**Q 1 Define:**

**Mole:** …………..The quantity of a given substance that contains as many molecules or formula units the number of atoms in exactly 12g of carbon -12.

**Stoichiometry:-……….**The quantitative relationships between reactants and products in a balanced equation

**Pressure: ……….**is a measure of the collisions of the atoms with the container. (Force per unit area, P = F/A)

**Boyle’s Law -** ………The volume of a fixed quantity of gas is inversely proportional to its pressure.

**Charles’ Law:** ………..The volume of a fixed amount of gas at constant pressure is directly proportional to the Kelvin (absolute) temperature. (Vα T, V1T2 = V2T1)

**Gay-Lussac’s Law:** ………. “The pressure of a sample of gas is directly proportional to the absolute temperature when volume remains constant.” (P1T2 = P2T1)

**Combined Gas Law:** …………Pressure and volume are inversely proportional to each other and directly proportional to temperature. (P1V1T2 = P2V2T1).

**Avogadro’s Law :** “At a fixed temperature and pressure, the volume of a gas is directly proportional to the amount of gas., V = c · n”

**Dalton’s Law of Partial Pressure**: …….“The total pressure of a mixture of gases is the sum of the partial pressures of the components of the mixture. Ptotal = PA + PB + ……”

**Graham’s Law:** ……..The rate of effusion of a gas is inversely proportional to the square root of its molecular mass.

**describe what a gas is?**

1. Gases consist of a large number of molecules in constant random motion.
2. Volume of individual molecules negligible compared to volume of container.
3. Intermolecular forces (forces between gas molecules) negligible.
4. Energy can be transferred between molecules, but total kinetic energy is constant at constant temperature.
5. Average kinetic energy of molecules is proportional to temperature

**Endothermic reaction:** **the reaction which is supplied by heat (absorb heat and Δ H is + Ve).**

**Exothermic reaction: the reaction which is accompanied with evolution of heat (evolve heat and Δ H is - Ve).**

**Internal energy (E): the total amount of energy contained within a system (kinetic and potential energies).**

**Heat content, Enthalpy (ΔH): it is a thermodynamic function describes the total energy of the system under constant pressure and it is given by H = E + PV**

**Specific heat; J g-1 oC-1: the amount of heat which required to raise the temperature of one gram of substance by one degree centigrade.**

**Heat capacity; J°C-1 Specific Heat x mass. “The amount of heat which required raising the temperature of the substance by 1°C.”**

**Molar heat capacity: “the amount of heat which required to raise the temperature of one mole of substance by 1°C molar heat capacity = specific heat x M.wt.”**

**First law of thermodynamics** Energy can neither created not destroyed, only transformed from one form to another

**Heat of combustion** **The change in enthalpy (heat evolved) when 1 mole of a substance is completely burnt in presence of excess oxygen.**

# ****Heat of Formation: ΔH°**f is the enthalpy change for the formation of one mole of the substance from its elements, at standard pressure (1 atm) and a specified temperature**

**Hess's law of constant heat summation** If a reaction is carried out in a series of steps, ΔH for the reaction is the sum of ΔH’s for each of the steps.

**Weight percent (wt%)** **Number of grams of solute which present in 100 gram of solution.**

**Mole fraction (X):The ratio of the number of moles of one component to the total number of moles of all components in the mixture solution**

**Molality** ,

**Molarity (M): The no. of moles of solute dissolved in 1 liter of solution**

**Henry’s Law –** The solubility of a gas increases in direct proportion to its partial pressure above the solution.

**Raoult’s Law –** The equilibrium vapor pressure of the solvent over the solution is directly proportional to the mole fraction of the solvent in the solution

**Le Chatelier's Principle: states that a system in equilibrium reacts to a stress in a way that counteracts the stress and establishes a new equilibrium state**

**Neutralization reaction: The reaction between an acid and a base, in which water is produced, is called neutralization.**

**H+ (aq) + OH- (aq) → H2O neutralization
HCI (aq) → H+ (aq) + Cl- (aq) acid
NaOH (aq) → Na+ (aq) + OH- (aq) base
H+(aq)+Cl- (aq)+Na+(aq) +OH- (aq) → Na+(aq)+ Cl- (aq) + H2O**

**Heterogeneous reactions the mixture is not uniform throughout.**

**e.g: C2H5OH(l) C2H4(g) + H2O
*ethanol ethane***

**The Bronsted-Lowry Concept**:

**This concept defines acids and bases (which may be molecules or ions) in terms of the exchange of a proton. In an acid - base reaction, an acid donates a proton to a base, which accepts it. In losing a proton, acid1 becomes base1 (The conjugate base of acid1), and gaining a proton, the original base2 becomes acid2 (the conjugate acid of base2).**

**Ideal Solution:**

1. **The force of attraction between all molecules are identical i.e. the attraction force is not affected by addition of other components A - A = B-B = A - B.**
2. **No heat is evolved or absorbed during mixing i.e. ∆H soln. = Zero**
3. **The volume of solution is the sum of volumes of the two liquids.**
4. **The solution obeys Raoult's law.**
* **Non- ideal solutions (Solutions deviate from ideal behavior).**

|  |  |
| --- | --- |
| **Negative deviation** | **Positive deviation** |
| **1- The force of attraction****increase by mixing** **A - A, B-B < A-B** | **The force of attraction decrease by mixing A-A , B-B > A-B** |
| **2- The vapor pressure will be lower than that given by Roault's law** | **The vapor pressure will be higher than that given by Raoult's law.** |
| **3- ΔH :- Ve (exothermic)** | **Δ H : + Ve (endothermic)** |
| **4- Temperature change when solution is formed: increase** | **Temperature change when solution is formed: decrease.** |
| **5- Example: Acetone-water****Vapour pressure of non-ideal solution (-ve deviation)** | **Ethanol-hexane****Vapour pressure of non-ideal solution (+ve deviation)** |

**Q 2 Choice ONLY one correct answer:**

|  |
| --- |
| 1. Which of the following is fundamental (or base) unit:
 |
| 1. Mol
 | 1. m/S
 | 1. g
 | 1. All of them
 |

1. "Pressure" of gas

a) Is a measure of the collisions of the atoms with the container?

b) Pascal (Pa), N/m2 is an unit of its measurement

c) Can be measured using manometer.

d) all of the above

|  |
| --- |
| 1. Ideal gas constant, R:
 |
| 1. 0.0821 liter torr K -1mol-1
 | c) 0.0821 ml atm K -1mol-1 |
| 1. 0.0821 liter atm K -1mmol-1
 | d) 0.0821 liter atm K -1mol-1 |

1. Giga, G is equal to

|  |  |  |  |
| --- | --- | --- | --- |
| 1. 109
 | 1. 106
 | 1. 10-9
 | 1. 10-6
 |

1. 149oF is equal to

|  |  |  |  |
| --- | --- | --- | --- |
| 1. 100 oC
 | 1. 65 oC
 | 1. 0oC
 | 1. – 13 oC
 |

1. Which is the limiting **reactant for the following equation**

**3 Fe + 4H2O 🡪 Fe3 O4 + 4H2**

|  |  |  |  |
| --- | --- | --- | --- |
| 1. **Fe**
 | 1. **H2O**
 | 1. **Fe3 O4**
 | 1. **H2**
 |

|  |
| --- |
| 1. Ideal gas constant, R:
 |
| 1. 0.0821 liter torr K -1mol-1
 | c) 0.0821 ml atm K -1mol-1 |
| 1. 0.0821 liter atm K -1mmol-1
 | d) 0.0821 liter atm K -1mol-1 |
| 1. What is the density of gas which its diffusion is 1.414 times of the rate of diffusion of CO2 at STP?

|  |  |
| --- | --- |
| a) ≈ 2.5 | c) ≈ 1 |
| b) ≈ 1.77 | d) none of the above  |

 |

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| 1. Which of the following compounds is more miscible in water

|  |  |
| --- | --- |
| a)CH3OH | c)CH3CH(OH)2 |
| b) CH3CH2OH | d) CHCl3 |

 |
| 1. Its solubility in water decreases by elevation of temperature :
 |
| a) CO­2 | b) glucose | c) NaCl | d) KMnO4 |

|  |
| --- |
| 1. Carbonated beverages is an example of:
 |
| a) liquid / liquid solution mixture | c) solid /liquid solution mixture  |
| b) gas /liquid solution mixture  | d) gas/gas solution mixture  |

1. An example of **weight to volume expression for concentration:**
2. **Molality b) mole fraction c)Molarity d) mass percent**
3. **Number of grams of solute which present in 100 gram of solution is**
	1. **Molality b) mole fraction c)Molarity d) mass percent**
4. Not all alcohols are miscible with water, the more C atoms in the alcohol,
	1. More soluble b) less soluble c) has no effect on solubility
5. **Has the same osmotic pressure as blood.**
	1. **NaCI (9%) b) NaCI (0.9%) c) NaCI (0.09%) d) NaCI (0.009%)**
6. **Has the same osmotic pressure as blood.**

**a) Hypertonic solution b) Isotonic solution c) Hypertonic solution**

1. **The amount of heat which required raising the temperature of one gram of substance by one degree centigrade.**
	1. **Enthalpy b) Specific heat c) Heat capacity d) Internal energy**

|  |
| --- |
| 1. The unit of heat capacity is:
 |
| a) JoC-1  | c) Jog-1  |
| b) Jog-1 C-1  | d) Kj |
| 1. The apparatus used for determination of heat of combustion is :
 |
| a) coffee –cup calorimeter  | b) Bomb calorimeter  |
| c) both a and b  | d) none of the above  |
| 1. Which of the following COMPLETE thermochemical equation :
 |

a) C2H4 + 3 O2 → 2CO2 (g) + 2H2O

b) H2 (g) + ½ O2 (g) → H2O (l) ΔH = -68.32 Kcal

c) ½H2 + ½Cl2 → HCl ΔH= -44.0 Kcal

d) CO (g) + O2 (g) → CO2 (g) ΔH= -284.5 Kj

1. If Kc is larger than Kp ,the forward reaction is:

a-Complete b-Incomplete

c-None of the above d-Reversible

1. If Q < Kc the reaction will …………

a-Move from left to right b-Move from right to left

c-Be at equilibrium d-Be reversible

1. The reaction between an acid and a base in which water is produced is called:

a-Acidification b-Neutralization

c-Alkalanization d- None of the above

1. In the Lewis concept , the formation of ………….is the basis for defining acid –base reaction.

a- A covalent bond b-Ionic bond

c-Double bond d-Triple bond

1. In case of 2 steps reactions, Kc equal to:

a-K1K2 b-K1=K2

c-K1/K2 d-K1

1. In case of reaction involving gases ……………

a-Kc=Kp b-Kc<Kp

c-Kp<Kc d-Kc ≠Kp

1. All …………tend to attain a state of equilibrium.

a- Reversible b-Irreversible

 c- Heterogeneous d- Homogeneous

**Q 3 Calculate:**

* 1. **Convert the quantity from 2.3 x 10-8 cm to nanometers (nm) First we will need to determine the conversion factors**
	2. **Convert the quantity from 31,820 mi2 to square meters (m2) (1 mile = 1.6093km)**

* 1. **Convert the quantity from 14 m/s to miles per hour (mi/hr).**
	2. **How many grams of Cu are there in 2.55 mole of Cu? (1 mole Cu = 63.5 g Cu)**
	3. **What number of moles of aluminum is present in 125 g of Al? (1 mole Al = 27.0g Al)**
1. **How many carbon atoms are there in a 1.0 carat diamond? Diamond is pure carbon and one carat is exactly 0.2g.**
2. **What is the percentage of Fe in Fe2 O3?**

**2 mole Fe = 2 X 55.8g Fe = 111.6g Fe**

**3 mole O = 3 X16.0 g O = 48 g O
 = 159.6 g Fe2O3
% of Fe in Fe2O3 is**

1. **A 1.26 g sample of pure caffeine contains 0.624g C, 0.065g H, 0.364g N and 0.208g 0. What is the empirical formula of Caffeine, and if the molecular weight of caffeine is 194. What is the molecular formula?**

**Division of each of these values by the smallest value (0.013) gives the ratio. 4 mole C : 5 mole H : 2 mole N : 1 mol O**

**The empirical formula of caffeine is C4H5N2O.**

**The formula weight indicated by C4H5N2O is 9**

**The molecular formula of caffeine is C8H10N4O2**

1. **A compound was analyzed and found to contain 13.5 g Ca, 10.8 g O, and 0.675 g H.  What is the empirical formula of the compound?**

* Divide each mole value by the smallest number of moles calculated.  Round to the nearest whole number.

* This is the mole ratio of the elements and is represented by subscripts in the empirical formula. **CaO2H2 = Ca(OH)2**
1. **A 14.0 g sample of a compound that contains only carbon and hydrogen is completely burned.  The resulting products are 20.49 g of CO2 and 12.59 g of H2O.   If the molar mass of the compound is known to be 30 g/mol, then what is the empirical and molecular formula of the compound?**
* Number of moles of CO2 = number of moles of C in the original compound. = (12 ) + 2(16.00) = 44 g/mol
(20.49 g CO2/44 g) = 0.4656 mol CO2 therefore 0.4656 mol C in the original compound.

- Calculate the number of moles of H2O that is formed. Since there are 2 H atoms per H2O molecule, this is 1/2 the amount of H in the original compound.
H2O molar mass = 2H + O = 2(1) + 16.00 = 18 g/mol
(12.59 g H2O /18 g) = 0.6987 mol H2O therefore twice as much H (1.397 mol) in the original compound.
- Substitute the moles into a trial formula C 0.4656, H 1.397
- Divide by the smallest trial subscript, so that you can simplfy to a whole number ratio
- Therefore, **CH3** is the empirical formula.
- Calculate a trial molar mass for the empirical formula
CH3 molar mass = 1C + 3H = (12) + 3(1) = 15g/mol.
- Divide the actual molar mass of the compound by the trial molar mass, to find how many formula units there are in the real compound
(30 g/mol)/(15g/mol) = 2
So there are 2(CH3) units in the actual molecular formula, **C2H6**
1. **How many grams of N2F4 can theoretically be prepared from 4.0g of NH3 and 14.0g of F2? 2NH3 + 5F2 🡪 N2F4 + 6HF and if 4.8g of N2F4 is obtained from the experiment, what is the percent yields?**

 **The stoichiometric relationship is 2 mole NH3 = 5 mole F2**

**The F2, therefore is the limiting reactant, since a smaller proportionate amount has been supplied**

**? g N2F4 = 0.368 mol F2**

**from the equation 1 mole N2F4 = 5mole F2**

**104g N2F4 = 5 mole F2**

1. **An ideal gas is enclosed in a Boyle's-law apparatus. Its volume is 247 ml at a pressure of 625 mmHg. If the pressure is increased to 825 mmHg, what will be the new volume occupied by the gas if the temperature is held constant?**



1. **A 4.50-L sample of gas is warmed at constant pressure from 300 K to 350 K. What will its final volume be?**

 (4.50 L)(350. K) = V2 (300. K) V2 = 5.25 L

1. **A sample of gas is pumped from a 12.0 L vessel at 27ºC and 760 Torr pressure to a 3.5-L vessel at 52ºC. What is the final pressure?**

Equation: or P1V1T2 = P2V2T1

(760 Torr)(12.0 L)(325 K) = ( P2)(3.5 L)(300 K)

P2 = 2.8 x 10³ Torr

1. **Suppose 5.00 g of oxygen gas, O2, at 35 °C is enclosed in a container having a capacity of 6.00 liters. Assuming ideal-gas behavior, calculate the pressure of the oxygen in millimeters of mercury. (Atomic weight: 0 = 16.0)**

One mole of O2 weighs 2(16.0) = 32.0 g. 5.00 g of O2 is, therefore, 5.00 g/32.0 g mol-1, or 0.156 mol. 35 °C is 35 + 273 = 308 K

PV = nRT



1. **A gaseous mixture made from 6.00 g O2 and 9.00 g CH4 is placed in a 15.0 L vessel at 0ºC. What is the partial pressure of each gas, and what is the total pressure in the vessel?**

nO2 = = 0.188 mol O2

nCH4 = = 0.563 mol CH4

PO2 === 0.281

PCH4==0.841 atm

Ptotal = PO2 + PCH4 = 0.281 atm + 0.841 atm

 Ptotal  = 1.122 atm

1. **A mixture of 40.0g of oxygen and 40.0g of helium has a total pressure of 0.9 atm. What is the partial pressure of oxygen?**

**No of mole of O2 = 40/32= 1.25 mol**

**No of mole of He = 40/4= 10 mol**

**XO2 = no2/ no2+nHe = 1.25 / (1.2+10) = 1.25/11.2 = 0.112**

**Po2 = Xo2 Ptotal =** 0.112x0.9 = 0.101 atm

PHe = Ptotal – Po2 = 0.900 – 0.101 = 0.799 atm

1. **A sample of air has the following partial pressures of components; 593.4 mm Hg N2, 159.2 mm Hg O2, 7.1 mm Hg Ar and 0.3 mm Hg CO2. What is the composition of the air in mole fraction?**

**Mole fraction of N2 = 593.4/760 = 0.7808**

**Mole fraction of O2 = 159.2/760= 0.2095**

**Mole fraction of Ar = 7.1 / 760 = 0.0093**

**Mole fraction of CO2 = 0.3/ 760 = 0.0004**

1. **Suppose 1 g of H2, N2, and O2 are placed together in 10 liter container at 125oC. Assume ideal gas , calculate the total pressure (At.wt, H=1 , O=16, N=14)**

No of mole of H2 = 1/(1+1) = 0.5

No of mole of N2 = 1/ (14+14) = 0.0357

No of mole of O2 = 1/(16+16) = 0.0313

ntotal = 0.5 + 0.0357 + 0.0313 = 0.567

ptotal = 0.567 x 0.0821x (125+273)/10 = 1.853

1. **The rate of effusion of an unknown gas (X) through a pinhole is found to be only 0.279 times the rate of effusion of hydrogen (H2) gas through the same pinhole, if both gases are at STP. What is the molecular weight of the unknown gas? (Atomic weight, H = 1.01.)**

Mx= 26.0

1. **What is the relative rate of diffusion of H2 and CO2 under the same condition?**

 **.**

1. **What is the density of gas which it’s diffusion is 1.414 times of the rate of diffusion of CO2 at STP?.**
2. **Calculate the pressure exerted by 10.0 g of methane, CH4, when enclosed in a 1.00-liter container at 25 °C by using (a) the ideal-gas law and (b) the van der Waals equation. (a=**  **and b=** **)**

The molecular weight of CH4 is 16.0; so n, the number of moles of methane, is 10.0 g/16.0 g mol-1, or 0.625 mol.

(a) Considering the gas to be ideal and solving for P, we obtain



(b) Treating the gas as a Van der Waals gas and solving for P, we have



1. **What is the molality of 12.5 % solution of glucose C6H12O6, in water? M.wt. of glucose is 180.0**

**In 12.5 % solution 12.5 gm C6H12O6 is dissolved in l00 gm solution.**

**W solvent = 100 - 12.5 = 87.5 g H2O**

**no. of moles glucose = 12.5/18**

1. **Calcium hydride, CaH2, reacts with water to form hydrogen gas: CaH2(s) + 2 H2O(l)** 🡪 **Ca(OH)2(aq) + 2 H2(g)**

**How many grams of CaH2 are needed to generate 10.0L of H2 gas if the pressure of H2 is 740 torr at 23oC?**



1. **What are the mole fractions of solute and solvent in a 1.0m aqueous solution?**

no of moles of H2O H2O

A 1.0 m aqueous solution contains

n solute =1.0 mol



The mole fractions are

X solute 

X water = 

1. **a) How many grams of concentrated nitric acid solution should be used to prepare 250 ml of 2.0M HNO3? The concentrated acid is 70.0 %**

 **b) If the density of the concentrated nitric acid solution is 1.42 g/ml. What volume should be used? M.wt. (HNO3) =63**

**70 gm HNO3 → l00 gm solution**

 

**Mass of pure HNO3** 

**mass of HNO3 solution =** 

**b) ml cone. NHO3 = (45/1.42) = 31.7 ml cone. HNO3**

1. **An aqueous solution of acetic acid was prepared by dissolving 164.2 gm of acetic acid in 800 ml of the solution. If the density of the solution was 1.026 gm/ml. M. wt of acetic acid = 60**

**Calculate:a) The molar concentration of the solution,**

**b) The molality**

**c) The mole fraction of both the solute and the solvent**

**d) The mole %, e) The weight %.**

a) 

**b) d = 1.026g/ml V = 800 ml**

**W solution = V x d = 800 x 1.026 = 820.8 gm**

**W slvent = 820.8 - 164.2 = 656.6 gm**



**c) no. of acetic acid moles = 164.2 / 60 = 2.737 mole**

 **no. of H2O moles = 656.6 / 18 = 36.44 mole**

**Mole fraction of acetic acid =**

**Mole fraction of H2O =** 

**d) mole % acetic acid = 0.0699 x 100 = 6.99 %**

 **mole % of H2O = 0.9299 x 100 = 92.99 %**

**e) percentage weight of acetic acid =** 

**percentage weight of H2O=** 

1. **Five grams of NaCl is dissolved in 25.0 g of H2O. What is the mole fraction, mole percent, molality and mass percent of NaCl in the solution?**

No of mole of NaCl = 5/58.44 = 8.56 x 10-2

No of mole of water = 25/18 = 1.39

XNaCl = 8.56 x 10-2 /(8.56 x 10-2 +1.39)= 0.0580

Mole % = 0.058 x 100 = 5.8 %

Molality = 8.56 x 10-2 / 25/1000 = 3.42 m

Mass % = 5/ (5 + 25) x100 = 16.7%

1. **Ten grams of ascorbic acid (vitamin C), C6H8O6, is dissolved in enough water to make 125 ml of solution. What is the molarity of the ascorbic acid? (Answer = 5.80 mol %)**

10/176= 0.0568 mol

M= 0.0568 mol/125/1000 = 0.454 m

1. **At 25°C oxygen gas collected over water at a total pressure of 101 kPa is soluble to the extent of 0.0393 g dm-3. What would its solubility be if its partial pressure over water were 107 kPa? The vapor pressure of water is 3.0 kPa at 25°C.**

**P total = PH2O + PO2**

**PO2 = P total - PH2O = 101-3 = 98 kPa**

1. **Heptane (C7H16) and octane (C8H18) form ideal solutions What is the vapor pressure at 40°C of a solution that contains 3.0 mol of heptane and 5 mol of octane? At 40°C, the vapor pressure of heptane is 0.121 atm and the vapor pressure of octane is 0.041 atm.**

**X heptane = 3.0/8.0 = 0.375 X octane = 5.0/8.0 = 0.625**

**∴Total = X heptane . Po heptane + X octane. Po octane**

**= 0.375 x 0.12 +0.625 x 0.04 = 0.045 atm + 0.026 atm. = 0.071 atm.**

1. **Assuming ideality, calculate the vapor pressure of 1.0 m solution of a non - volatile, on dissociating solute in water at 50°C. The vapor pressure of water 50°C is 0.122 atm.**

See problem no 26

**the mole fraction of water in 1.0m solution is 0.982.**

**PH2O = XH2O PH2O = 0.982 x 0.122 = 0.120 atm.**

1. **A solution is prepared by mixing 5.81 g acetone C3H6O, (M. wt = 58.1 g/mole) 11.9 g chloroform (CHCI3 M.wt 119.4 g/mole). At 35°C this solution has a total vapor pressure of 260 torr. Is this an ideal solution? Comment? The vapor pressure of pure acetone and pure CHCI3 at 35°C are 345 and 293 torr, respectively.**

**n acetone =** 

**n CHCl3=** 

nt = 0.1 + 0.1 = 0.2 mole

**Xacetone  =** 

**XCHCl3 = 0.5**

**Pt = + = 345 x 0.5 + 293 x 0.5 = 319 torr. .**

**The observed value = 260 torr**

**• By comparing the 2 values, shows that the solution is not ideal.**

**• The observed value = 260 less than the expected value = 319 this is a negative deviation from Roault's law.**

1. **What is the normal boiling point of a 2.50 *m* glucose, C6H12O6, solution?**

1. **Automotive antifreeze consists of ethylene glycol, C2H6O2, a nonvolatile nonelectrolyte. Calculate the boiling point of a 25.0 mass percent solution of ethylene glycol in water.**

Boiling point = (normal b.p of solvent + ∆ T

(250x 1000)/(750x 62.1) = 5,38 m

Δ Tb = Kbm = 0.52 x 5.38 = 2.8oC

Bp = 100 + 2.8 = 102.8

1. **Calculate the freezing point of a 2.50 m aqueous glucose solution.**

Δ Tf = Kfm = 1.86 x 2.5 = 4.65 Fp = 0.00 + 4.65 = - 4.65OC

1. **A 37.0 g sample of a new covalent compound, a nonelectrolyte, was dissolved in 2.00 x 102 g of water. The resulting solution froze at -5.58oC. What is the molecular weight of the compound?**

1. **A 1.00 g sample of a biological material was dissolved in enough water to give 1.00 x 102 mL of solution. The osmotic pressure of the solution was 2.80 torr at 25oC. Calculate the molarity and approximate molecular weight of the material.**

1. **What would be the freezing point and boiling point of a solution containing 6.50 g of ethylene glycol (C2H6O2) in 200.0 g of H2O? KfH2O = 1.86°C /m, kbH2O = 0.512°C /m Solution m = W/M.wt x 1000/W(solvent) = 6.5/62.1 x 1000/200 = 0.525**

**ΔTf =tsolvent – tsolution  = m Kf = 1.86 x 0.525 = 0.977oC**

 **tsolution = zero - ΔTf = - 0.977 oC**

 **ΔTb =tsolvent – tsolution = 0.512x 0.525 = 0.267oC**

 **0.267oC = Tsolution  - 100 Tsolution = 100.267oC**

1. **What are the boiling point and freezing point of a solution prepared 2.40 g of biphenyl (C12H10) in 75.0 g of benzene? b.p of benzene = 80.1°C f.p. of benzene = 5.5°C. Solution : m= 2.4/154 x 1000/75 = 2.208 m**

**ΔTb = 2.208 x 2.53= 0.526OC Bp of solution= 80.1+0.526 = 80.626**

**ΔTf = 2.208 x 5.12 = 1.06 OC Fp of solution = 5.5 – 1.1 = 4.4 oC**

1. **A solution prepared by dissolving 0.30 g of an unknown nonvolatile solute in 30.0 g of CCI4 has a boiling point that is 0.392°C higher than that of pure CCI4. What is the molecular weight of the solute? Kb = 5.02°C/m.**

**ΔTb = m Kb +0.392oC=m x 5.02 m= 0.0781**

**0.0781= 0.3/Mwt x 1000/30 Mwt = 128g/mole**

1. **Find the osmotic pressure of blood at normal body temperature (37°C) if blood behaves as if it were a 0.296 M solution of a nonionizing solute.**

**Π= MRT = 0.246x 0.0821 x(273 +27) = 7.53 atm**

1. **An aqueous solution contains 30.0 g of a protein in l.0 L. The osmotic pressure of the solution is 0.0167 atm at 25°C. What is the approximate molecular weight of the protein?**

**Π= n/v RT 0.0167 = n/1 x 0.0821 x 248 n=6 Π=6.83 x 10 -4 mol**

**Mwt = mass/n = 30 /6.83 x 10 -4 mol = 4.39 x 104 g/mole**

1. **For the reaction N2O4(g)** ⮀ **2 NO2(g) The concentrations of the substances present in an equilibrium mixture at 25°C are [N2O4] = 4.27 x 10-2 mol/L , [NO2] = 1.41 x 10-2 mol/L , what is the value of Kc for this temperature.**

**Kc =**

1. **At 500 K. 1.0 mol of ONCI(g) is introduced into a one - liter container. At equilibrium the ONCI(g) is 9.0% dissociated:**

**2 ONCI(g)** ⮀ **2 NO(g) + CI2(g)**

**Calculate the value of Kc for equilibrium at 500 K.**

**The concentration of ONCI(g) is 1 mol/L since ONCl is 9.0% dissociated,**

**Number of moles dissociated** 

**The concentration of ONCI at equilibrium, therefore, is**

 **[ONCI] = 1.0 mol/L - 0.09 mol/L = 0.91 mol/L**

**We can derive the amounts of CI2 from the coefficient of the chemical equation: 2 ONCI** ⮀  **2NO + CI2
 2 x x
 0.09 mol 0.045 mol**

|  |  |
| --- | --- |
|  | **2ONCI ⮀ 2NO + CI2** |
| **at start**  |  | **-** | **-** |
| **Change**  |  | **+ 0.09** | **+ 0.045** |
| **at equilibrium** | **0.91** | **0.09** | **0.045** |

**Kc =**=

1. **For the reaction 2 SO3(g)** ⮀ **2 SO2(g) + O2(g) at 1100 K**

 **Kc is 0.0271 mol/L. what is Kp at same temperature.**

**Δn = 3-2 =1**

**Kp = Kc (RT)+1**

**= 0.0271 mol/L x (0.0821 L. atm / K. mol) (1100 K) = 2.45 atm**

1. **What is Kc for the reaction? N2(g) + 3 H2(g)** ⮀ **2 NH3(g), At 500°C if Kp is 1.5 x 10-5 / atm-2 at this temperature.**

**Δn = 2 - 4 = -2 T = 273 + 500 = 773 K**

**Kp = Kc (RT)Δn Kp = Kc (RT)-2**

**Kc = Kp (RT)2 =** **x [0.0821.atm/K.mol x 773K]2**

**= (1.5 x I0-5/atm2) (4.03 x 103L2. atm2/mol2) = 6.04 x 10-2L2/ mol2**

1. **For the reaction 2 SO2(g) + O2(g)** ⮀ **2 SO3(g), at 827°C, kc is 36.9 L / mol. If 0.05 mol of SO2(g), 0.03 mol of O2(g), and 0.125 mol of SO3(g) are mixed in a 1.0 L container at 827°C, in what direction will the reaction proceed?**

**Q =**

1. **Kc for the HI equilibrium at 425°C is 54.5: H2(g) + I2(g)** ⮀ **2 HI(g) A quantity of HI(g) is placed in a 1.01 container and allowed to come to equilibrium at 425°C What are the concentrations of H2(g) and I2(g) in equilibrium with 0.5 mol/L of HI(g)**

 **The [H2] = [I2] at equilibrium since they are produced in equal amounts by decomposition of HI(g) ∴[H2] = [I2] = x [HI] = 0.5 mol/L**

  **x2 =** 

**∴ x = 0.068 mol/L**

**The equilibrium concentration is: [HI] = 0.5 mol/ L [H2] = [I2] = 0.068 mol/ L**

1. **For the reaction H2(g) + CO2(g)** ⮀ **H2O(g) + CO(g)  kc is 0.771 at 750°C. If 0.01 mol of H2 and 0.01 mol of CO2 are mixed in 1 liter container at 750°C, what are the concentrations of all substances present at equilibrium?**

**If x mol of H2 reacts with x mol of CO2 out of the total amount supplied, x mol H2O and x mol CO will be produced. Hence**

|  |  |
| --- | --- |
|  |  **H2(g) + CO2(g)** ⮀ **H2O(g)**  + **CO(g)** |
| **At start** | **0.01 mol/L** | **0.01 mol/L** | **----** | **-----** |
| **Change** | **- x** | **- x** | **+ x** | **+ x** |
| **at equilibrium** | **0.01 -x** | **0.01 -x** | **X** | **X** |

 

**If we extract the square root of both sides of this equation:**

**∴**  X = 0.0878 – 0.878 X X= 0.00468 mol/l

At equilibrium, therefore

**[H2] = [CO2] = 0.01 mol/L - 0.00468 mol/L**

**= 0.0053 mol/L**

**[H2O] = [CO] = 0.00468 mol/L**

1. **For the reaction C(s) + CO2(g)** ⮀ **2 CO(g) Kp is 167.5 atm at 1000°C. What is the partial pressure of CO(g) in an equilibrium system in which the partial pressure ofCO2(g) is 0.1atm?**



PCO2= 16.8 PCO= 4.10 atm

1. **Kp for the equilibrium: FeO(s) + CO(g)** ⮀ **Fe(s) + CO2(g)**

**at 1000°C is 0.403. If CO(g) at a pressure of 1.0 atm, and excess FeO(s) are placed in a container at 1000°C, what are the pressures of CO(g) and CO2(g) when equilibrium is attained?**

**Let x equal the partial pressure of CO2 when equilibrium is attained**

|  |  |
| --- | --- |
|  | **FeO(s) + CO(g)** ⮀ **Fe(s) + CO2(g)** |
| **At start** **Change** | **1.0 atm****- x** | **---****+ x** |
| **At equilibrium** | **1.0 - x atm** | **x** |



X = PCO2 = 0.287 atm

1.0 – X = Pco = 0.713

1. **What are [H+] and [OH-] in a 0.02 M HCI solution.**

**Since HCI is a strong electrolyte ∴[H+] = 0.02M**

**[H+] [OH-] = 1.0 x 10-14 (2 X 10-2) [OH-] = 1.0 x 10-14**

**[OH-] =** 

1. **What are [H+] and [OH-] in a 0.005 M solution of NaOH?**

**NaOH is a strong electrolyte, NaOH → Na++ OH-**

**Therefore, [OH-] = 5.0 x 10-3 M [H+] [OH-] = 1.0 x 10-14**

 **[H+] [5.0 x 10-3] = 1.0 x 10-14 [H+] =** 

1. **What is the pH and pOH of a solution that is 0.05 M in H+?**

 **[H+] = 5.0 x 10-2 M pH = - log 5.0 x 10-2 = 1.3 pOH = 14-1.3 = 12.7**

1. **What is the pH of a solution for which [OH] = 0.03M**

 **[OH-] = 3.0 x 10-2 pOH = - log [OH-] = - log 3.0 x 10-2 = 1.52**

**pH = 14 - 1.52 = 12.48**

1. **What is the [H+] of a solution with a pH of 10.6.**

 **[H+] = antilog - 10.6 = 2.5 x 10-11 M**

1. **Nicotinic acid is a monoprotic acid with the formula HC6H4NO2. A solution that is 0.012 M nicotinic acid has a pH of 3.39 at 25°C. Calculate the acid ionization constant, Ka, for this acid? and what is the degree of ionization of nicotinic acid in this solution.**

**1) Ka =**  **or [H+] =** **[H+]2 = Ka Ca**
**∴ Ka =** 

**2) The value of [H+] can be obtained from the pH**

 **[H+] = antilog [-pH] = antilog - 3.39 = 4.1 x 104 M**

**∴Ka =** 

**Degree of ionization =** 

**The percent ionization = 0.034 x 100 = 3.4%**

1. **A student prepared a 0.01 M NH3 solution, and found that NH3 has undergone 4.2% ionization. Calculate the Kb for NH3, and the pH of the solution.**

**Answer Kb = 1.8 x 10-5**

1. **What are the concentration at 25°C of nicotinic acid, hydrogen ion, and nicotinate ion in a solution of 0.1 M nicotinic acid? What is the pH the solution? What is the degree o f ionization of the acid? Ka = 1.4 x 10-5**

**Answer: pH = 2.92 Degree of ionization = 0.012**

1. **What is the pH of a solution that contains, 0.1 M HCl and 0.1 M acetic acid ? Ka (acetic acid) = 1.8 x 10-5**

**The solution contains 0.1 M H+ (from the strong acid HCI) + [H+] from the weak acetic acid.**

**[H+] from acetic acid =** 

**=****1.34 x 10-3**

**We see that [H+] from acetic acid is very small compared to 0.1, so in the solution [H+] ≈ 0.1 M. This gives the pH of 1.0.**

**Complete:**

1. **The reaction between an acid and base in which water is produced is called ……………………..**
2. **An acid base reaction always proceeds from the ……………. acid and base to the weaker ……………….. acid and base.**
3. **In the Lewis concept, the formation of …………………is the basis for defining acid – base reaction**
4. **All …………… processes tend to attain a state of equilibrium**
5. **The addition of a catalyst causes a system to achieve ……………… ………but does not alter the position of equilibrium.**
6. **Le Chatelier's Principle, states that:**

**……………………………………………………………………………**

**……………………………………………………………………………**

1. **Theory can be defend as "……………**

**……………………………………………………………………………**

1. **1st Law of Thermodynamics state that:**

**………………………………………………………………………**

1. **Kirchoff's equations at constant at constant pressure = …………………… And at constant volume=**
2. **Thermochemistry it is concerns the study …………………….**
3. **Heat of transition is …………………**
4. **Real gases is derived from ideal at ……………………. And ……….**
5. **The pressure of gas equals ………….. and it is measured by ……………… in units …………., ……………..**
6. **Thermochemical equations It must essentially:**

**a)………………………………………………………………………**

**b)………………………………………………………………………**

**c)………………………………………………………………………**

1. **Compare between types of solution deviate from ideal behavior**

|  |  |
| --- | --- |
| **Negative deviation**  | **Positive deviation**  |
| ***a)*** |  |
| **b)** |  |
| **c)** |  |
| **d)** |  |

1. **Describe what a gas is according to kinetic theory:**

**a)………………………………………………………………………**

**b)………………………………………………………………………**

**c)………………………………………………………………………**

**d)………………………………………………………………………**

**e) ………………………………………………………………………**

**11. Define and mention laws of:**

**a)Graham’s Law (Molecular Effusion and Diffusion):**

**……………………………………………………………………………**

**……………………………………………………………………………**

**b) Daltons Law (Gas Mixtures and Partial Pressures):**

**……………………………………………………………………………**

**……………………………………………………………………………**

**c) Henry’s Law –**

**……………………………………………………………………………**

**……………………………………………………………………………**

**d) Raoult’s Law**

**……………………………………………………………………………**

**……………………………………………………………………………**

**……………………………………………………………………………**

1. **Hess's law of constant heat summation**

**……………………………………………………………………………**

**……………………………………………………………………………**

**Illustrate Hess’s law by burning of carbon to CO2…………..**

**……………………………………………………….**

**…………………………………………………………**