

# Atoms and Molecules

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## Where it fits:

This chapter deals with the basic building blocks of chemistry which will be useful almost everywhere, especially in physical chemistry.

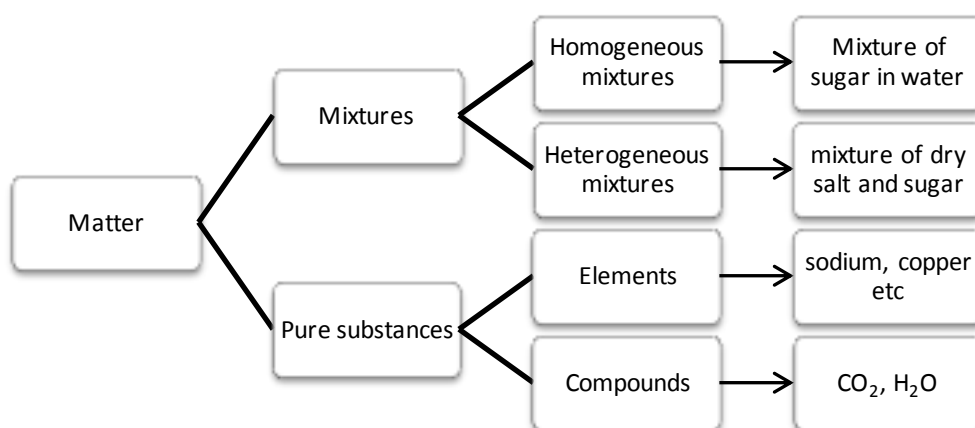
## Motivation:

A building is as strong as its foundation. This chapter is the foundation of chemistry. In a way if your foundation is weak then no matter what you try to build above it will remain weak. Let's get started with some of the basics of chemistry.

## Matter

Anything which has mass and occupies space is called *matter*. The three physical states of matter are solid, liquid and gas. You must be already familiar with these states. *Are there any more states?*

## Chemical Classification of Matter



## Measurement of properties of matter

Many properties of matter such as length, volume etc. is quantitative in nature and their measurement is represented by a number followed by units. E.g. length of a copper wire is 2 m. Here 2 is a number and m denotes *metre* (unit in which length is measured). A universally accepted system of units is SI units.

## International System of Units (SI)

It has the following 7 base units.

Physical Quantity	Length	Mass	Time	Current	Temperature	Amount of substance	Luminous Intensity
Symbol	L	M	T	I	T	n	I <sub>v</sub>
SI unit	metre	kilogram	second	Ampere	Kelvin	mole	candela
Symbol	m	kg	s	A	K	mol	cd

*The units for other physical quantities such as speed, volume, pressure, density, force etc. can be derived from these base units.*

## Metric prefixes for powers of 10

Power	-15	-12	-10	-9	-6	-3	-2	-1	1	2	3	6	9
Prefix	femto	pico	angstrom	nano	micro	milli	centi	deci	deca	hecto	kilo	mega	giga
Symbol	f	p	Å	n	μ	m	C	d	da	h	k	M	G

## Uncertainty in measurement

Measurement of a continuous variable (e.g. mass of substance) can only be as precise as the skill of the measurer and the choice of the measuring apparatus. Error in measurement depends both on the observer and the apparatus. For e.g. the weighing machine used in vegetable market cannot give you the exact weight of the vegetables. The digital one used by shopkeepers is somewhat good measurer of weight.

### Precision

Degree to which further measurements or calculations show the same or similar results.

The volume of an acid is given as 16.0 ml. A student tries to measure the volume of acid using a measuring cylinder. He takes four readings and the values are 15.4 ml, 15.5 ml and 15.4, 15.4ml. As seen the outcome of each result is very close. This is called *precision*.

### Accuracy

Degree of closeness of a measured or calculated quantity to its actual (true) value.

Another student recorded the values as 15.9 ml, 15.8 ml, 16.1 and 16.2 ml. As seen the average value of volume is 16.0 ml. This is called *accuracy*. The example given in precision is not accurate as the average is 15.42 which is very far away from 16.0 ml

### Scientific notation (*exponential notation*)

Any number can be written as  $N \times 10^n$ , where  $n$  is any integer and  $N$  is any real number between 1 and 10. It becomes easier to multiply and divide the numbers in this notation. e.g. 243.702 as  $2.43702 \times 10^2$  and 0.00045 as  $4.5 \times 10^{-4}$

Write the following in scientific notation

- (a) 142
- (b) 12.3
- (c) 12.0
- (d) 1457.234
- (e) 0.0023

Ans: ( $1.42 \times 10^2$ ,  $1.23 \times 10^1$ ,  $1.20 \times 10^1$ ,  $1.457234 \times 10^3$ ,  $2.3 \times 10^{-3}$ )

### Significant figures

The reliability of a measurement is indicated by the number of digits used to represent it. For greater accuracy, we express it with digits that are known with certainty. These are called as Significant figures. They contain all the certain digits plus one doubtful digit in a number. For example if you say the length of a cloth is 12.5 cm. Here it is confirmed that the length is greater than 12 cm and less than 13. Length can be 12.6 or 12.7 or 12.4 etc.

So here there are two certain digits (12) and one doubtful digit .5.

#### Rules for determining number of significant figures

- All non-zero digits are significant. e.g. 34.567 has 5 significant figures.
- Zeroes preceding to first non-zero digit are not significant. Such zero indicates the position of decimal point. e.g. 345, 0.234, 0.004 has 3, 3, 1 significant figures respectively.
- Zeroes between two non-zero digit are significant. e.g. 4003, 0.00302 has 4 and 3 significant figures respectively.
- Zeroes at the end or right of a number are significant provided they are on the right side of the decimal point. e.g. 14.400 have 5 significant figures. 400, 400., 400.0 has 1, 3 and 4 significant figures respectively. This will be clear if we represent them in scientific notations. Such as  $400 = 4 \times 10^2$ ,  $400. = 4.00 \times 10^2$  and  $400.0 = 4.000 \times 10^2$

- Counting numbers of objects have infinite significant figures. e.g. dozen of bananas, as it has 12 bananas, it can be written as 12.0000.....(infinite number of times)

Find the number of significant figures in the following

- (a) 2345
- (b) 234.56
- (c) 0023
- (d) 23400
- (e) 0.00240
- (f) 3 apples
- (g) 500.00
- (h) 2300005
- (i) 230004.5004600

Ans: (4, 3, 3, infinite, 5, 7, 13)

#### Addition and subtraction of significant figures

The result cannot have more digits to the right of the decimal point than either of the original numbers. e.g. adding two numbers 23.45 and 34.567, first one has 2 digits after decimal point while second has 3. So the result *cannot have more than 2 digits after decimal point*.

Find the result in each case

- (a)  $23.45 + 23.453$
- (b)  $24.45 + 22.4 + 234.456$
- (c)  $23 + 24.5 + 22$

#### Multiplication and division of significant figures

The results should have the same number of significant number as the one of the original numbers with least significant figures. When reducing the result to required significant figure, you should not forget rounding off. e.g. Multiplying two numbers 1.30 and 24.56. The significant figure in first one is 3 while that in second is 4. So the result has 3 significant figures (least from original numbers).

#### Rounding off

- If the rightmost digit to be removed is greater than 5, the preceding number is increased by 1. 23.47 can be rounded to 23.5

- If the rightmost digit to be removed is less than 5, the preceding number is not changed. 23.43 can be rounded to 23.4
- If the rightmost digit to be removed is 5, then the preceding number is not changed if it is an even number but it is increased by one if it is an odd number. 23.65 can be rounded to 23.6 and 23.75 can be rounded to 23.8

Find the result in each case:

- (a)  $1.3 \times 3.65$   
 (b)  $234 \times 1.34$   
 (c)  $0.013 \times 123$   
 (d)  $1.5 \times 25$

Ans: (4.7, 314, 1.60, 38)

### Dimensional analysis

It is always required to convert units from one system to other while using them in calculations. The method used to do this is called dimensional analysis.

What is the distance travelled by light in 2 nanoseconds (ns)?

Speed of light is  $3 \times 10^8$  m/s and the time given is 2 ns.

Therefore we need to convert the given time in the required units (here it is second as velocity is in m/s), so time is  $2 \times 10^{-9}$  s.

Now you can find the distance as  $\text{velocity} \times \text{time} = 3 \times 10^8 \times 2 \times 10^{-9} = 0.6$  m

Convert these:

- (a) 54 km/hr in m/s  
 (b)  $1000 \text{ kg/m}^3$  in  $\text{g/cm}^3$

Ans: (15 m/s,  $1 \text{ g/cm}^3$ )

**Example 1:** The correctly reported difference of 16.567 and 6.0 will be

**Soln:**  $16.567 - 6.0 = 10.567$ . The number with least number of decimal places is 6.0

So the answer should have one decimal place and follow the rounding off rules above, you will get 10.6

**Example 2:** The actual product of 8.568 and 1.30 is 11.13840. The correctly reported answer will be

**Soln:** 1.30 is the number with least (3) significant figures

So, the product has 3 significant figures so reported answer will be 11.1

**Example 3:** The value of Planck's constant is  $6.62618 \times 10^{-34}$  Js. Find the number of significant numbers in it. If it has to be expressed to three significant figures, then what will be its value?

**Soln:** The number of significant figures is 6

If it has to be reduced to three, it will be reduced to  $6.63 \times 10^{-34}$  Js.

**Example 4:** Convert the density in SI units. 3 g/ml

**Soln:** Density is expressed as mass per unit volume. The given value has mass in g and volume in ml. To convert this in SI units we have to convert mass in kg and volume in  $m^3$ .

Now  $1 \text{ kg} = 1000\text{g}$  and  $1 \text{ m}^3 = 10^6 \text{ cm}^3 = 10^6 \text{ ml}$

So  $1 \text{ g} = 10^{-3} \text{ kg}$  and  $1 \text{ ml} = 10^{-6} \text{ m}^3$

Density =  $3\text{g}/1\text{ml} = 3 \times 10^{-3}\text{kg}/10^{-6}\text{m}^3 = 3 \times 10^3 \text{ kg/m}^3 = 3000 \text{ kg/m}^3$

## Laws of chemical combination

The elements combine to form compounds and these combinations are governed by some laws.

### 1. Law of conservation of mass:

Matter can neither be created nor destroyed. It means in any chemical process, the total weight of the products is equal to the total weight of reactants. For example 2 gm of A reacts completely with 4 gm of B to form C.  $A+B \rightarrow C$  weight of C formed is  $2+4=6$  gm (*What happens when one of the products is a gas?*)

### 2. Law of Definite proportions:

The same chemical compound, irrespective of its source, always contains the same elements combined in the same fixed proportion by weight.

$\text{CO}_2$  is a compound found in nature, if that compound is prepared in laboratory, it will again be  $\text{CO}_2$  only.

### 3. Law of multiple proportions:

When two elements combine to form two or more compounds, the different weights of one element combining with a fixed weight of the other element are in the simple whole number ratios.

Nitrogen and Oxygen combine to form  $\text{NO}$ ,  $\text{N}_2\text{O}_3$ , 14 gm of N requires 16 gm of O for first and 24 gm of O for second so the ratio of weights of O is 16:24 i.e. 2:3 (simple ratio)

#### 4. *Law of reciprocal proportions:*

When two different elements combine separately with the same weight of a third element, the ratio of the masses in which they do so will be the same or some simple multiple of the mass ratio in which they combine with each other.

Sulphur and oxygen form  $\text{H}_2\text{S}$  and  $\text{H}_2\text{O}$  with hydrogen and  $\text{SO}_2$  among themselves. 2 gram of hydrogen combines with 32 g of sulphur and 16 gram of oxygen ratio of weights of S:O is  $32/16=2:1$ . In  $\text{SO}_2$ , 32 g of sulphur combines with 32 grams with oxygen, the ratio is  $32:32=1:1$ . So the ratio is  $2/1:1/1=2:1$

#### 5. *Gay Lussac's Law of gaseous volumes:*

It is Law of definite proportions by volume for gases. When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume provided all gases are at same temperature and pressure.

100 ml of Hydrogen combines with 50 ml of oxygen to give 100 ml of water vapour. Volume of hydrogen and oxygen is 100:50 i.e. 2:1 (simple ratio)

### Dalton's atomic theory

- Matter is made of atoms. An atom is the smallest indivisible particle of an element, which can take part in a chemical change.
- Atoms of same element have identical properties including identical mass
- Atoms of different elements have different properties and different mass
- Compounds are formed when atoms of different elements combine in a fixed ratio
- Atoms cannot be created or destroyed. So a chemical reaction is rearrangement of atoms.

### Atoms and Molecules

#### Atomic mass

Atom is the smallest particle of an element that takes place in a chemical reaction. It may or may not be capable of free existence. Carbon-12 is assigned a mass of exactly 12 atomic mass unit (amu). One atomic mass unit is one-twelfth the mass of one Carbon-12 atom and numerically

**1 amu =  $1.66 \times 10^{-24}$  g** (see periodic table for elements and their atomic masses).

$$\text{Atomic mass of an element} = \frac{\text{mass of 1 atom of element}}{1/12 \times \text{mass of 1 atom of carbon} - 12}$$

For elements having isotopes, average atomic mass is calculated.

Atomic mass of an element expressed in grams is the gram atomic mass or **gram atom**.

Mass of one atom of oxygen =  $2.656 \times 10^{-23}$  g

Atomic mass of oxygen =  $2.656 \times 10^{-23} / 1.66 \times 10^{-24} = 16$  amu

Find the atomic masses of the following from the periodic table.

- (a) H
- (b) He
- (c) Na
- (d) C
- (e) N
- (f) F
- (g) Cl

Ans: (1, 4, 23, 12, 14, 19, 35.5)

### Molecular Mass

Molecule is the smallest particle of an element or a compound that can exist free and retain all its properties. The number of atoms of an element in a molecule of the element is called its **atomicity**. Molecular mass is the sum of the atomic masses of the elements present in a molecule.

$$\text{Molecular mass of a substance} = \frac{\text{mass of 1 molecule of substance}}{1/12 \times \text{mass of 1 atom of carbon} - 12}$$

Molecular mass of a substance expressed in grams is called **gram molecular mass**.

Find the molecular mass of sugar ( $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ )?

Atomic mass of C, H, O are 12, 1, 16 respectively.

So the molecular mass of sugar is  $12 \times 12 + 22 \times 1 + 11 \times 16 = 144 + 22 + 176 = 342$

Find the molecular mass of the following:

- (a)  $\text{H}_2$
- (b)  $\text{N}_2$
- (c)  $\text{O}_2$
- (d)  $\text{Cl}_2$
- (e)  $\text{CH}_4$
- (f)  $\text{CO}_2$
- (g)  $\text{NH}_3$
- (h)  $\text{C}_2\text{H}_6$
- (i)  $\text{CH}_3\text{OH}$
- (j)  $\text{C}_6\text{H}_{12}\text{O}_6$

Ans: (2, 28, 32, 71, 16, 44, 17, 30, 32, 180)

### Formula Mass (Formula Weight)

A more general term that can be applied to compounds that are not composed of molecules, such as ionic compounds e.g. NaCl (formula mass =  $23 + 35.5 = 58.5$  amu)

Find the formula mass of the following

- (a) ZnO
- (b) Na<sub>2</sub>O
- (c) K<sub>2</sub>CO<sub>3</sub>
- (d) CaCl<sub>2</sub>

Ans: (81.5, 62, 138, 111)

### Avogadro law

Equal volumes of gases at the same temperature and pressure contain equal number of molecules. Avogadro made a distinction between atoms and molecules. This law helps in deducing the atomicity of elementary gases.

For example, consider the formation of water from hydrogen and oxygen. According to Dalton's theory it should have been,  $2\text{H} + \text{O} \rightarrow \text{H}_2\text{O}$ . But Avogadro found that, 2 volumes of hydrogen and 1 volume of oxygen gives 2 volumes of water vapour. This can be explained when we consider hydrogen and oxygen to be diatomic and the reaction to be  $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ . (This hypothesis as you all must be aware of gave rise to the concept of molecules).

### Mole concept

One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g of the C-12 isotope. The number of particles in one mole of substance is called **Avogadro number** ( $N_0$ ) and is equal to  $6.022 \times 10^{23}$ . The mass of one mole of a substance in grams is called its **molar mass**.

$$1 \text{ mole} = 6.022 \times 10^{23} \text{ particles}$$

- mass in gram of 1 amu =  $(1/N_0)$  g
- 1 mole of atoms = Gram atomic mass (or 1 g atom) =  $N_0$  atoms
- 1 mole of molecules = Gram molecular mass (or 1 g molecule) =  $N_0$  molecules
- 1 mole of ionic compound = Gram formula mass =  $N_0$  formula units

The volume occupied by 1 mol of a substance is called **molar volume**. Molar volume of a substance depends on the temperature and pressure of the substance.

*Molar volume of gas molecules is 22.4 L at STP (273 K and 1 atm. pressure)*

**Example:** What is the weight of one atom of oxygen?

**Soln:** Weight of one mole of Oxygen = 16 g

Weight of one atom of Oxygen =  $16/N_0 = 16/(6.022 \times 10^{23}) = 2.66 \times 10^{-23}$  g

**Example:** 20 cc of CO<sub>2</sub> are passed over red hot coke. The volume of CO evolved is (reaction is C + CO<sub>2</sub> → 2CO)

**Soln:** 1 cc of CO<sub>2</sub> gives 2 cc of CO, 20 cc of CO<sub>2</sub> gives  $2 \times 20 = 40$  cc of CO

Find the no. of atoms in:

- (a) 3 mole of Na
- (b) 0.04 mole of C

Find the weight of one atom of:

- (a) Carbon
- (b) Nitrogen
- (c) Oxygen
- (d) Hydrogen

Ans: (  $12/N_0$ ,  $14/N_0$ ,  $16/N_0$ ,  $1/N_0$  )

Applications of Avogadro law:

Moles of gas = (Volume in litre at STP) / 22.4

**Example:** Carbon dioxide contains 27.27 % by weight of carbon. With what weight of oxygen will 24 g of carbon combine? Which law of combination is required here?

**Soln:** 27.27 g of carbon combines with 72.73 g of oxygen.

Hence 24 g of carbon combines with  $(72.73/27.27) \times 24 = 64$  g of oxygen.

This is in accordance with law of definite proportions

**Example:** Find the number of molecules in 24 g of oxygen.

**Soln:** Molecular weight of oxygen = 32

Number of moles of oxygen = weight of oxygen / molecular weight of oxygen =  $24/32 = 0.75$

No. of molecules of oxygen =  $0.75 \times 6.02 \times 10^{23}$

Example: Calculate the volume at STP occupied by 99 g of CO<sub>2</sub>.

Soln: Moles of CO<sub>2</sub> = 99/44=2.25

Volume of 1 mol at STP = 22.4 L

Volume of 2.25 mol = 22.4\*2.25 = 50.4 L

## Chemical Formula

The formula that gives the symbolic representation of the actual number of atoms of various elements present in one molecule of the compound is called the *molecular formula*.

*Empirical formula* is defined as the simplest formula of the substance, which gives the relative number of atoms of each element present in the molecule of that substance. This is also called as stoichiometric formula. It gives the simplest whole number ratio between the numbers of atoms of all the elements present in the compound.

C<sub>6</sub>H<sub>12</sub>O<sub>6</sub> is the molecular formula of glucose, while CH<sub>2</sub>O is the empirical formula of glucose.

Empirical formula can be determined if the percentage composition of each constituent elements present in the compound are known.

Percentage composition is defined as the mass percentage of each constituent element present in any compound

$$\text{Mass \% of the element} = \frac{\text{mass of element in 1 molecule of the compound}}{\text{Molecular mass of the compound}} \times 100$$

### Relationship between Molecular formula and empirical formula

Molecular formula = n × empirical formula

Where, n is a whole number

## Chemical Equation

A chemical equation represents a chemical reaction which actually takes place. The substances that react among themselves to bring about the chemical changes are known as reactants, whereas the substances that are produced as a result of the chemical change are known as products.

A chemical equation should be balanced before taking out any information from them.

## Balanced chemical equation

It is in accordance with the law of conservation of mass. A balanced chemical equation has the same number of atoms of element on both sides of equation.

## Balancing of chemical equation

Most of the chemical equations are balanced by hit and trial method. See *Example* for balancing of a chemical equation. There is a more scientific way, but that we will study later.

Example: Write the chemical reaction for combustion of butane ( $C_4H_{10}$ ) and balance it.

Soln:  $C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

First balance the C atoms, it becomes  $C_4H_{10}(g) + O_2(g) \rightarrow 4CO_2(g) + H_2O(l)$

Now balance H atoms, it becomes  $C_4H_{10}(g) + O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

Now balance O atoms, it becomes  $C_4H_{10}(g) + (13/2)O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

Multiply whole by 2, it becomes  $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$

## Limiting reagent

When reactants are not present in the amounts as required by a balanced chemical reaction, some amount of one reactant (which is in excess) will remain unreacted. The reactant which gets consumed fully is called limiting reagent.

It is like if you are set to manufacture a cycle and make 500 tyres and 200 chains, the chains will limit the cycles that can be made. After 200 cycles, 100 tyres will remain unused.

Example: 60 g of Magnesium and 60 g of oxygen are reacted, then the residual mixture contains

Soln:  $2Mg + O_2 \rightarrow 2MgO$

48 g of Mg reacts with 32 g of oxygen to give 80 g of MgO

60 g of Mg reacts with 40 g of oxygen to give 100 g of MgO

So the residual mixture contains 100 g of MgO and 20 g of  $O_2$

## Molarity

Most of the reactions carried out are in solution. Therefore it is important to know how the concentration of a substance is expressed in solutions.

$$\text{Molarity (M)} = \frac{\text{No. of moles of solute}}{\text{volume of solution in litres}}$$

Example 13: What is the concentration of sugar ( $C_{12}H_{22}O_{11}$ ) in  $\text{mol L}^{-1}$  if its 68.4 g are dissolved in enough water to make a final volume upto 2 L?

Soln: Molecular wt. of sugar=342, moles of sugar=68.4/342=0.2

Molarity = moles/volume=0.2/2= 0.1 M

## Assignments

### Try these Questions

- A balanced chemical equation is in accordance with
  - Avogadro's law
  - Law of constant proportions
  - Law of gaseous volumes
  - Law of conservation of mass
- Which of the following has the largest number of atoms?
  - 0.5 g atom of Na
  - 0.23 g of Na
  - 0.25 mole of Na
  - $2.3 \times 10^{20}$  amu of Na
- Which has maximum molecules?
  - 7 g  $N_2$
  - 2 g  $H_2$
  - 16 g  $NO_2$
  - 16 g  $O_2$
- 2 grams mixture of  $H_2$  and He gases have mole fraction of  $H_2$  as 0.5. Find the moles of He.(mole fraction=no. of moles/total moles)
- 5.6 litres of a gas at STP are found to have a mass of 7 g. The molecular mass of the gas is
- Using trial and error, balance the equations
  - $C_2H_6 + O_2 \rightarrow CO_2 + H_2O$
  - $Fe_2O_3 + H_2SO_4 \rightarrow Fe_2(SO_4)_3 + H_2O$
  - $CO_2 + H_2 \rightarrow CH_4 + H_2O$
- What is an isotope?
  - Carbon has 2 major isotopes.  $C^{12}$  and  $C^{13}$ , each in ratio 185:1. Find the atomic weight of carbon.
  - Find the atomic number of carbon.
- $N_2 + O_2 \rightarrow N_2O_5$  Weight of  $N_2O_5$  formed from 5.6 gm of  $N_2$  and 14.4 gram of oxygen (in gm)
- Balance the equation (Using trial and error)  $Fe_2(SO_4)_3 + KSCN \rightarrow K_3Fe(SCN)_6 + K_2SO_4$
- How many moles are there in one atom?
- How many grams of each of the following elements must be taken to get 1 mol of the element?
  - Na
  - Chlorine
  - Copper
- What is the mass in g of
  - Single atom of chlorine
  - Single molecule of  $CO_2$ ?
- Calculate the number of constituent atoms in 53 g of  $Na_2CO_3$ .

(Masses Na = 23u, C = 12u, O = 16u).

14.  $10^{22}$  atoms of an element X are found to have a mass of 930mg. calculate the molar mass of the element X.
15. How many grams of Na will have the same number of atoms as 6g of Mg ( Na = 23, Mg = 24).
16. The density of liquid is  $13.6 \text{ g/cm}^3$ . How many moles of Hg are there in 1 litre of the metal (atomic mass of Hg = 20u).
17. Calculate the molecules of methane, C and H atoms in 25g of methane.
18. If a dot of a lead pencil has a mass of  $1 \mu\text{g}$ . Calculate the no. of C atoms needed to make such a dot.
19. Calculate the following quantities in 11.2g of nitrogen. Atomic mass of N = 14amu.
- No. of moles of  $\text{N}_2$ .
  - No. of molecules of  $\text{N}_2$ .
  - No. of atoms of Nitrogen.
  - Volume occupied by the gas under NTP conditions.
20. Calculate the volume of 12g of oxygen under NTP conditions.
21. Given 0.7 mol of  $\text{S}_8$ . calculate
- Number of molecules.
  - Number of S atoms.
  - Number of moles of S atoms.
22. How many (i) grams (ii) molecules (iii) atoms of hydrogen are there in 1 mole of  $\text{H}_2$ ?
23. Calculate the no. of moles of  $\text{CO}_2$  in 5.6 litre of the gas at 273K & 1 atm pressure.
24. Determine the mass of
- 0.5 mole of  $\text{O}_2$  gas.
  - 0.5 mole of O atoms.
25. How many atoms of each of the constituent elements are there in 100g of  $\text{CaCO}_3$ ? (Ca =40, C=12, O =16).
26. What is the mass of  $10^{-3}$  mol of glucose? How many C atoms are there in  $10^{-3}$  mol of glucose?