I. Fundamental concepts

1. Structure of Matter

The things in our environment can be classified into two main categories as matter and energy. Those that occupy space and have a mass are called matter. The classification of matter according to their physical nature and chemical composition is shown in the following chart.

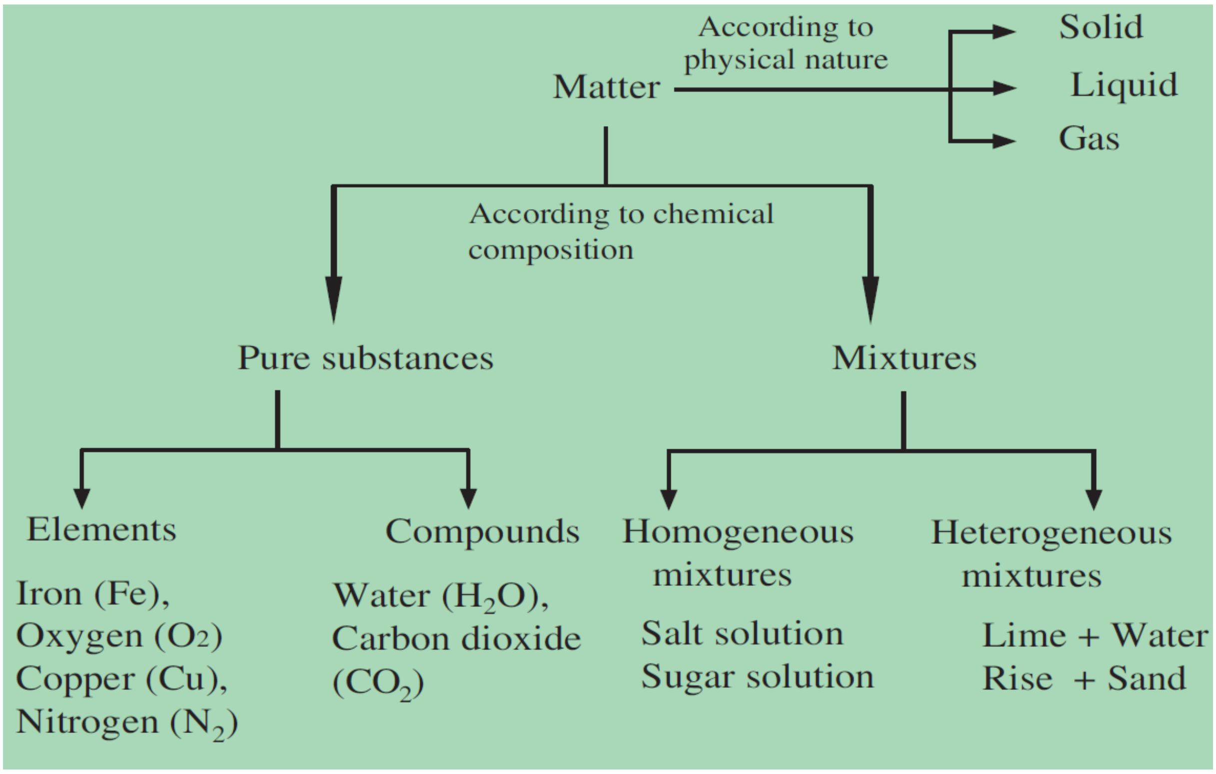

Figure 1: Sub-divisions of Matter.

2. Change of State of Matter

A physical change in a matter is referred to as a change of condition. They are reversible alterations that do not entail any modifications to the matter's chemical composition. Deposition, melting, sublimation, freezing, vaporization and condensation are examples of state transitions. The modifications are depicted in the diagram below.

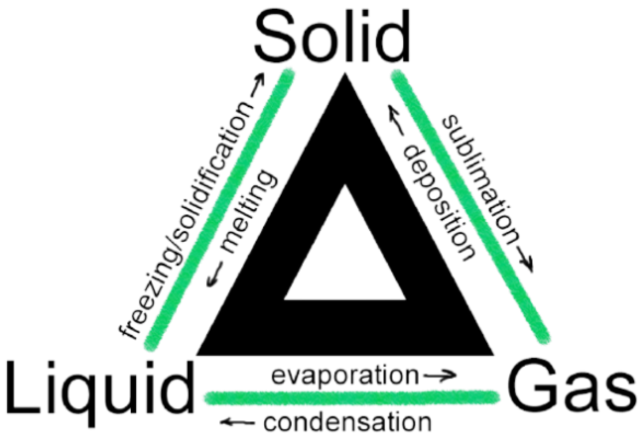

Figure 2: Transformation of matter

3. Notions of atom, isotope and molecule

3.1 Atom

An atom is a particle of matter that uniquely defines a chemical element. An atom consists of a central nucleus that is surrounded by one or more negatively charged electrons. The nucleus is positively charged and contains one or more relatively heavy particles known as protons and neutrons. (Figure I.7).

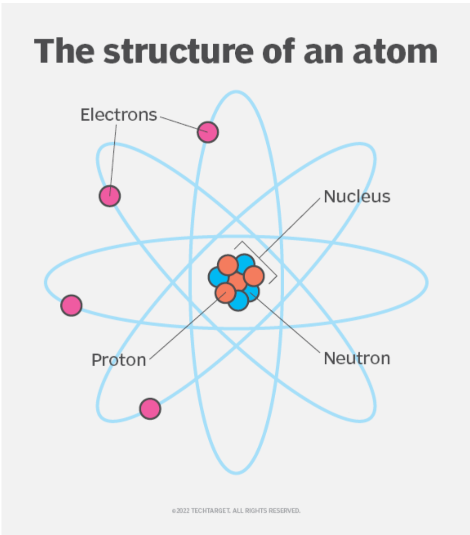

Figure 3: Structure of atom

Protons and electrons have equal and opposite charges. Protons have a positive charge and electrons a negative charge. Normally, atoms have equal numbers of protons and electrons, giving them a neutral charge. The atomic number (represented by the letter Z) of an element is the number of protons in the nucleus of each atom of that element.

The mass number (represented by the letter A) is defined as the total number of protons and neutrons in an atom

Figure 4: Atomic symbol

Example 1

Calculate each of the three subatomic particles and give specific group or period names for each atom.

a. mercury
b. platinum
c. bromine

Solutions

a. Hg (transition metal)- has 80 electrons, 80 protons, and 121 neutrons
b. Pt (transition metal)- has 78 electrons, 78 protons, and 117 neutrons
c. Br (halogen)- has 35 electrons, 35 protons, and 45 neutrons

Example 2

Write both A/Z and symbol-mass formats for the atoms in Example 1

Solutions

a. $^{201}_{80}Hg$
b. $^{195}_{78}Pt$
c. $^{80}_{35}Br$

Example 3

Identify the elements based on the statements below.

a. Which element has 25 protons?
b. Which element has 0 neutrons?
c. Which element has 83 electrons?

Solutions

a. manganese

b. hydrogen

c. bismuth

3.2 Isotope

Isotopes are atoms that have the same number of protons but different numbers of neutrons. The atomic masses of isotopes differ. For example, hydrogen has three isotopes: protium (1H), deuterium (2H), and tritium (3H), with one, two, and three neutrons respectively.

- 1 proton and without neutron $\quad {}_{1}^{1}H$
- 1 proton and 1 neutron $\quad {}_{1}^{2}H$
- 1 proton et 2 neutrons $\quad {}_{1}^{3}H$

The percentage of atoms with a specific atomic mass found in a naturally occurring sample of an element is known as its relative abundance (x_i).

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$$x_i\,(\%) = \frac{number\ of\ atoms\ in\ a\ given\ isotope}{Total\ number\ of\ atoms\ of\ all\ isotopes\ of\ this\ element} \times 100$$

Because an element is composed of a mixture of various isotopes with constant proportions, it becomes possible to define an average molar mass for each element, taking its composition into account:

$$\mathrm{M} = \sum \mathrm{x_i} \times \mathrm{M_i}$$

x_i designating the natural abundance of the isotope i of molar mass M_i.

Example

Natural copper is composed of two stable isotopes with respective atomic masses of 62.929 (A1 = 63) and 64.927 (A2 = 65). The atomic number of copper is Z=29.

Indicate the composition of the two isotopes.

Knowing that the molar mass of the natural isotopic mixture is 63.540, calculate the abundance of the two isotopes.

Cu: Z = 29

Isotope 1: M1 = 62.929 g mol^{-1}

29 protons; 29 electrons and 34 neutrons

Isotope 2: M2 = 64.927 g mol^{-1}

29 protons; 29 electrons and 36 neutrons

$$\mathrm{M} = \sum \mathrm{x_i} \times \mathrm{M_i}$$

$M_{Cu} = x_1 M_1 + x_2 M_2$

$\Sigma xi = 1 \Rightarrow x_1 + x_2 = 1 \Rightarrow x2 = 1 - x_1$

$M_{Cu} = x_1 M_1 + (1 - x_1) M_2 = x_1 M_1 + M_2 + x_1 M_2$

$M_{Cu} - M_2 = x_1 (M_1 - M_2)$

$x_1 = (M_{Cu} - M_2) / (M_1 - M_2) = (63.540 – 64.927) / (62.929 – 64.927) = 0.6942$

$$^{63}_{29}Cu = 69.42\%$$

$$^{65}_{29}Cu = 30.58\%$$

3.3 Molecules

Molecules are made when two or more atoms chemically bond together, represented by $C_xH_yO_z$

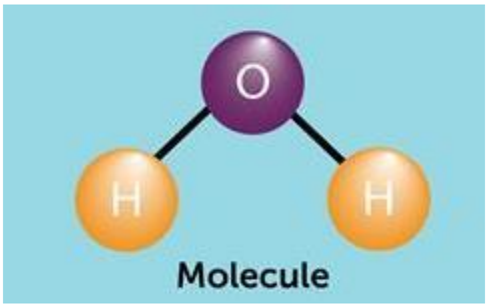


Figure 5: A schematic representation of the water molecule

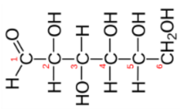
For example

- A hydrogen chloride molecule is composed of one hydrogen atom and one chlorine atom, so its formula is HCl.
- the water molecule, shown above, is composed of two hydrogen atoms (subscript 2) and one oxygen atom (subscript 1 is not written), its formula is H2O

Any molecule can be represented as a molecular formula, a structural formula, and an empiric formula. It is important to be aware that it may be possible for the same atoms to be arranged in different ways: Compounds with the same molecular formula may have different atom-to-atom bonding and therefore different structures.

Example: A glucose molecule can be represented as (a) a molecular formula, (b) a structural formula, and (c) an empiric formula, respectively.

$C_6H_{12}O_6$



$(CH_2O)_n$

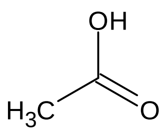
(a) a molecular formula (b) a structural formula c) an empiric formula

Example:

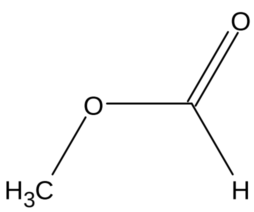
Could there be another compound with the same formula as acetic acid, $C_2H_4O_2$

Solution

Methyl formate molecules have one of the oxygen atoms between the two carbon atoms, differing from the arrangement in acetic acid molecules.



acetic acid



methyl formate

Example:

Molecules of glucose (blood sugar) contain 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms. What are the molecular and empirical formulas of glucose?

Solution

The molecular formula is $C_6H_{12}O_6$ because one molecule actually contains 6 C, 12 H, and 6 O atoms. The simplest whole-number ratio of C to H to O atoms in glucose is 1:2:1, so the empirical formula is CH_2O.

Example:

Dichloromethane has a molecular weight of 99 u. Analysis of a sample shows that it contains 24.3% carbon, 4.1% hydrogen and 71.6% chlorine. What is its molecular formula?

Solution:

Assume 100 g of compound. We have 71.6 g of Cl, 24.3 g of C, and 4,1 g of H. Look up the molar masses of Cl, C, and H and convert these to moles.

$24.3/12 = 2.1$ mol

$4.1/1 = 4.1$ mol

$71.6/35.5 = 2.1$ mol

Dichloromethane has a molecular weight of 99 u , by dividing it by the empirical formula mass we find the value of n.

n = 99/49.5 = 2

So the molecular formula of dichloromethane is $C_2H_4Cl_2$

4. Law of Conservation of Mass

"Nothing is lost, nothing is created, everything is transformed" (Lavoisier 1743-1794)

According to this law, during any physical or chemical change, the total mass of the products remains equal to the total mass of the reactants, as described in the examples below.

- When Matter Undergoes a Chemical Change

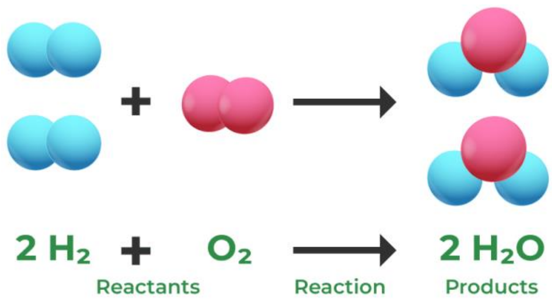


Figure 6: Chemical change of matter

- When Matter Undergoes a Physical Change

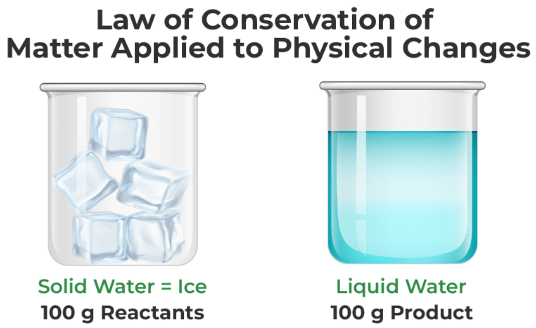


Figure 7: Physical change of matter

Example

If heating 10 grams of CaCO3 produces 4.4 g of CO2 and 5.6 g of CaO, show that these observations are in agreement with the law of conservation of mass.

Solution

Mass of the reactants: 10g

Mass of the products: 4.4g+5.6g=10g

Because the mass of the reactants is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

5. Qualitative aspects of the matter

5.1 Matter composition

The state of matter can be an element, a compound or a mixture.

5.1.1 Element

An element is matter made of only one kind of atom.

• There are 115 known elements.

• Ninety elements are naturally occurring.

• The elements are organized according to their properties in the Periodic Table.

• Examples – Hydrogen, Carbon, Nitrogen, Calcium, Sodium,Oxygen

5.1.2 Compounds

Compounds are 2 or more elements that are chemically combined.

 Compounds cannot be easily separated into their elements.

Examples;

1. H2O Water

2. NaCl Salt

3. $C_6H_{12}O_6$ Sugar/Glucose

H_2 N_2 O_2

 The gases of hydrogen, nitrogen and oxygen naturally exist as compounds of 2 atoms of their element

5.1.3 Mixtures

Mixtures are made of different compounds that are mixed together.

Mixtures can be easily separated into the original compounds.

Homogeneous – substances evenly mixed

Heterogeneous – substances not evenly mixed

5.1 Solutions

A solution is a mixture of two (or more) substances, consisting of a **solvent**, which is usually present in large quantities, and a **solute** (solid, liquid or gas), which is usually present in smaller quantities

Dilute solution: contains a low concentration of solute

Concentrated solution contains a high concentration of solute

Saturated solution: When the introduced solute can no longer dissolve in the solvent and forms a precipitate.

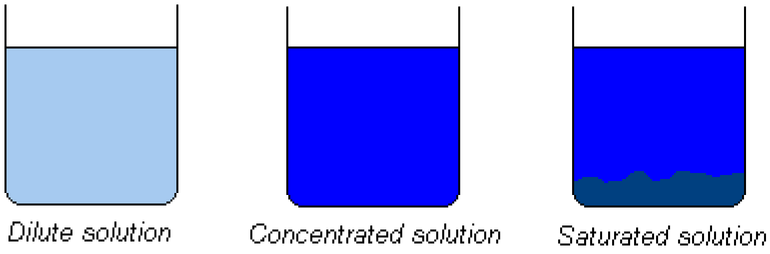


Figure 8: Representative schematic of a dilute, concentrated and saturated solution

6. Quantitative aspects of the matter

6.1 Mole number (n)

The mole (abbreviated mol) is the SI measure of quantity of a "chemical entity," such as atoms, electrons, or protons. It is defined as the amount of a substance that contains as many particles as there are atoms in 12 grams of pure carbon-12.

The number of units in one mole of any substance is called Avogadro's number or Avogadro's constant It is equal to **$6.023\ 10^{23}$ mol^{-1}.**

The mole number (n) is calculated:

- *For a liquid or solid* $n = \frac{m}{M}$

m: mass of compound in g

M: molar mass of compound in $g{\cdot}mol^{-1}$

Example:

Calculate the number of moles of hydrated copper(II) sulfate (CuSO4, 5H2O) contained in 5.00 g of this substance.

Solution:

$M_{(CuSO4\ \cdot\ 5H2O)}$ = (63.5 + 32 + 4 x 16 + 5 x 18) = 249.5 g/mol

The amount of copper sulphate in moles.

$$n = \frac{m_{(CuSO4\ \cdot\ 5H2O)}}{M_{(CuSO4\ \cdot\ 5H2O)}} = \frac{5}{249{,}5} = 0.02\ \text{mol}$$

- *For a gaz*

$$n = \frac{V}{V_m}$$

V: gas volume in L (liters)

Vm: molar volume of compound in $L \cdot mol^{-1}$

6.2 Atomic mass unit (amu)

The atomic mass unit is defined as the mass of 1/12 of the carbon-12 atom.

$1\ u = 1{,}661.10^{-27}$ kg

Example: m_e=0.000548 u, m_p=1.00728 u, m_n= 1.008 u

6.2 Molar mass (M)

The molar mass (M) of a chemical compound is defined as the mass of one mole of a substance (a simple body, a chemical compound). It is expressed in grams per mole (g mol^{-1} or g mol^{-1}).

$$M = \frac{m}{n}$$

Example: M_H = 1.0 g mol^{-1}; M_C = 12.0 g mol^{-1} ; M_O = 16.0 g mol^{-1}.

6.3 Percent composition

In a chemical compound, the elements of which it is composed are always present in the same mass ratio. The percentage composition can be calculated using the following formula:

$$x(\%) = \frac{m_x}{m_T} \times 100$$

Where: m_x corresponds to the elementary mass and m_t to the total mass.

Example 1

Calculate the nitrogen mass percentage in ammonium nitrate NH_4NO_3

m(N) = 14 u we have two nitrogen atoms => m(N) = 28

$m(NH_4NO_3)$ = 14 + 4 x 1 + 14 + 16 x 3 = 80 u

$$N(\%) = \frac{m_N}{m_T} \times 100$$

$$N(\%) = \frac{28}{80} \times 100 = 35\%$$

6.4 Molar concentration (molarity)

The molar concentration, or molarity, or sometimes molar rate of a chemical species, is the quantity of matter, expressed in moles per unit volume (1); this concept is mainly used for species in solution. The molar concentration of a species X is denoted cX or [X].

$$c_x = \frac{moles\ of\ solutes}{solution\ volume} = \frac{n}{V}$$

In the same field, the molality is the quantity of matter, expressed in moles per unit mass (mol kg^{-1}). Compared with molarity, the number of moles is replaced by the mass.

$$b = \frac{moles\ of\ solutes}{solution\ mass} = \frac{n}{m}$$

Example:

We dissolve 90 g of glucose (C6H12O6) in 1 liter of water. What are the molarity and molality?

Solution:

$M(C_6H_{12}O_6)$ = (6 x 12 + 12 + 6 x 16) = 180 g/mol.

$$\mathrm{n(C6H12O6)} = \frac{\mathrm{m_{(C6H12O6)}}}{\mathrm{M_{(C6H12O6)}}} = \frac{90}{180} = 0{,}5\ \mathrm{mol}$$

$$\mathrm{molarity}\ (\mathrm{c_x}) = \frac{\mathrm{n(\,glucose)}}{\mathrm{v\ (solution)}} = \frac{0{,}5}{1} = 0{,}5\,\frac{\mathrm{mol}}{\mathrm{l}}\ \text{or 0.5 molar.}$$

In the case of molality, the mass of a liter of water is 1 kg, so:

$$\mathrm{molality} = \frac{\mathrm{n(\,glucose)}}{\mathrm{m\ (solution)}} = \frac{0{,}5}{1} = 0{,}5\ \mathrm{mol/kg}$$

6.5 Molar volume

The molar volume of a substance is the volume occupied by one mole of that substance under specific conditions of temperature and pressure.

Molar volume can be described as:

$$V_m = \frac{V}{n} \qquad (m^3{\cdot}mol^{-1})$$

V: volume in cubic meters or liters;

n: quantity of substance in moles.

Example:

What volume of carbon dioxide can be obtained by decomposing 50 g of calcium carbonate under normal conditions of pressure and temperature (0°C and 1atm)?

Solution:

We write the equilibrium equation for the reaction:

$$CaCO3(s) \rightarrow CaO(s) + CO2(g)$$

First, we calculate the number of moles of CaCO3 involved in this reaction:

$$n = \frac{m}{M}$$

$m_{CaCO3} = 50g$ and $M_{CaCO3} = 40 + 12 + 16x3 = 100$ g/mol ⇒ $n = \frac{50}{100} = 0.5$ mol

We have 1 molecule of CaCO3 which gives 1 molecule of CO2

So 1 mol CaCO3 gives 1 mol CO2

and 0.5 mol CaCO3 induces 0.5 mol CO2

We know that one mole of gaseous substance occupies a volume of 22.4L, i.e. 0.5 mol of CO2 occupies a volume equal to 0.5 x 22.4= 11.2 l.

6.7 Mass concentration

The mass concentration of a solution corresponds to the mass of solute (g) dissolved or diluted in 1 L of solvent. This parameter is used for mixtures.

$$Cp = \frac{m_i}{V_{sol}}$$

Example:

Dissolve 0.02 mol of glucose (C6H12O6) in 1 liter of water. What is the mass concentration?

Solution:

$$\left(m_{(C6H12O6)}\right) = n_{(C6H12O6)} \times M_{(C6H12O6)} = 0.02 \times 249.5 = 5 \text{ g}$$

So the mass concentration is $Cp = \frac{5}{1} = 5\ g\ L^{-1}$

6.8 Molar fraction

It represents the ratio of the number of moles of a specific component (n_i) to the total number of moles in the mixture (n_t). The formula for calculating the mole fraction (X_i) of a component in a mixture is:

$$x_i = \frac{n_i}{n_{tot}}$$

The mole fraction is between 0 and 1, and the sum of mole fractions is 1.

We can use the mass fraction (w_i), which is the ratio between the mass (m_i) and the total mass (m_{tot}) of the mixture.

$$w_i = \frac{m_i}{m_{tot}}$$

Example:

Calculate the mass and mole fraction of oxygen (0.5 mol) mixed with nitrogen (0.7 mol).

Solution:

We begin by calculating the mass of oxygen and nitrogen:

$$m_{O2} = n_{O2} \times M_{O2} = 0.5 \times 32 = 16\ g$$

$$m_{N2} = n_{N2} \times M_{N2} = 0.7 \times 28 = 19.6\ g$$

Mass fraction of oxygen: $w_{O2} = \frac{16}{35.6}$

Molar fraction of oxygen: $x_{O2} = \frac{0,5}{1.2} = 0.41$

6.9 Normality (N)

Normality is the number of equivalents (Z) of solute per liter of solution (mol/l), it describes only the moles of reactive species per liter of solution, it is always a multiple of molarity and it describes the molar equivalents of the reagents involved in chemical reactions.

Normality can be calculated using the following formula

$$C_n = M*Z$$

C_n : Normality

M : molarity

Z: the number of protons or electrons exchanged in the reaction

For example a solution of H2SO4 worth 0.1 molar (0.1 M) implies a normality equal to 0.2 N, as there are 2 times more moles of H+ ions in solution than moles of H2SO4 introduced. A solution of HCl or NaOH has the same normality as its molarity:

$$H_2SO_4 \rightarrow 2H + SO_4^{2-}$$

$$NaOH \rightarrow H + OH^-$$

And in its oxidation reaction, permanganate (KMnO4) exchanges 5é, so if its molarity is 0.01 molar its normality is therefore equal to 0.05 normal:

$$MnO_4^- + 8\ H^+ + 5\ e^- \rightarrow Mn^{2+} + 4\ H_2O$$

6.10 Volumetric mass density or specific mass

Is a physical quantity that characterizes the mass of a material per unit volume (kg/cm^3).

$$\rho = \frac{m}{V}$$

Example

The mass of mercury contained in 54 cm^3 of volume equals 0.7344 kg. What is the density of mercury?

Solution:

Applying the relationship ρ=m/V we obtain a density equal to 0.0136 kg cm^{3-}

6.11 Density

The density of a solid or liquid is the ratio of its density to the density of a reference body. For liquids and solids, the reference body is pure water at 4°C.

$$d = \frac{\rho_{(s\ ou\ l)}}{\rho_{H2O}} = \frac{\frac{m}{v}}{\frac{m'}{v}} = \frac{m}{m'}$$

The density of a gas is the ratio of its density to the density of air under the same conditions of temperature and pressure.

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$$d_{gaz} = \frac{\rho_{gaz}}{\rho_{air}} = \frac{M_{gaz}}{M_{air}} = \frac{M_{gaz}}{28}$$

Example

Knowing that air is composed of approximately 80% nitrogen and 20% oxygen, determine the molar mass of air

$\mathrm{M_{air}} = 0.8 * \mathrm{M_{N2}} + 0.2 * \mathrm{M_{O2}} = 28.8\ \mathrm{g\ mol^{-1}}$

7. Application exercises

Exercise 1

1. Calculate the amount of substance in 5.0 g of silver, with a molar mass of 107.9 g/mol.
2. Calculate the number of silver atoms present in these 5.0 g.
3. Deduce the mass of a silver atom.
4. Calculate the mass of a silver atom knowing that the studied silver is: (_47^107)Ag. Compare this value with question 3.

Given: $m_{proton} = 1.67 \times 10^{-27}$ kg (mass of electrons is negligible in this exercise and mass of a neutron = mass of a proton).

Exercise 2

Calculate the molar mass of the following compounds: carbon dioxide, sucrose, ammonia, methane, hydrogen chloride, sulfur dioxide.

Exercise 3

A cylinder weighs 37.6 g when empty; it weighs 53.2 g when filled up to 7.4 ml with an unknown liquid. Calculate the density of the liquid.

Exercise 4

A cylindrical rod weighing 45.0 g measures 2.00 cm in diameter and 15.0 cm in length (L). Find its density.

Exercise 5

Calculate the mass required to prepare 250 ml of a solution of AgNO3 with a concentration of 0.125 M, using solid AgNO3.

Exercise 6

What volume of 0.25 M Na2CrO4 will be needed to obtain 8.1 g of Na2CrO4?

Exercise 7

Calculate the mass required to prepare a solution with a mole fraction of sucrose (C12H22O11) at 0.0348, using 100 g of water.

Exercise 8

Concentrated sulfuric acid is labeled with a density of d = 1840 g/l and is 96% pure by weight. Calculate the molarity of this solution.

Exercise 9

What volume of 18.0 M H2SO4 is needed for the preparation of 2.00 liters of 3.00 M H2SO4?

Exercise 10

Calculate the relative atomic mass (RAM) of boron from the following data:

Isotope	Isotopic Mass (u)	Abundance (%)
10B	10.0129	19.91
11B	11.0093	80.09

Exercise 11

When a sample of aluminum (Al) is placed in a graduated cylinder with a volume of 25 ml containing 10.5 ml of water, the water level rises to 13.5 ml. What is the mass of the aluminum? dAl = 2.7 g/mL.

Exercise 12

What is the mass percentage of NaHCO3 in a solution containing 20 g of NaHCO3 dissolved in 600 mL of H2O?

Exercise 13

What is the volume percentage of ethanol in a solution containing 35 ml of ethanol dissolved in 155 ml of water?

Exercise 14

What is the molality of a solution containing 16.3 g of potassium chloride dissolved in 845 g of water?

Exercise 15

A piece of magnesium burns in the presence of oxygen (O2), forming magnesium oxide (MgO), according to the following equation:

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

How many moles of oxygen are needed to produce 12 moles of magnesium oxide?