$\text{Number of moles} = \frac{\text{Mass of substance}}{\text{Molar mass of a substance}} \times \frac{12 \times 10^{-3}\,g}{12\,g\,mol^{-1}} = 1 \times 10^{-3}$$

We can find number of atoms as given:

Number of atoms = Number of moles × Avogadro's constant

Number of atoms of carbon = 1×10^{-3} mol × 6.22×10^{23} atoms/mol

= 6.022×10^{20} atoms

Number of moles of carbon in his homework writing = 4.167 × 10^{-4} mol

Number of atoms of carbon in 12 mg homework writing = 6.022 × 10^{20} atoms

10) Ans.

Average atomic mass of neon (Ne)

$$= \frac{\text{Atomic mass of } ^{20}Ne \times \% + \text{Atomic mass of } ^{21}Ne \times \% + \text{Atomic mass of } ^{22}Ne \times \%}{100}$$

$$= \frac{(19.9924u)(90.92) + (20.9940u)(0.26) + (21.9914u)(8.82)}{100} = 20.1707u$$

11) Ans.

Mass of urea = 5.6 g

Molecular mass of urea, NH_2CONH_2

= 2 (average atomic mass of N) + 4 (average

atomic mass of H) +1 (average atomic mass of C)

+ 1 (average atomic mass of O)

= 2 x 14u+1 x 12u+4 x 1 u+ 1 x 16 u = 60u

$\therefore$ molar mass of urea = 60 g mol^{-1}

Number of moles

$$= \frac{\text{Mass of urea in g}}{\text{Molar Mass of urea in g mol}^{-1}}$$

$$= \frac{5.6\text{ g}}{60\text{ g mol}^{-1}} = 0.0933\text{ mol}$$

Number of molecules = Number of moles x

Avogadro's constant

Number of molecules of urea

$= 0.0933 \times 6.022 \times 10^{23}$ molecules/mol

$= 0.5618 \times 10^{23}$ molecules

$= 5.618 \times 10^{22}$ molecules

Ans : Number of moles = 0.0933 mol

Number of molecules of urea

$= 5.618 \times 10^{22}$ molecules

12) Ans. Molar mass of ethane, C2H6 $= 2 \times 12 + 1 \times 6 = 30\text{ g mol}^{-1}$

Number of moles of C2H6 $= \frac{W}{M}$

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

$\because$ 1 mol of C_2H_6 at STP occupies 22.4 dm^3

$\therefore$ 2 mol of C_2H_6 at STP will occupy,

$V = 2 \times 22.4$

$= 44.8\ dm^3$

Q.7 Answer the following Questions: **12**

1) Ans. (i) This law was put forward by Gay Lussac's in the year 1808.

(ii) The law states "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

(iii) Under the same conditions of pressure and temperature 100mL of hydrogen combines with 50 mL of oxygen to give 100 mL of water.

(iv) Hydrogen(g) + Oxygen(g) → Water

100mL 50mL 100mL

(v) Thus, they bear a simple ratio of 2:1:2.

(vi) Under the same conditions of pressure and temperature 1L of nitrogen combines with 3L of Hydrogen to give 2L of ammonia.

(vii) Nitrogen(g) + Hydrogen(g) → Ammonia(g)

1L 3L 2L

(viii) Thus they bear a simple ratio of 1:3:2.

2) Ans. (i) This theory was proposed by Dalton.

(ii) This theory came into existence in 1808.
(iii) Dalton published the theory as “A New system of chemical philosophy”.
(iv) Matter consists of tiny, indivisible particles called atoms.
(v) All atoms of a given element have identical properties like mass , shape , size etc
(vi) Atoms of different elements are different.
(vii) Compounds are formed when atoms of different elements combine together in a fixed simple whole number ratio.
(viii) Chemical reactions involve only reorganization of atoms.
(ix) Atoms are neither created nor destroyed during a chemical reaction.

3) Ans. (i) Atomic mass is expressed in amu. In 1961 the present system of weight of amu was agreed upon.
(ii) One atomic mass unit is defined “as a mass exactly equal to one twelfth of the mass of one carbon-12 atom.”
(iii) Later on the exact value of 1 amu was experimentally determined.
(iv) 1 amu = 1/12 × mass of one C-12
$1/12 \times 1.9926 \times 10^{-23}$g
$= 1.66056 \times 10^{-24}$g
(v) Recently atomic mass unit has been replaced by unified mass unit called Dalton
(vi) Symbol for unified mass unit is u or Da
(vii) For example : amu of 1 molecule of oxygen is
Molecular mass of oxygen = 2 × 16 u
Therefore, mass of 1 molecule of oxygen = 32u
Mass of 1 molecule of oxygen in grams $= 32 \times 1.660566 \times 10^{-24} = 53.1379 \times 10^{-24}$g

Q.8 Answer the following

1) Ans. 1. Mixtures: (i) Sea water; (ii) Gasoline.
Pure substances: (iii) Skin; (iv) A rusty nail; (v) a page of the textbook; (vi) diamond.
2.

Mass of an atom of oxygen in gram is 26.56896×10^{-24}g , and

$1.66056 \times 10^{-24} g = 1 u$

$\therefore 26.56896 \times 10^{-24} g = ?$

$$= \frac{26.56896 \times 10^{-24} g}{1.66056 \times 10^{-24} g} = 16.0 u$$

Similarly Mass of an atom of hydrogen

$= 1.0080$ u

2) Ans. 1.

No.	Material	Element or compound
i.	Mercuric oxide	Compound
ii.	Helium gas	Element
iii.	Water	Compound
iv.	Table salt	Compound
v.	Iodine	Element
vi.	Mercury	Element
vii	Oxygen	Element
viii.	Nitrogen	Element

2.

Molecular mass of $O_2 = 2 \times 16$ u

$\therefore$ Mass of 1 Molecular= 32 u

$\therefore$ Mass of 1 Molecular of O_2

$= 32.0 \times 1.66056 \times 10^{-24}$ g

$= 53.1379 \times 10^{-24}$ g

3) Ans. 1. (A) An aqueous solution of sugar is an example of a homogeneous mixture.
(B) A mixture of oil and water is an example heterogeneous mixture.
© Cu is an element.
(D) Water, H_2O is a compound.

2. (a) NH_3 = 14 u + 3 x 1 u = 17 u.
(b) CH_3COOH = 2 x 12 u + 4 x 1 u + 2 x 16 u = 60 u.
(c) C_2H_5OH = 2 x 12 u + 6 x 1 u + 16 u = 46 u.