Solution

A solution of glucose in water is labelled as 10% w/ w, what would be the molality and mole fraction of each component in the solution? If the density of solution is 1.2 g mL^{-1}, then what shall be the molarity of the solution?

Ans: Mass of solute (W_B) = 10 g

Mass of solvent (W_A) = 90 g

Molar mass of solute (M_B) = 180 g/mol

Density of solution = 1.2 g/mL

Molality (m) = $\frac{W_B}{M_B} \times \frac{1000}{W_A} = \frac{10}{180} \times \frac{1000}{90} = 0.62$ m

Density = $\frac{Mass}{Volume} \Rightarrow$ Volume = $\frac{Mass}{Density} = \frac{100}{1.2}$

Molarity = $\frac{W_B}{M_B} \times \frac{1000}{v(ml)} = \frac{10}{180} \times \frac{1000}{100} \times 1.2 = 0.66$ M

H_2S, a toxic gas with rotten egg like smell, is used for the qualitative analysis. If the solubility of H_2S in water at STP is 0.195 m, calculate Henry's law constant.

Answer

It is given that the solubility of H_2S in water at STP is 0.195 m, i.e., 0.195 mol of H_2S is dissolved in 1000 g of water.

Moles of water = $\frac{1000g}{18g\ mol^{-1}}$

= 55.56 mol

Mole fraction of H_2S, x = $\frac{\text{Moles of } H_2S}{\text{Moles of } H_2S + \text{Moles of } water}$

= $\frac{0.195}{0.195 + 55.56}$

= 0.0035

At STP, pressure (p) = 0.987 bar

According to Henry's law: $p = K_H x$

$\Rightarrow K_H = \frac{p}{x}$

$= \frac{0.987}{0.0035}$

= 282 bar

Calculate **(a)** molality **(b)** molarity and **(c)** mole fraction of KI if the density of 20% (mass/mass) aqueous KI is 1.202 g mL^{-1}.

Answer

(a) Molar mass of KI = 39 + 127 = 166 g mol^{-1}

20% (mass/mass) aqueous solution of KI means 20 g of KI is present in 100 g of solution.

That is,

20 g of KI is present in (100 − 20) g of water = 80 g of water

Therefore, molality of the solution = $\frac{\text{Moles of KI}}{\text{Mass of water in kg}}$

= $\frac{\frac{20}{166}}{0.08}$ m

= 1.506 m

= 1.51 m (approximately)

(b) It is given that the density of the solution = 1.202 g mL^{-1}

Volume of 100 g solution = $\frac{Mass}{Density}$

= $\frac{100g}{1.202g\ \text{mL}^{-1}}$

= 83.19 mL

= 83.19×10^{-3} L

Therefore, molarity of the solution = $\frac{\frac{20}{166} mol}{83.19\times10^{-3} L}$

= 1.45 M

(c) Moles of KI = $\frac{20}{166}$ = 0.12 mol

Moles of water = $\frac{80}{18}$ = 4.44 mol

Therefore, mole fraction of KI = $\frac{\text{Moles of KI}}{\text{Moles of KI} + \text{Moles of water}}$

= $\frac{0.12}{0.12+4.44}$ = 0.0263

An aqueous solution of 2% non-volatile solute exerts a pressure of 1.004 bar at the normal boiling point of the solvent. What is the molar mass of the solute?

Answer
Here,
Vapour pressure of the solution at normal boiling point (p_1) = 1.004 bar
Vapour pressure of pure water at normal boiling point $\left(p_1^0\right)$ = 1.013 bar
Mass of solute, (w_2) = 2 g
Mass of solvent (water), (w_1) = 98 g
Molar mass of solvent (water), (M_1) = 18 g mol^{-1}
According to Raoult's law,

$$\frac{p_1^0 - p_1}{p_1^0} = \frac{w_2 \times M_1}{M_2 \times w_1}$$

$$\Rightarrow \frac{1.013 - 1.004}{1.013} = \frac{2 \times 18}{M_2 \times 98}$$

$$\Rightarrow \frac{0.009}{1.013} = \frac{2 \times 18}{M_2 \times 98}$$

$$\Rightarrow M_2 = \frac{1.013 \times 2 \times 18}{0.009 \times 98}$$

= 41.35 g mol^{-1}
Hence, the molar mass of the solute is 41.35 g mol^{-1}.

Boiling point of water at 750 mm Hg is 99.63°C. How much sucrose is to be added to 500 g of water such that it boils at 100°C. Molal elevation constant for water is 0.52 K kg mol^{-1}.
Answer

Here, elevation of boiling point ΔT_b = (100 + 273) − (99.63 + 273)
= 0.37 K
Mass of water, w_l = 500 g
Molar mass of sucrose ($C_{12}H_{22}O_{11}$), M_2 = 11 × 12 + 22 × 1 + 11 × 16
= 342 g mol^{-1}
Molal elevation constant, K_b = 0.52 K kg mol^{-1} We know that:

$$\Delta T_b = \frac{k_b \times 1000 \times w_2}{M_2 \times w_1}$$

$$\Rightarrow w_2 = \frac{\Delta T_b \times M_2 \times w_1}{1000 \times k_b}$$

$$= \frac{0.37 \times 342 \times 500}{1000 \times 0.52}$$ = 121.67 g (approximately)

Hence, 121.67 g of sucrose is to be added.

Calculate the osmotic pressure in pascals exerted by a solution prepared by dissolving 1.0 g of polymer of molar mass 185,000 in 450 mL of water at 37°C.
Answer
It is given that:

Volume of water, V = 450 mL = 0.45 L

Temperature, T = (37 + 273)K = 310 K

Number of moles of the polymer, $n = \frac{1}{185000}$ mol
We know that:

Osmotic pressure, $\pi = \frac{n}{V} \mathrm{R}T$

$= \frac{1}{185000} \text{mol} \times \frac{1}{0.45 \text{ L}} \times 8.314 \times 10^3 \text{ Pa L K}^{-1}\text{mol}^{-1} \times 310 \text{ K}$

= 30.98 Pa

= 31 Pa (approximately)

What tow difference between ideal solution and non-ideal solution.

Define the term: Mole fraction, molarity and molality.

What role does the molecular interaction play in solution of alcohol and water?

Electrochemistry

Can you store copper sulphate solution in a Zn pot?

Answer
Zinc is more reactive than copper. Therefore, zinc can displace copper from its salt solution. If copper sulphate solution is stored in a zinc pot, then zinc will displace copper from the copper sulphate solution.
Hence, copper sulphate solution cannot be stored in a zinc pot.

Calculate the potential of hydrogen electrode in contact with a solution whose pH is 10.

Given the standard electrode potentials,

K^+/K= - 2.93 V, Ag^+/Ag = 0.80 V, Hg^{2+}/Hg = 0.79 V, Mg^{2+}/Mg = -2.37 V, Cr^{3+}/Cr = -0.74 V Arrange these metals in increasing order of their reducing power.

Write the Nernst equation and emf of the following cells at 298 K:
(i) Mg(s) | Mg2+(0.001M) || Cu2+(0.0001 M) | Cu(s)
(ii) Fe(s) | Fe2+(0.001M) || H+(1M)|H2(g)(1bar) | Pt(s)

Answer
(i) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E^{\ominus}_{\text{cell}} - \frac{0.0591}{n}\log\frac{\left[\text{Mg}^{2+}\right]}{\left[\text{Cu}^{2+}\right]}$$

$$= \left\{0.34 - (-2.36)\right\} - \frac{0.0591}{2}\log\frac{.001}{.0001}$$

$$= 2.7 - \frac{0.0591}{2}\log 10$$

= 2.7 − 0.02955
= 2.67 V (approximately)
(ii) For the given reaction, the Nernst equation can be given as:

$$E_{\text{cell}} = E^{\ominus}_{\text{cell}} - \frac{0.0591}{n}\log\frac{\left[\text{Fe}^{2+}\right]}{\left[\text{H}^{+}\right]^2}$$

$$= \left\{0 - (-0.44)\right\} - \frac{0.0591}{2}\log\frac{0.001}{1^2}$$

$$= 0.44 - 0.02955(-3)$$

= 0.52865 V
= 0.53 V (approximately)

The cell in which the following reactions occurs:

$$2\text{Fe}^{3+}{}_{(aq)} + 2\text{I}^{-}{}_{(aq)} \rightarrow 2\text{Fe}^{2+}{}_{(aq)} + \text{I}_{2(s)}$$

Has E^{0}_{cell} = 0.236 V at 298 K.

Calculate the standard Gibbs energy and the equilibrium constant of the cell reaction.
Answer
Here, *n* = 2, $E^{\ominus}_{\text{cell}} = 0.236\text{ V},$ T = 298 K
We know that:

$$\Delta_r G^{\ominus} = -n\text{F}E^{\ominus}_{\text{cell}}$$

= −2 × 96487 × 0.236
= −45541.864 J mol^{-1}
= −45.54 kJ mol^{-1}

Again, $\Delta_r G^{\ominus} =$ −2.303 R *T* log *K* c

$$\Rightarrow \log K_c = -\frac{\Delta_r G^{\ominus}}{2.303\ \text{R}T}$$

$$= -\frac{-45.54 \times 10^3}{2.303 \times 8.314 \times 298}$$

= 7.981
*K*c = Antilog (7.981)
= 9.57 × 10^{7}

Define conductivity and molar conductivity for the solution of an electrolyte. Discuss their variation with concentration.

Answer

Conductivity of a solution is defined as the conductance of a solution of 1 cm in length and area of cross-section 1 sq. cm. The inverse of resistivity is called conductivity or specific conductance. It is represented by the symbol *κ*. If ρ is resistivity, then we can write:

$$\kappa = \frac{1}{\rho}$$

The conductivity of a solution at any given concentration is the conductance (*G*) of one unit volume of solution kept between two platinum electrodes with the unit area of cross-section and at a distance of unit length.

$$G = \kappa \frac{a}{l} = \kappa \cdot 1 = \kappa$$

(Since *a* = 1, *l* = 1)

Conductivity always decreases with a decrease in concentration, both for weak and strong electrolytes. This is because the number of ions per unit volume that carry the current in a solution decreases with a decrease in concentration.

Molar conductivity:

Molar conductivity of a solution at a given concentration is the conductance of volume V of a solution containing 1 mole of the electrolyte kept between two electrodes with the area of cross-section *A* and distance of unit length.

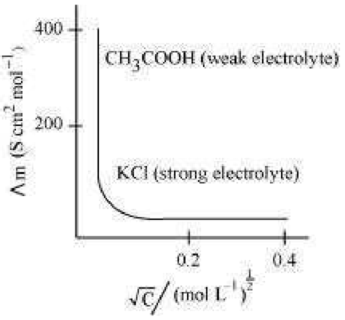
$$\Lambda_m = \kappa \frac{A}{l}$$

Now, *l* = 1 and *A* = V (volume containing 1 mole of the electrolyte).

$$\Lambda_m = \kappa V$$

Molar conductivity increases with a decrease in concentration. This is because the total volume V of the solution containing one mole of the electrolyte increases on dilution.

The variation of Λ_m with $\sqrt{c}$ for strong and weak electrolytes is shown in the following plot:



The resistance of a conductivity cell containing 0.001M KCl solution at 298 K is 1500 Ω. What is the cell constant if conductivity of 0.001M KCl solution at 298 K is 0.146 × 10−3 S cm−1.

Answer

Given,

Conductivity, *κ* = 0.146 × 10−3 S cm−1

Resistance, *R* = 1500 Ω

Cell constant = *κ* × *R*

= 0.146 × 10−3 × 1500

= 0.219 cm−1

A solution of Ni(NO3)2 is electrolysed between platinum electrodes using a current of 5 amperes for 20 minutes. What mass of Ni is deposited at the cathode?

Answer

Given,

Current = 5A
Time = 20 × 60 = 1200 s
Charge = current × time
= 5 × 1200
= 6000 C
According to the reaction,

$$Ni^{2+}_{(aq)} + 2e^- \longrightarrow \underset{58.7\ g}{Ni_{(s)}}$$

Nickel deposited by 2 × 96487 C = 58.71 g

Therefore, nickel deposited by 6000 C $= \frac{58.71 \times 6000}{2 \times 96487}\ g$
= 1.825 g
Hence, 1.825 g of nickel will be deposited at the cathode.

Suggest two materials other than hydrogen that can be used as fuels in fuel cells.

Answer

Methane and methanol can be used as fuels in fuel cells.

Explain how rusting of iron is envisaged as setting up of an electrochemical cell.

Answer

In the process of corrosion, due to the presence of air and moisture, oxidation takes place at a particular spot of an object made of iron. That spot behaves as the anode. The reaction at the anode is given by,

$$Fe_{(s)} \longrightarrow Fe^{2+}_{(aq)} + 2e^-$$

Electrons released at the anodic spot move through the metallic object and go to another spot of the object.

There, in the presence of H^+ ions, the electrons reduce oxygen. This spot behaves as the cathode. These H^+ ions come either from H_2CO_3, which are formed due to the dissolution of carbon dioxide from air into water or from the dissolution of other acidic oxides from the atmosphere in water.

The reaction corresponding at the cathode is given by,

$$O_{2(g)} + 4H^+_{(aq)} + 4e^- \longrightarrow 2H_2O_{(l)}$$

The overall reaction is:

$$2Fe_{(s)} + O_{2(g)} + 4H^+_{(aq)} \longrightarrow 2Fe^{2+}_{(aq)} + 2H_2O_{(l)}$$

Also, ferrous ions are further oxidized by atmospheric oxygen to ferric ions. These ferric ions combine with moisture, present in the surroundings, to form hydrated ferric oxide $(Fe_2O_3 . xH_2O)$ i.e., rust.

Hence, the rusting of iron is envisaged as the setting up of an electrochemical cell.