

CHEMICAL ENGINEERING

Process Calculations



Comprehensive Theory
with Solved Examples and Practice Questions





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Process Calculations

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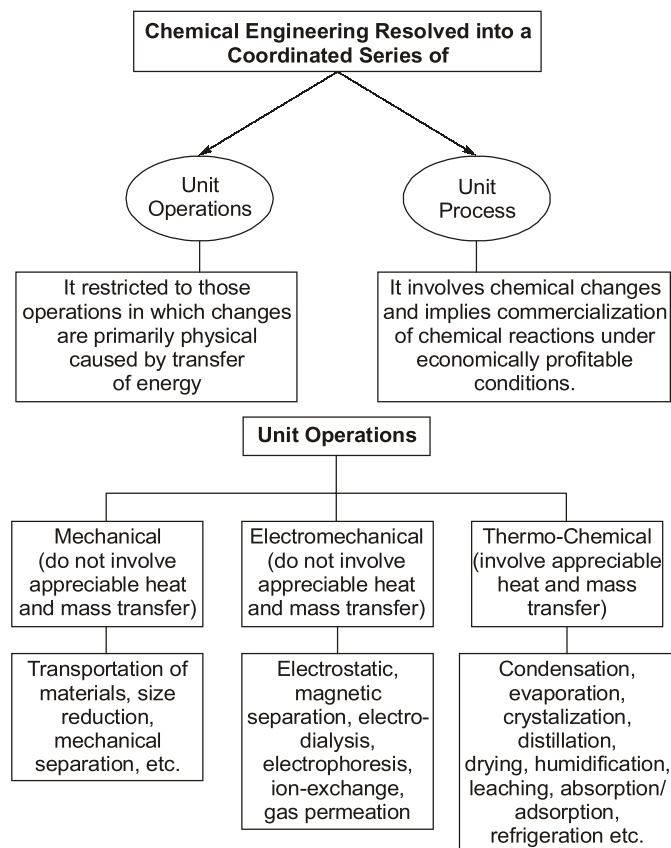
Introduction and Basic Concepts

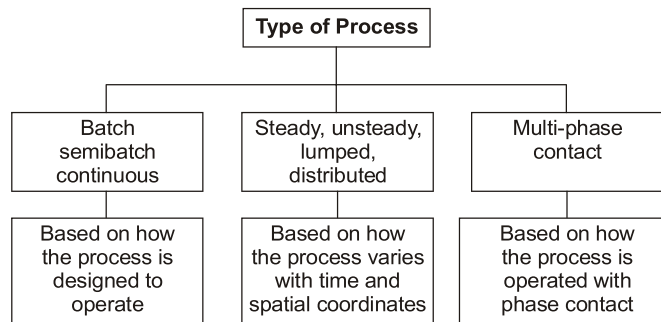
LEARNING OBJECTIVES

The reading of this chapter will enable the students

- To understand unit and dimensions.
- To understand the basic chemical calculations.
- To understand about ideal gas law and their applications.
- To understand about Raoult's and Henry's law for gas-liquid system.

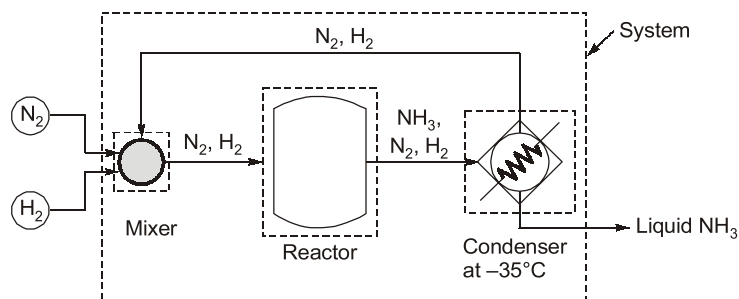
1.1 INTRODUCTION





What is System?

- Any specified arbitrary portion or whole of a process analyzing the problem is defined as system.
- It depends on what information is provided and what needs to be determined.
- A system may contain more than one process unit.
- In figure, the entire process is a system which consists of three process units.



1.2 UNIT AND DIMENSIONS

Unit	Dimensions
<ul style="list-style-type: none"> • The "unit" indicates what the measured quantity represents. 	<ul style="list-style-type: none"> • The "dimension" is the measurable quantity that the unit represents. Example : length, mass, time and temperature
<ul style="list-style-type: none"> • A measured or counted quantity has a numerical value and a unit. 	<ul style="list-style-type: none"> • It also be calculated by multiplying or dividing other dimensions. Example : length/time = velocity, length³ = volume and mass/length³ = density
<ul style="list-style-type: none"> • Measurable units are specific values of dimensions that have been defined by convention. Example: grams for mass, seconds for time and centimeters for length 	

The system of units composed of (1) Base units, (2) Derived units and (3) Multiple units.

1. **Base units:** These are the units for basic quantities such as length, mass, time, etc.
2. **Derived units:** These are the units obtained by multiplying and dividing base units. e.g., cm², m/s, etc.

3. **Multiple units:** These are the units which are multiple or fractions of base units. e.g., hour, minute, second, etc.

Unit Systems

The various systems of units and the basic/fundamental quantities associated with them are given below:

Fundamental Quantity	Systems of Units				Dimensions
	SI	MKS	CGS	FPS	
Length	Meter (m)	Meter (m)	Centimeter (cm)	Foot (ft)	L
Mass	Kilogram (kg)	Kilogram (kg)	Gram (g)	Pound (lb)	M
Time	Second (s)	Second (s)	Second (s)	Second (s)	θ
Temperature	Kelvin (K)	Celsius (°C)	Celsius (°C)	Fahrenheit (°F)	T

°C = Degrees Celsius

K = Kelvin

SI = International system of units

Symbolic abbreviations of the units are given in brackets.

- (i) **Force:** According to Newton’s law of motion, force is proportional to the product of mass and acceleration.

$$F \propto m.a$$

$$F = k m.a$$

In CGS system, dyne is defined as the force necessary to accelerate one gram mass at 1 cm/s².

In the SI system, Newton (N) is defined as the force necessary to accelerate one kilogram mass at 1 m/s². The engineer’s unit of force in the MKS system is kilogram-force (kgf). The kilogram force is the force necessary to accelerate 1 kg mass at 9.81 m/s².

In the SI system, the unit of force has been named as Newton (N) in honour of the scientist Newton. 1 N is equal to 1 (kg.m)/s².

- (ii) **Kilogram force (kgf):** The **kilogram force** is a metric unit of force (kgf). The kilogram-force is equal to a mass of one kilogram multiplied by the standard acceleration due to gravity on Earth, which is defined as exactly 9.80665 meter per second². Then one (1) kilogram-force is equal to 1 kg × 9.80665 meter per second² = 9.80665 kilogram × meter per second² = 9.80665 newton (1 N).

Note: A kilogram-force (kgf), also called kilopond (kp), is a gravitational metric unit of force.

- (iii) **Pressure:** Pressure is defined as the force per unit area.

$$P = \frac{F}{A}$$

The units of pressure in SI, MKS and FPS systems are N/m² (known as Pascal, abbreviated as Pa), kgf/cm² and lbf/in² (psi) respectively. The relationship between the absolute, atmospheric and gauge pressure is

$$\text{Absolute pressure} = \text{Gauge pressure} + \text{Atmospheric pressure}$$

Gauge pressure:

$$\text{SI unit} = \text{N/m}^2 \text{ (or Pa)}$$

CGS unit = dyn/cm²

AES unit = lbf/in² (or psi)

- The gauge pressure of the fluid which is the pressure of the fluid relative to atmospheric pressure (reference pressure).
- A gauge pressure of 0 indicates that the absolute pressure of the fluid is equal to the atmospheric pressure.

$$P_{\text{absolute}} = P_{\text{gauge}} + P_{\text{atmospheric}}$$

Types of Pressures

- **Atmospheric pressure:** P_{atm} , is the pressure caused by the weight of the earth's atmosphere. Often atmospheric pressure is called barometric pressure.
- **Absolute pressure:** P_{abs} , is the total pressure. An absolute pressure of 0.0 is a perfect vacuum. Absolute pressure must be used in all calculations, unless a pressure difference is used.
- **Gauge pressure:** P_{gauge} , is pressure relative to atmospheric pressure.
- **Vacuum pressure:** P_{vacuum} , is a gauge pressure that is below atmospheric pressure.

$$P_{\text{gauge}} = P_{\text{absolute}} - P_{\text{atmospheric}}$$

$$P_{\text{vacuum}} = P_{\text{atmospheric}} - P_{\text{absolute}}$$

$$P_{\text{absolute}} = P_{\text{atmospheric}} + P_{\text{gauge}}$$

- (iv) **Work:** Work is defined as the product of the force acting on body and the distance travelled by the body in the direction of force applied.

The SI units for work are the joule (J) or Newton meter ($N \times m$), from the function $W = F \times s$ where W is work, F is force, and s is the displacement. The joule is also the SI unit of energy.

$$\text{Joule} = N \cdot m = \text{kg} \cdot \text{m}^2/\text{s}^2$$

- (v) **Power:** Power is the time rate at which work is done or energy is transferred. In calculus terms, power is the derivative of work with respect to time.

The SI unit of power is the watt (W) or joule per second (J/s). Horsepower is a unit of power in the British system of measurement.

$$\text{Power} = \frac{\text{Work}}{\text{Time}}$$

or

$$P = \frac{W}{t}$$

- (vi) **Heat:** It is a form of energy that flows from one body to another as a result of a difference in temperature. It cannot be stored as such within the system. The units of heat in SI, MKS and CGS systems are Joule (J), kilocalorie (kcal) and calorie (cal) respectively.

Example 1.1

In a multiple effect evaporator system, the second effect is maintained under vacuum of 475 torr (mm Hg). Find the absolute pressure in kPa.

Solution:

$$\begin{aligned} \text{Absolute pressure} &= \text{Atmospheric pressure} - \text{Vacuum} \\ &= 760 - 475 = 285 \text{ torr (mm Hg)} \end{aligned}$$

$$\text{Absolute pressure} = 285 \text{ mm Hg} \times \left(\frac{101.325 \text{ kPa}}{760 \text{ mm Hg}} \right) = 38 \text{ kPa}$$

Example 1.2 A force of 19.635 kgf is applied on a piston of diameter 5 cm. Find the pressure exerted on the piston in kPa.

Solution:

$$P = \frac{F}{A}$$

$$A = \frac{\pi}{4}d^2 = \frac{\pi}{4}(5)^2$$

$$F = 19.635 \text{ kgf}$$

$$P = \frac{19.635}{19.635} = 1 \text{ kgf/cm}^2$$

$$P = 1 \text{ kgf/cm}^2 \times \frac{101.325 \text{ kPa}}{1.033227 \text{ kgf/cm}^2} = 98.066 \text{ kPa}$$

Example 1.3 Consider the equations $S = 7t + 8t^2$, where $S = m$ and $t = s$. What are the dimensions and units of 7 and 8?

Solution:

Unit of 7 is m/s and 8 is m/s².

Example 1.4 The thermal conductivity k of a liquid metal is predicted via the empirical equation $k = A \exp(B/T)$, where k is in J/(s.m.K) and A and B are constants, T is absolute temperature. What are the units of A and B ?

Solution:

A has the same unit as k i.e., J/(s.m.K), B has the unit of T i.e., K.

1.3 BASIC CHEMICAL CALCULATIONS

(i) **Atomic weight:** The atomic weight of an element is the mass of the atom of this element based on the scale that assigns carbon a mass of exactly twelve.

(ii) **Molecular weight:** The molecular weight of a compound is the sum of the atomic weights of atoms that constitute a molecule of the compound.

The molecular weight of a monoatomic element (e.g., sodium) is its atomic weight and the molecular weight of a diatomic element (e.g., oxygen, chlorine) is twice that of its atomic weight.

(iii) **Gram atom:** It is used to specify the amounts of chemical elements. It is defined as the mass in grams of an element which is numerically equal to its atomic weight.

$$\text{Gram atoms of an element} = \frac{\text{Weight in grams}}{\text{Atomic weight}} \quad \dots(1.1)$$

$$\therefore \text{Gram atoms of element } A = \frac{\text{Weight in grams of } A}{\text{Atomic weight of } A} \quad \dots(1.2)$$

Similarly, the mass in kilograms of a given element that is numerically equal to its atomic weight is called a kilogram-atom.

$$\text{Similarly, kilogram atoms of element } A = \frac{\text{Weight in kilograms of } A}{\text{Atomic weight of } A} \quad \dots(1.3)$$

For chemical compounds, a mole is defined as the amount of substance equal to its molecular weight/formula weight.

(iv) **Gram mole:** It is used to specify the amounts of chemical compounds. It is defined as the mass in grams of a substance that is equal numerically to its molecular weight.

$$\therefore \text{Gram moles of compound } B = \frac{\text{Weight in grams of } B}{\text{Molecular weight of } B} \quad \dots(1.4)$$

A gram mole of a substance is the mass in grams of the substance that is numerically equal to its molecular weight.

$$\text{Similarly, Kg. moles of compound } B = \frac{\text{Weight in kilograms of } B}{\text{Molecular weight of } B} \quad \dots(1.5)$$

A mole is defined as the amount of a substance equal to its molecular weight.

Molecular weight of a compound is found from the atomic weight of elements involved in formation of the compound.

Example 1.5

Calculate the kilogram atoms of carbon which weighs 36 kg.

Solution:

$$\text{Atomic weight of carbon} = 12$$

$$\therefore k\text{-atom of carbon} = \frac{\text{Weight in kilograms of carbon}}{\text{Atomic weight of carbon}} = \frac{36}{12} = 3$$

Example 1.6

Calculate the kilograms of Na of which the amount is specified as 3 k-atom.

Solution:

Basis: 3 k-atom Na.

$$\text{Atomic weight of Na} = 23$$

$$\therefore k\text{-atom of Na} = \frac{\text{kg of Na}}{\text{Atomic weight of Na}}$$

$$\begin{aligned} \therefore \text{kg of Na} &= k\text{-atom of Na} \times \text{Atomic weight of Na} \\ &= 3 \times 23 = 69 \end{aligned}$$

(v) **Equivalent weight of an element or a compound** is defined as the ratio of the atomic weight or molecular weight to its valence. The valence of an element or a compound depends on the number of hydroxyl ions (OH⁻) donated or the hydrogen ions (H⁺) accepted for each atomic weight or molecular weight.

$$\therefore \text{Equivalent weight} = \frac{\text{Molecular weight}}{\text{Valence}} \quad \dots(1.6)$$

The concentration of a solution containing a solid or a liquid solute can be expressed in terms of normality, molarity and molality.

(vi) **Normality** is designated by the symbol N and is defined as the number of gram-equivalent of solute dissolved in one litre of solution [for preparation of 1 N NaOH solution, we have to dissolve 40 grams of NaOH (\equiv to 1 g equivalent of NaOH) in one litre of solution (solvent-water)].

Example 1.18 How many grams of carbon dioxide gas is dissolved in a 1 L bottle of carbonated water if the manufacturer uses a pressure of 2.4 atm in the bottling process at 25°C?

Given: H of CO_2 in water = 1.65×10^3 atm/(mol fraction) at 25°C

Solution:

Mole fraction of CO_2 dissolved in water

$$x_{\text{CO}_2} = \frac{p}{H} = \frac{2.4}{(1.65 \times 10^3)} = 0.001452$$

$$\text{Moles of } \text{CO}_2 / \text{moles of water} = \frac{x_{\text{CO}_2}}{(1 - x_{\text{CO}_2})} = 0.00415$$

$$1 \text{ lit. of water} = \frac{1000}{18} = 55.55 \text{ moles of water}$$

Therefore, Mole of CO_2 dissolved = $0.00145 \times 55.55 = 0.081$ moles
= $0.081 \times 44 = 3.553$ gms

1.6 P-V-T BEHAVIOUR OF REAL GASES

Van der Waal's Equation of State: The P-V-T behaviour of actual/real gases (non-ideal gases) differ from that predicted by the ideal gas law. J.D. Van der Waals made the successful attempt to modify the ideal gas law in order to make it applicable to real gases. J.D. Van der Waals proposed the following equation of state (known as the Van der Waals equation) to explain the P-V-T behaviour of real gases (which takes into account for non-ideal behaviour of gases).

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT$$

For 1 mole, where a and b are called Van der Waals constants for a given gas and V is the molal/molar volume (m^3/kmol). The Van der Waals constants depend upon the nature of the gas and can be determined from the experimental P-V-T data.

The values of constants a and b are evaluated using the following equations :

$$a = 27 \frac{R^2 T_c^2}{64 P_c} \text{ (l}^2 \text{ MPa/mol}^2\text{) or } \frac{(\text{m}^3)^2 \cdot \text{MPa}}{(\text{kmol})^2}$$

and
$$b = \frac{RT_c}{8P_c}, \text{ l/mol or m}^3/\text{kmol}$$

where T_c and P_c are the critical temperature and the critical pressure of the gas respectively.

Critical temperature: It is the maximum temperature at which a gas can be liquefied.

Critical pressure: It is the saturation pressure corresponding to the critical temperature.

Critical volume: It is the volume occupied by a gas at critical conditions.

Note that at temperatures above the critical conditions, a gas cannot be liquefied irrespective of the pressure.

Example 1.19 Carbon dioxide weighing 1.10 kg occupies a volume of 33 litres at 300 K. Calculate the pressure using the Van der Waals equation of state.

Data: $a = 3.60 \text{ (m}^3\text{)}^2, \text{ kPa}/(\text{kmol})^2$ and $b = 4.3 \times 10^{-2} \text{ m}^3/\text{kmol}$ for CO_2

Solution:

Basis: 1.10 kg of CO_2 gas at 300 K.

The Van der Waals equation is

$$\left(P + \frac{a}{V^2}\right)(V - b) = RT$$

$$\text{Amount of } \text{CO}_2 \text{ gas} = 1.10 \text{ kg} = 0.025 \text{ kmol}$$

$$\text{Volume occupied by this gas} = 33 \text{ l} = 0.033 \text{ m}^3$$

$$\begin{aligned} \therefore V &= \frac{0.033}{0.025} = 1.32 \text{ m}^3/\text{kmol} \\ a &= 3.60 \text{ (m}^3\text{)}^2 \text{ kPa}/(\text{kmol})^2 \\ b &= 4.3 \times 10^{-2} \text{ m}^3/\text{kmol} \\ R &= 8.31451 \text{ (m}^3 \cdot \text{kPa)} / (\text{kmol} \cdot \text{K}) \\ T &= 300 \text{ K} \end{aligned}$$

$$\left[P + \frac{3.60}{(1.32)^2}\right][1.32 - 4.3 \times 10^{-2}] = 8.31451 \times 300$$

Solving we get,

$$P = 1951 \text{ kPa} = 1.951 \text{ MPa}$$

Compressibility factor: Compressibility factor (Z) is defined as

$$Z = \frac{PV}{nRT}$$

$$Z = \frac{\text{Actual volume of gas at a given temperature and pressure}}{\text{Volume of the ideal gas at the same } T \text{ and } P}$$



Student's Assignments

- Q.1** Convert a pressure of 2 atm to mm Hg.
- Q.2** Convert a volumetric flow rate of 2 m³/s to l/s.
- Q.3** Convert 88 kg of carbon dioxide into its amount in molar units.
- Q.4** Find the moles of oxygen present in 500 grams.
- Q.5** Find moles of K_2CO_3 that will contain 117 kg of K.
- Q.6** How many kilograms of carbon are present in 64 kg of methane?
- Q.7** 98 grams of sulphuric acid (H_2SO_4) are dissolved in water to prepare one litre of solution. Find normality and molarity of solution.
- Q.8** A solution of caustic soda contains 20% NaOH by weight. Taking density of the solution as 1.196 kg/l, find the normality, molarity and molality of the solution.
- Q.9** H_2SO_4 solution has a molarity of 11.24 and molality of 94. Calculate the density of solution.
- Q.10** At 298 K (25°C) the solubility of methyl bromide in methanol is 44 kg per 100 kg. Calculate : (i) the weight fraction and (ii) the mole fraction of methanol in the saturated solution.

- Q.11** Calculate the weight of sulphur dioxide in a vessel having 2 m³ volume, the pressure and temperature being 97.33 kPa and 393 K (120°C).
- Q.12** A certain quantity of gas contained in a closed vessel of volume 1 m³ at a temperature of 298 K (25°C) and pressure of 131.7 kPa is to be heated such that the pressure should not exceed 303.98 kPa. Calculate the temperature of gas attained.
- Q.13** A mixture of nitrogen and carbon dioxide at 298 K (25°C) and 101.325 has an average molecular weight of 31. What is the partial pressure of nitrogen?
- Q.14** 960 cc of a gas weighs 2.5 g at 750 torr and 300 K (27°C). Calculate the molecular weight of gas.
- Q.15** A sample of gas having volume of 10 l at 101.325 kPa pressure and at temperature of 298 K (25°C) is compressed to a high pressure so that its volume reduces by 4.5 l. If the pressure rises by 0.1 MPa, what will be the rise in temperature?

ANSWERS

1. (1520) 2. (2000) 3. (2) 4. (15.625)
 5. (1.5) 6. (48) 7. (2, 1)
 8. (5.98, 5.98, 6.25) 9. (1.2205)
 10. (0.6944, 0.871) 11. (3.8144) 12. (687.82)
 13. (82.33) 14. (65) 15. (27.6)

Explanation

1. (1520)

Basis : 2 atm pressure.

Conversion factor between atm and mm Hg is
 1 atm = 760 mm Hg.

$$\therefore 2 \text{ atm} = ?$$

$$\therefore \text{Pressure} = 2 \text{ atm} \times \left(\frac{760 \text{ mm Hg}}{1 \text{ atm}} \right) \\ = 1520 \text{ mm Hg}$$

2. (2000)

Basis : Volumetric flow rate of 2 m³/s.

Relationship between volume in m³ and volume in l is

$$1 \text{ m}^3 = 1000 \text{ l}$$

\therefore Volumetric flow rate

$$= 2 \text{ m}^3/\text{s} \times \frac{1000 \text{ l}}{1 \text{ m}^3} = 2000 \text{ l/s}$$

3. (2)

Basis : 88 kg of carbon dioxide.

Molecular formula of carbon dioxide = CO₂

Atomic weight: C = 12 and O = 16

Molecular weight of CO₂ = 1 × 12 + 2 × 16 = 44

$$\text{kmol of CO}_2 = \frac{\text{kg of CO}_2}{\text{Molecular weight of CO}_2}$$

$$= \frac{88}{44} = 2$$

\therefore 88 kg of CO₂ = 2 kmol CO₂

4. (15.625)

Basis : 500 g of oxygen.

Molecular weight of O₂ = 2 × 16 = 32

$$\text{Moles of O}_2 = \frac{500}{32} = 15.625 \text{ mol}$$

5. (1.5)

Basis : 117 kg of K.

Atomic weight of K = 39

$$\text{Atoms of K} = \frac{117}{39} = 3 \text{ katom}$$

Each mole of K₂CO₃ contains 2 atom of K

2 atom of K \equiv 1 mole of K₂CO₃

2 katom of K \equiv 1 kmol of K₂CO₃

\therefore Moles of K₂CO₃ = $\frac{1}{2} \times 3 = 1.5 \text{ kmol}$

6. (48)

Basis : 64 kg of methane.

Atomic weight of C = 12

Molecular weight of CH₄ = 16

1 katom of carbon \equiv 1 kmol of CH₄

\therefore 12 kg of carbon \equiv 16 kg of CH₄

i.e., in 16 kg of CH₄, 12 kg of carbon are present.

So, amount of carbon present in 64 kg of methane

$$= \frac{12}{16} \times 64 = 48 \text{ kg}$$

7. (2, 1)

Basis : One litre of solution.

Amount of H_2SO_4 dissolved = 98 g

Molecular weight of H_2SO_4 = 98

To find the normality we must know the gram equivalents of H_2SO_4 . So, find the gram equivalents using equivalent weight of H_2SO_4 .

$$\therefore \text{Equivalent weight of } H_2SO_4 = \frac{98}{2} = 49$$

$$\text{Gram-equivalents (g eq.) of } H_2SO_4 = \frac{98}{49} = 2$$

$$\therefore \text{Normality, } N = \frac{\text{g eq of } H_2SO_4}{\text{Volume of solution in l}} \\ = \frac{2}{1} = 2$$

$$\text{Moles of } H_2SO_4 = \frac{98}{98} = 1.0 \text{ mol}$$

$$\therefore \text{Molarity, } M = \frac{\text{Moles of } H_2SO_4}{\text{Volume of solution in l}} \\ = \frac{1}{1} = 1$$

8. (5.98, 5.98, 6.25)

Basis : 100 kg of solution.

The solution contain 20 kg NaOH and 80 kg water (solvent).

Density of solution = 1.196 kg/l

$$\text{Volume of solution} = \frac{100}{1.196} = 83.62 \text{ l}$$

$$\text{Moles of NaOH in solution} = \frac{20}{40} \\ = 0.5 \text{ kmol} = 500 \text{ mol}$$

$$\therefore \text{Molarity, } M = \frac{\text{gram moles of NaOH}}{\text{Volume of solution in litre}} \\ = \frac{500}{83.52} = 5.98$$

For NaOH as valence = 1,

Equivalent weight = Molecular weight

\therefore Normality, N = Molarity, M = 5.98

$$\text{Molality} = \frac{\text{gram moles of NaOH}}{\text{kg of solvent}} \\ = \frac{500}{80} = 6.25 \text{ mol/kg}$$

9. (1.2205)

Basis : 1 litre of solution.

Molarity = 11.24 and Molality = 94

$$\text{Now, Molarity} = \frac{\text{gram moles of } H_2SO_4}{\text{Volume of solution}}$$

$$\therefore 11.24 = \frac{\text{gram moles of } H_2SO_4}{1}$$

$$\therefore \text{Moles of } H_2SO_4 = 11.24 \times 98 \\ = 1101.52 \text{ g} = 1.101 \text{ kg}$$

Molality = mol H_2SO_4 / kg of solvent

\therefore Amount of solvent

$$= \frac{11.24}{94} = 0.1195 \text{ kg}$$

Amount of solution

$$= 1.101 + 0.1195 = 1.2205 \text{ kg}$$

Density of H_2SO_4 solution

$$= \frac{1.2205}{1} = 1.2205 \text{ kg/l}$$

10. (0.6944, 0.871)

Basis : 100 kg of methanol in the saturated solution.

Solution contains 44 kg of methyl bromide

Weight of the saturated solution

$$= 100 + 44 = 144 \text{ kg}$$

Weight fraction of methanol in the saturated solution

$$= \frac{100}{144} = 0.6944$$

Molecular weight of CH_3OH = 32

Molecular weight of CH_3Br = 94.91

\therefore Moles of CH_3OH in solution

$$= \frac{100}{32} = 3.125 \text{ kmol}$$

and Moles of CH_3Br in solution

$$= \frac{44}{94.91} = 0.4636 \text{ kmol}$$

\therefore Total moles of solution

$$= 3.125 + 0.4636$$

$$= 3.5886 \text{ kmol}$$

Mole fraction of methanol in saturated solution

$$= \frac{3.125}{3.5886} = 0.871$$

11. (3.8144)**Basis :** 2 m³ volume of SO₂ gas

$$PV = nRT$$

$$\therefore n = \frac{PV}{RT}$$

where, $P = 97.33 \text{ kPa}$
 $V = 2 \text{ m}^3$, $T = 393 \text{ K}$
 $R = 8.31451 \text{ m}^3 \cdot \text{kPa}/(\text{kmol} \cdot \text{K})$

 \therefore Moles of SO₂ gas

$$= \frac{97.33 \times 2}{8.31451 \times 393} = 0.0596 \text{ kmol}$$

$$\text{Weight of SO}_2 \text{ gas} = 0.0596 \times 64 = 3.8144 \text{ kg}$$

12. (687.82)**Basis :** 1 m³ volume of gas at 298 K.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

where, $P_1 = 131.7 \text{ kPa}$, $P_2 = 303.98 \text{ kPa}$,
 $T_1 = 298 \text{ K}$, $V_1 = V_2 = 1 \text{ m}^3$
(as vessel being closed), $T_2 = ?$

$$\therefore \frac{131.7 \times 1}{298} = \frac{303.98 \times 1}{T_2}$$

$$\therefore T_2 = 687.82 \text{ K}$$

$$\therefore \text{Temperature of gas attained} \\ = 687.82 \text{ K} (414.82^\circ\text{C})$$

13. (82.33)**Basis :** Average molecular weight of 31 of gas mixture.Let x_{N_2} and x_{CO_2} be the mole fractions of N₂ and CO₂ respectively.Molecular weight of N₂ = 28, Molecular weight of CO₂ = 44

$$M_{\text{avg}} = \sum M_i x_i$$

$$M_{\text{avg}} = M_{\text{N}_2} \cdot x_{\text{N}_2} + M_{\text{CO}_2} \cdot x_{\text{CO}_2}$$

$$\therefore 31 = 28x_{\text{N}_2} + 44x_{\text{CO}_2} \quad \dots(1)$$

$$\sum x_i = 1$$

$$x_{\text{N}_2} + x_{\text{CO}_2} = 1 \quad \dots(2)$$

$$\therefore x_{\text{CO}_2} = 1 - x_{\text{N}_2} \quad \dots(3)$$

Put the value of x_{N_2} from equation (2) in equation (1) and solve for x_{N_2} .

$$31 = 28x_{\text{N}_2} + 44(1 - x_{\text{N}_2})$$

$$16x_{\text{N}_2} = 13$$

$$\therefore x_{\text{N}_2} = 0.8125$$

$$\therefore x_{\text{CO}_2} = 1 - 0.8125 = 0.1875$$

Partial pressure of N₂ = $x_{\text{N}_2} \cdot P$

$$= 0.8125 \times 101.325$$

$$= 82.33 \text{ kPa} (617.52 \text{ torr})$$

14. (65)**Basis :** 2.5 g of gas at 300 K.Volume of gas = $V = 960 \text{ cc} = 0.96 \times 10^{-3} \text{ m}^3$

$$P = 750 \text{ torr} = 99.99 \text{ kPa}$$

$$PV = nRT$$

$$\therefore PV = \frac{m}{M} RT$$

where, $m = \text{Mass of gas}$
 $= 2.5 \text{ g} = 2.5 \times 10^{-3} \text{ kg}$,
 $M = \text{Molecular weight of gas}$
 $R = 8.31451 \text{ m}^3 \cdot \text{kPa}/(\text{kmol} \cdot \text{K})$

$$\therefore 99.99 \times 0.96 \times 10^{-3} = \frac{2.5 \times 10^{-3}}{M} \times 8.31451 \times 300$$

$$M = 64.69 \approx 65$$

15. (27.6)**Basis :** 10l of gas at 298 K

$$PV = nRT$$

where, $V = 10 \text{ l} = 10 \times 10^{-3} \text{ m}^3$
 $R = 8.31451 \text{ m}^3 \cdot \text{kPa}/(\text{kmol} \cdot \text{K})$
 $T = 298 \text{ K}$
 $P = 201.325 \text{ kPa}$

$$\therefore n = \frac{PV}{RT}$$

$$= \frac{101.325 \times 10 \times 10^{-3}}{8.31451 \times 298}$$

$$= 4.09 \times 10^{-4} \text{ kmol}$$

Final pressure = $101.325 + 100 = 201.325 \text{ kPa}$
(as pressure increases by 0.1 MPa i.e., 100 kPa)

$$\text{Final volume} = 10 - 4.5 = 5.5 \text{ l} = 5.5 \times 10^{-3} \text{ m}^3$$

$$PV = nRT$$

$$\text{where, } P = 201.325 \text{ kPa,}$$

$$V = 5.5 \times 10^{-3} \text{ m}^3$$

$$201.325 \times 5.5 \times 10^{-3} = 4.09 \times 10^{-4} \times 8.31451 \times T$$

$$\therefore T = 325.6 \text{ K}$$

$$\therefore \text{Rise in temperature} = 325.6 - 298 \\ = 27.6 \text{ K} (27.6^\circ\text{C})$$

