Chemical Engineering (GATE & PSUs)

Postal Correspondence

GATE & Public Sectors

Process Calculations

CONTENT

1. INTRODUCTION 03-14

2. MATERIAL BALANCE CALCULATION 15-29

3. ENERGY BALANCE CALCULATION 30-42

4. COMBUSTION CALCULATION 43-51

5. SATURATION AND HUMIDIFY 52-58

6. PRACTICE SET-I 59-68

7. PRACTICE SET-II 69-72

8. PRACTICE SET-III 73-76

9. PRACTICE SET-IV 77-89

10. GATE PREVIOUS PAPER 90-105

11. SUBJECTIVE QUESTION WITH SOLUTION 106-129

GATE 2015 Cut-off Marks				
BRANCH	**GENERAL**	**SC/ST/PD**	**OBC(Non-Creamy)**	**Total Appeared**
Chemical Engineering	27.52	18.34	24.77	15874

CHAPTER-1

INTRODUCTION

Units & Dimensions

(1.) Newton (N): The SI unit of force. It is equal to the force that would give a mass of one kilogram an acceleration of one m/s^2, and is equivalent to 100,000 dynes.

Metric unit of force, used also as a unit of weight (force due to gravity).

1 Newton = 1 kg • m/s^2

The definition of the standard metric unit of force is stated by the above equation. One Newton is defined as the amount of force required to give a 1-kg mass an acceleration of 1 m/s^2.

(2.) Kilogram-force (kg_f): 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 (1N).

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

(3.) Mole: Mole is a unit of measurement used in **chemistry** to express amounts of a **chemical** substance, defined as the amount of any substance that contains as many elementary entities as there are atoms in 12 grams of pure carbon-12, the isotope of carbon with relative atomic mass of exactly 12 by definition.

In other words a **mole** is simply a unit of measurement. Units are invented when existing units are inadequate. Chemical reactions often take place at levels where using grams wouldn't make sense, yet using absolute numbers of atoms/molecules/ions would be confusing, too.

A mole is the quantity of anything that has the same number of particles found in 12.000 grams of carbon-12. That number of particles is Avogadro's Number, which is roughly 6.02×10^{23}. A mole of carbon atoms is 6.02×10^{23} carbon atoms. A mole of chemistry teachers is 6.02×10^{23} chemistry teachers.

Examples: 1 mole of NH_3 has 6.022×10^{23} molecules and weighs about 17 grams. 1 mole of copper has 6.022×10^{23} atoms and weighs about 63.54 grams.

$$\textbf{Mole} = \frac{\text{mass in gram}}{\text{molecular weight}}$$

(4.) Avogadro's number: Avogadro's number is the number of particles found in one mole of a substance. It is the number of atoms in exactly 12 grams of carbon-12. This experimentally determined value is approximately 6.022 x 10^{23} particles per mole. Also known as Avogadro's constant

$$\boxed{1\text{mol}=6.023\times10^{23}\text{atoms}}$$

(5.) Pressure: The ratio of force to the area over which that force is distributed.

Pressure is force per unit area applied in a direction perpendicular to the surface of an object.

$$\boxed{P=\frac{\text{Force}}{\text{Area}}=\frac{F}{A}=\frac{F.d}{A.d}=\frac{W}{V}=\frac{\text{Energy}}{\text{Volume}}}$$

(6.) Gauge pressure: The amount by which the pressure measured in a fluid exceeds that of the atmosphere. Gauge pressure is the pressure relative to atmospheric pressure. Gauge pressure is positive for pressures above atmospheric pressure, and negative for pressures below it.

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

(7.) Absolute pressure is zero-referenced against a perfect vacuum, so it is equal to gauge pressure plus atmospheric pressure.

A *barometer* is a device that measures atmospheric pressure.

(8.) Vacuum pressure - Pressures below atmospheric pressure are called vacuum pressures and are measured by vacuum gages that indicate the difference between the atmospheric pressure and the absolute pressure.

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

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

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

(9.) Molarity (M): Molarity's definition is a unit of measurement used to denote the concentration of a particular substance or solution. Molarity is expressed as moles of solute over liters of solution. It is denoted either by a capital M or by the term molar.

$$\boxed{\text{Molarity (M)} = \frac{\text{No. of gram moles of solute}}{\text{Volume of solution in Litre}}}$$

(10.) Normality(N): Normality is a measure of concentration equal to the gram equivalent weight per liter of solution. Gram equivalent weight is the measure of the reactive capacity of a molecule.

$$\boxed{\text{Normality (N)} = \frac{\text{Gram equivalent of solute}}{\text{Volume of solution in Litre}}}$$

For acid reactions, a 1 M H_2SO_4 solution will have normality (N) of 2 N because 2 moles of H^+ ions are present per liter of solution. For sulfide precipitation reactions, where the SO_4^- ion is the important part, the same 1 M H_2SO_4 solution will have a normality of 1 N.

(11.) Molality: Molality is the number of moles of solute per kilogram of solvent. It is important the mass of solvent is used and not the mass of the solution. Solutions labeled with molal concentration are denoted with a lower case m.

$$\boxed{\text{Molality (m)} = \frac{\text{No. of gram moles of solute}}{\text{Weight of solvent in Kilogram}}}$$

The important part of remembering the difference is

Molarity - M $\rightarrow$ moles per liter solution

Molality - m $\rightarrow$ moles per kilogram solvent

(12.) Concentration : It is defined as the amount of solute in gram dissolved in one litre of solution. It is denoted by symbol 'C'

$$\boxed{\text{Concentration (C)} = \frac{\text{Mass of solute in gram}}{\text{Volume of solution in litre}} = \text{Normality} \times \text{Equivalent weight}}$$

(13.) Equivalent weight: The equivalent weight to an element or a compound is equal to the atomic weight or molecular weight divided by valence.

$$\boxed{\text{Equivalent weight} = \frac{\text{Atomic weight or Molecular weight}}{\text{Valency}}}$$

(14.) Valency: Valency of an element or compound depends on number of hydrogen ions accepted or the hydroxyl ions donated for each atomic weight or molecular weight.

(15.) 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 * m), from the function W = F * s where W is work, F is force, and *s* is the displacement. The joule is also the SI unit of energy.

$$\boxed{\text{Joule} = \text{N.m} = \frac{\text{kg. m}^2}{s^2}}$$

(16.) 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.

Power = Work / time or **P = W / t**

Ideal gas Laws & their Applications

(17.) **Boyle's law:** Boyle's Law states that the product of the pressure and volume for a gas is a constant for a fixed amount of gas at a fixed temperature. Written in mathematical terms, this law is,

$$P \propto \frac{1}{V} \qquad P = \frac{C}{V} \qquad \boxed{PV = C}$$

Where P = Absolute pressure, C = constant

V = Volume occupied by gas P V = constant

If a sample of gas initially at pressure P_i and volume V_i is subjected to a change that does not change the amount of gas or the temperature, the final pressure P_f and volume V_f are related to the initial values by the equation $P_i V_i = P_f V_f$

Key Points

- According to Boyle's Law, if pressure doubles, volume halves.
- Boyle's Law holds true only if the number of molecules n and the temperature are both constant.
- Boyle's Law is used to predict the result of introducing a change, in volume and pressure only, to the initial state of a fixed quantity of gas.

Important Terms

Boyle's law

Boyle's law (sometimes referred to as the Boyle–Mariotte law) states that the absolute pressure and volume of a given mass of confined gas are inversely proportional, if the temperature remains unchanged within a closed system.

Ideal gas

An ideal gas is a theoretical gas composed of a set of randomly-moving, non-interacting point particles.

Isothermal

In thermodynamics, a curve on a p-V diagram for an isothermal process.

Boyle's law : Practice Questions

1. Boyle's Law states that...
   (a) P1V1 = P2V2
   (b) P1V2 = P2V1
   (c) P2V1 = P1V2
   (d) all of the above
2. Boyle's Law also states that...
   (a) at a constant temperature, the pressure exerted by a gas varies inversely with the volume of that gas.
   (b) at a constant temperature, the pressure exerted by a gas varies directly with the volume of that gas.
   (c) at a constant temperature, the pressure exerted by a gas remains the same no matter what the volume of that gas is.
   (d) at a constant temperature, the volume of a gas remains the same no matter what the pressure exerted by that gas is.
3. An application of Boyle's Law is...
   (a) at a constant temperature, as the pressure on a gas increases, the volume decreases.
   (b) at a constant temperature, as the volume of a gas increases, the pressure decreases.
   (c) at a constant temperature, as the pressure on a gas increases, the volume increases.

(d) A and B.

4. As a volume of a gas increases, the number of molecules of that gas that collide with the walls of the container decreases.
   (a) True
   (b) False

Answer Keys : 1. a 2. a 3. d 4. a

(18.) CHARLES' Law:

Charles's law states that if a given quantity of gas is held at a constant pressure, its volume is directly proportional to the absolute temperature. The expression of Charles's law that you should memorize:

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

Try using Charles's law to solve the following problem.

Charle's Law : Practice Questions

Example: A sample of gas at 15ºC and 1 atm has a volume of 2.50 L. What volume will this gas occupy at 30ºC and 1 atm?

Solution: The pressure remains the same, while the volume and temperature change—this is the hallmark of a Charles's law question.

So, $\frac{V_1}{T_1} = \frac{V_2}{T_2}$, then 2.50 L/288K = V_2/303K, and V_2 = 2.63 L

This makes sense—the temperature is increasing slightly, so the volume should increase slightly. Be careful of questions like this—it's tempting to just use the Celsius temperature, but you **must first convert to Kelvin temperature** (by adding 273) to get the correct relationships.

Example: A 600 mL sample of nitrogen is heated from 27 °C to 77 °C at constant pressure. What is the final volume?

Solution: T (K) = 273 + °C

T_i = initial temperature = 27 °C

T_i (K) = 273 + 27=300 K

T_f = final temperature = 77 °C

T_f K = 273 + 77 = 350 K

The next step is to use Charles' law to find the final volume. Charles' law is expressed as:

$V_i/T_i = V_f/T_f$

V_f = (600 mL)(350 K)/(300 K) = 700 mL

Answer: The final volume after heating will be 700 mL.

(20.) At NTP Condition:

Pressure = 1.01325×10^5 N/m^2, T = 273 k

For an ideal gas V = 22.4 m^3/ kmol

(21.) **Density of gas:** It is given by $\rightarrow$ $\boxed{\rho = \frac{P}{MRT}}$

Where P = Pressure

M = Molecular weight

R = Universal gas constant

T = Temperature

(22.) **Specific gravity of gas**: The specific gravity of a gas is the ratio of densities of a gas relative to dry air.

(23.) **Partial pressure** $\rightarrow$ In a mixture of gases, each gas contributes to the total pressure of the mixture. This contribution is the partial pressure. The partial pressure is the pressure of the gas, if the gas were in the same volume and temperature by itself.

(24.) **Dalton's Law of Partial Pressures**

It states that the total pressure exerted by a mixture of gases is the sum of partial pressure of each individual gas present. Each gas is assumed to be an **ideal gas**.

$$P_{total} = \overline{P}_1 + \overline{P}_2 + \overline{P}_3 + \ldots$$

where $\overline{P}_1$ and $\overline{P}_2$ are the partial pressure of gas 1and gas 2 in the mixture. Since each gas behaves independently, the ideal gas law can be used to calculate the pressure of that gas if we know the number of moles of the gas, the total volume of the container, and the temperature of the gas.

$$P_1 = \frac{n_1 RT}{V}$$

Each gas exerts the same pressure they would exert if they were in the container alone. For example, if a mixture of gases at 298 K in a 1.00 L container consists of 1.00 g mol of H_2, 1.50 g mol of N_2, and 2.00 g mol of Br_2. The partial pressure of each gas can be written

$$P_{N_2} = \frac{n_{N_2} RT}{V} \qquad P_{H_2} = \frac{n_{H_2} RT}{V} \qquad P_{Br_2} = \frac{n_{Br_2} RT}{V}$$

The total pressure can be expressed as

$$P_T = P_{N_2} + P_{H_2} + P_{Br_2} = \frac{n_{N_2} RT}{V} + \frac{n_{H_2} RT}{V} + \frac{n_{Br_2} RT}{V}$$

The total pressure in this example is 13.8 atm.

Note: The following observations about gas mixtures are important to remember.

- Each gas occupies the entire volume of the container.
- The gases will mix homogeneously.
- The gases should not react (no chemical reaction should occur between the gases in the mixture).
- The type of gas has no bearing on the partial pressure of the gas.

Dalton's Law: Practice Questions

Example: How can two or more gases in a closed container each occupy the entire volume of its container?

Answer: Ideal gases are considered to be point particles, the volume of each gas molecule is considered so small, it is not important. Ideal gases do not attract or repel other gases, therefore the intermolecular forces among the gas molecules is not important. Because gases consist mostly of empty space, the distance between gas molecules is large enough so that each gas molecule does not "see" another gas molecule. This is another reason why there are no interactions with other particles.

(25.) **Ideal gas equation:** This equation is a combination of 3 simpler gas relationships:

Boyle's law: $V \propto \frac{1}{P}$ (at constant n and T)

Charles' law: $V \propto T$ (at constant n and P)

Avogadro's law: $V \propto n$ (at constant P and T)

by above laws, we get

$$V \propto \frac{nT}{P}$$

$$V = \text{constant} \times \frac{nT}{P} = R\frac{nT}{P} \quad ; \quad R \text{ is the gas constant}$$

$$\boxed{PV = nRT}$$

Where, P = absolute pressure V = Volume of n kmol gas

The conditions $0\ ^0C$ and 1 atm are called **Normal temperature and pressure (NTP).**

Experiments show that at NTP, 1 mole of an ideal gas occupies 22.414 L.

So 1 mole = 22.414L

1 kmol = $22.414 m^3$

Ideal Gas: Practice Questions

Problem:

What is the volume (in liters) occupied by 49.8g of HCl at a temperature of 25.0ºC and a pressure of 0.950 atm?

$$T = 25.0 + 273.15 = 298.15K$$

$PV = nRT$ $\qquad P = 0.950$ atm

$$V = \frac{nRT}{P} \qquad n = 49.8g \times \frac{1 \text{ mol HCl}}{36.45 \text{ g HCl}} = 1.37 \text{ mol}$$

$$V = \frac{1.37 \text{ mol} \times 0.0821 \frac{L.atm}{mol.K} \times 298.15K}{0.950 \text{ atm}}$$

$V = 35.3$ L $\qquad \boxed{\text{At NTP this same mass of HCl occupied a volume of } 30.62\ L}$

(26.) **Gas Constant (R)**

$$R=\frac{PV}{nT}=\frac{(1atm)(22.414L)}{(1mol)(273.15K)}$$

$$R=0.082057\frac{L.atm}{mol.K};$$

$$R=8.314\frac{J}{mol.K}\quad ;\ R=62358.9\frac{mm\,Hg.L}{kmol.K}$$

(27.) **Mole Fraction:** The partial pressure of a gas can also be expressed in terms of the **mole fraction** of each gas. It is denoted by x_i (for liquid) and y_i (for gas)

So, $\boxed{y_i=\frac{n_i}{n_{total}}==\frac{n_i}{n_A+n_B+n_C+....}}$

There is a relationship between the partial pressure of a gas, the total pressure of the system and the mole fraction of the gas.

$\overline{P}_{N_2}=P_{total}\cdot y_{N_2}$ ...(1) where, y_{N_2} = mole fraction of N_2

Equation (1) is known as Raoult's law.

Key Concepts

- Partial pressure of a gas in a mixture of gases is the pressure which that gas would exert if it were the only gas present in the container.
- Dalton's Law of Partial Pressures states that the total pressure in a gas mixture is the sum of the partial pressures of each individual gas.

$$P_{total}=\overline{P}_1+\overline{P}_2+\overline{P}_3+...$$

- Dalton's Law of Partial Pressures assumes each gas in the mixture is behaving like an ideal gas.

(28.) **Average molecular weight**: It is given as $\boxed{M_{av}=\Sigma M_i y_i}$

where M_i= molecular weight of component i

Y_i = mole fraction of component i in the mixture.

(29.) **S.I. (International System of Units):** The 'International System of Units' is the standard modern form of the metric system. The International System of Units is a system of measurement based on 7 base units: the metre (length), kilogram (mass), second (time), ampere (electric current), Kelvin (temperature), mole (quantity), and candela (brightness). These base units can be used in combination with each other. This creates SI derived units, which can be used to describe other quantities, such as volume, energy, pressure, velocity etc.

(30.) Composition:

(i.) Solid composition expressed in weight% or in mole% wherever nothing say about composition i.e. whether it is mole% or weight% it is taken as weight%

(ii) Volumetric composition of a liquid solution will change with temperature. The system composition expressed in mole percent will not vary with the temperature.

(iii.) For an ideal gases, the composition in mole percent exactly same as the composition in volumetric percent and exactly same as the composition of pressure percent. This statement only applies for gases and does not apply to liquid and solid system.

$$\boxed{\text{Volume\%} = \text{Pressure\%} = \text{Mole\%}}$$

(31.) ppm (parts per million parts):

(i.) Very often the impurities present in solid or liquid compound are expressed in ppm. This is defined on weight basis. If solution is watery. $\boxed{1\ \text{mg/lit} = 1\ \text{ppm}}$

(ii.) This unit is use in analysis of water treatment problem and effluent treatment problem.

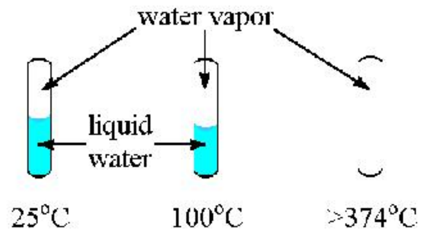
(32.) Critical temperature: Gases can be converted to liquids by compressing the gas at a temperature less than equal to critical temperature.

Gases become more difficult to liquefy as the temperature increases because the kinetic energies of the particles that make up the gas also increase.

The **critical temperature** of a substance is the temperature at and above which vapor of the substance cannot be liquefied, no matter how much pressure is applied.

Every substance has a critical temperature. Some examples are shown below.

Substance	Critical Temperature (°C)
NH_3	132
O_2	-119
CO_2	31.2
H_2O	374



Tubes containing water at several temperatures. Note that at or above 374°C (the critical temperature for water), only water vapor exists in the tube.

(33.) **Critical pressure:** The critical pressure of a substance is the pressure required to liquefy a gas at its critical temperature. Some examples are shown below.

Substance	Critical Pressure (atm)
NH_3	111.5
O_2	49.7
CO_2	73.0
H_2O	217.7

(34.) **Critical volume**: The volume occupied by a gas under critical condition (at critical T & P)

(35.) **Specific gravity of solid or liquid (SG):** it is defined as

$$\boxed{(SG)_{T_1/T_2} = \frac{\text{Density of solid or liquid at } T_1 \text{ K}}{\text{Density of water } T_2 \text{ K}}}$$

(36.) **Scales use for specific gravity:**

(i.) Twaddel scale: Used for liquid heavier than (^{o}Twaddell or ^{o} Tw) water

(ii.) Baume gravity scale (^{o}Be)

(iii.) API scale: For petroleum products ^{o} (API)

(iv.) Brix scale: (^{o} Brix) measure directly sugar solution concentration

(37.) **p_H:** It is used to defined acidity or alkalinity of a solution. It is defined as

$$\boxed{p_H = -\log_{10}(H^+)}$$

$\Rightarrow$ if $p_H < 7 \rightarrow$ Acid , $p_H = 7 \rightarrow$ Neutral

$p_H > 7 \rightarrow$ Base

Where H^+ is the hydrogen ion concentration in g_{eq}/L

(38.) **Duhring's plot** is a graphical representation of such a relationship, typically with the pure liquid's boiling point along the x-axis and the mixture's boiling point along the y-axis; each line of the graph represents a constant concentration. It is used in the problems relating evaporation.

(39.) **Duhring's rule :**The rule that a plot of the temperature at which a liquid exerts a particular vapor pressure against the temperature at which a similar reference liquid exerts the same vapor pressure produces a straight or nearly straight line.

(40.) **Amagat's law:** The volume of an ideal gas mixture (V) is equal to the sum of the component volumes (V_j's) of each individual component in the gas mixture at the same temperature (T) and total pressure (P) of the mixture. For example,

Gas A T and P V_A	+	Gas B T and P V_B	=	Gas Mixture T and P $V_A + V_B = V$
$PV_A = n_A RT$		$PV_B = n_B RT$		$PV = nRT$

$$\frac{V_A}{n_A} = \frac{V_B}{n_B} = \frac{V}{n} \quad \text{thus} \quad \frac{n_A}{n} = \frac{V_A}{V} = y_A \quad \text{and} \quad \frac{n_B}{n} = \frac{V_B}{V} = y_B$$

(41.) **Compressibility factor:** Compressibility factor (Z) is defined as

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

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

KEY POINTS TO REMEMBER

(Introduction)

(1) **Mole** = $\frac{\text{mass in gram}}{\text{molecular weight}}$

(2) **1 mole** = 6.023×10^{23} atoms

(3) **Absolute pressure** = gauge pressure + atmospheric pressure

(4) **Absolute pressure** = atmospheric pressure – vacuum

(5) **Ideal gas equation** $\Rightarrow$ $\boxed{PV=nRT}$

Where P = Absolute pressure

(6) **Boyle's law:** $\boxed{PV=\text{constant}}$

V = Volume occupied by gas

(7.) **CHARLES's Law: -** $\boxed{\frac{V}{T}=\text{constant}}$ where T = Absolute temperature

(8.) **Gas density** = $\dfrac{\text{molecular weight of gas}}{\text{molal volume of gas}}$

(9.) **Specific gravity of gas** = $\dfrac{\text{molecular weight of gas}}{\text{molal weight of air}}$

(10.) **pH**= $-\log_{10}(H^+)$

(11.) **Compressibility factor:** $\boxed{Z=\dfrac{PV}{nRT}}$

NUMERICALS

Question-1: A compound whose molecular weight is 106, & composition on weight basis

C – 85% H_2 – 5% N_2 – 10%

Find formula of this compound.

Solution: Basis: compound is 106 kg (1 kg mol)

$$\text{Carbon amount} = \frac{85\times106}{100} = 90.1\text{kg}$$

$$\text{kg atoms of carbon} = \frac{90.1}{12} = 7.51 \simeq 8$$

$$\text{Amount of hydrogen} = \frac{5\times106}{100} = 5.3$$

$$\text{kg atoms of hydrogen} = \frac{5.3}{1} = 5.3 \simeq 5$$

$$\text{Nitrogen amount} = \frac{10\times106}{100} = 10.6$$

kg atoms of nitrogen = $\sum$

Hence , the compound formula is given by: $\boxed{C_8H_5\,N}$

Sample Study Materials