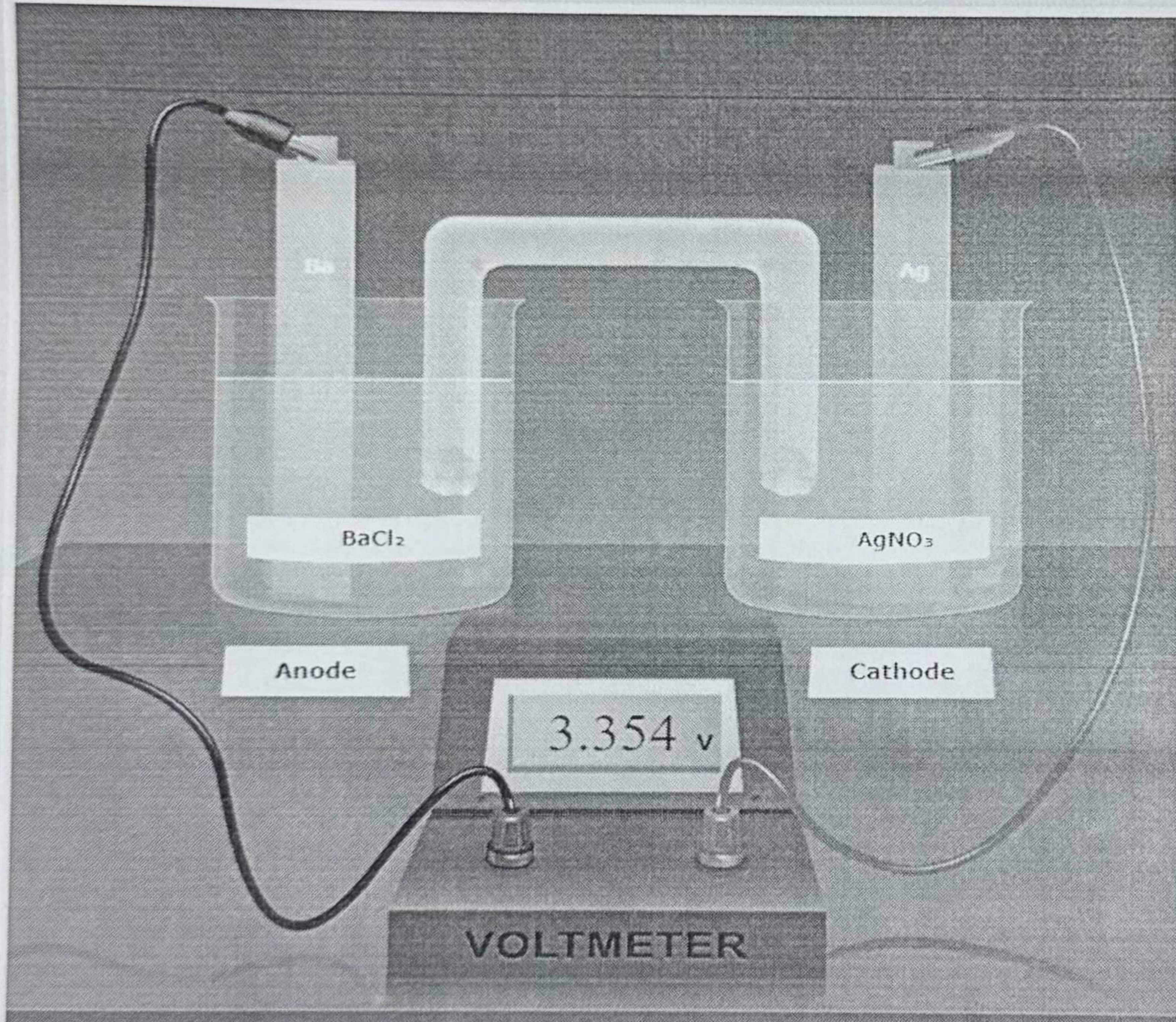
Ba
Ag
BaCl₂
AgNO₃
Anode
Cathode
3.354 v
VOLTMETER

Name:

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Department of Science and Humanities
Applied Chemistry Laboratory

Subject: Engineering Chemistry

Therefore, EMF of the cell = 3.354 V

The Gibb's free energy change of the cell reaction, ΔG = Negative

The Equilibrium constant of the cell reaction, K = >1

The spontaneity of the cell reaction = Spontaneous

Anode : $Ba(s) \rightarrow Ba^{2+}(aq) + 2e^-$

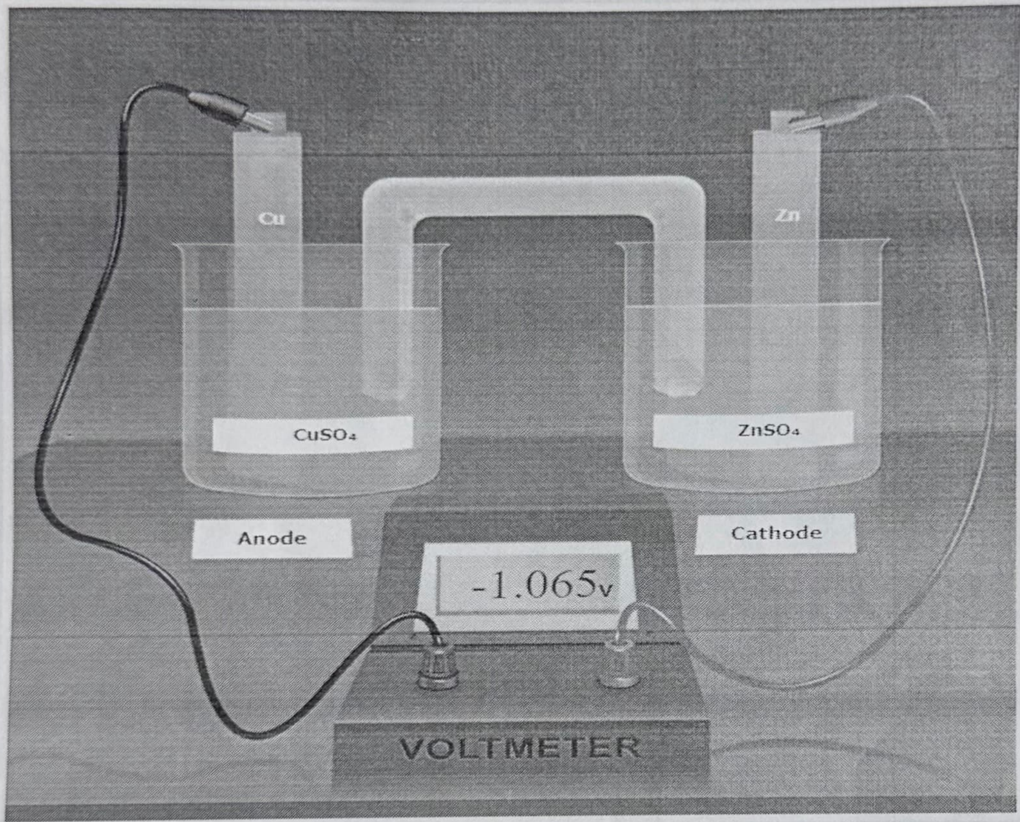
Cathode : $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$

Overall Reaction: $Ba(s) + 2Ag^+(aq) \rightarrow Ba^{2+}(aq) + 2Ag(s)$

Cell Reaction:

$Ba_{(s)} \mid BaCl_{2\,(aq)} \mid\mid AgNO_{3(aq)} \mid Ag_{(s)}$

2)



Temperature = 20° C
Cathode used = Zinc
Concentration of electrolyte = 10 M
Anode used = Copper
Concentration of electrolyte = 3 M
Therefore, EMF of the cell = -1.065 V

The Gibb's free energy change of the cell reaction, ΔG = Positive

The Equilibrium constant of the cell reaction, K = >1

The spontaneity of the cell reaction = Non-Spontaneous

Anode : $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

Cathode : $Zn^{2+}(aq)+2e^{-}\text{-->}Zn(s)$

Overall Reaction: $Cu(s) + Zn^{2+}(aq) \text{-->} Cu^{2+}(aq) + Zn(s)$

Cell Representation:

$Cu_{(s)} \mid CuSO_{4(aq)} \mid\mid ZnSO_{4(aq)} \mid Zn_{(s)}$

Calculations:
1)

$E^{o}_{Cell} = E^{o}_{cathode} - E^{o}_{anode}$

$E^{o}_{Cell} = (0.80) - (-2.90)$

$E^{o}_{Cell} = 3.7$

E cell = E°cell-(RT/nF) ln ($[Ba^{2+}]/[Ag^{+}]^2$)

$$E_{cell} = 3.7 - \frac{2.303 \times 8.314 \times 303.15}{2 \times 96500} \; log \frac{0.01^2}{10^2}$$

$E_{cell} = 3.7 - 0.288$

$\underline{\mathbf{E_{cell} = 3.412}}$

$Electric\ work\ done = nFE_{cell}$

$-\Delta G = nFE_{cell}$

Therfore,

$\Delta G = -nFE_{cell}$

$= -2 \times 96500 \times 3.412$

$\Delta G = -658.516$

$-\Delta G^0 = nFE^0_{cell}$

$\Delta G^o = -nFE^o_{cell}$

$= -2 \times 96500 \times 3.7$

$\Delta G^o = -714.100$ kJ

$\Delta G^0 = -RT\ln K$

$-714.100 = -8.314 \times 303.15 \ln K$

$\ln K = -714.100 / -2520.38$

$\ln K = 0.28$

$K = 1.32$

2)
$E^o_{Cell} = E^o_{cathode} - E^o_{anode}$

$E^o_{Cell} = (-0.76) - (0.34)$

$E^o_{Cell} = -1.1$

$E_{cell} = E^\circ_{cell} - (RT/nF) \ln ([Cu^{2+}]/[Zn^{2+}])$

$E_{cell} = -1.1 - \frac{2.303 \times 8.314 \times 293.15}{2 \times 96500} \log \frac{3^2}{10^2}$

E_{cell} =-1.1 – (-0.069)

E_{cell} = -1.1 + 0.069

E_{cell} = - 1.031

$$\text{Electric work done} = nFE_{cell}$$

$$-\Delta G = nFE_{cell}$$

Therfore,

$\Delta G = -nFE_{cell}$

$= -2 \times 96500 \times -1.031$

ΔG = 198.983 kJ

$$-\Delta G^0 = nFE^0_{cell}$$

$\Delta G^o = -nFE^o_{cell}$

$= -2\times 96500 \times -1.1$

ΔG^o = 212.3 kJ

$$\Delta G^0 = -RT\ln K$$

212.3 = – 8.314 x 293.15 lnK

lnK = 212.3 / – 2437.24

lnK = -0.08

K = 0.92

Assignment:

1. What is the electrode potential of electrode in which conc. of Mg2+ is 0.01 M? =-2.36 V.

Ans: EMg2+/Mg = E0Mg2+/Mg +0.0591/n x log [Mg2+]/[Mg]

EMg2+/Mg = -2.36 + 0.0591/2 x log (0.01/1)

EMg2+/Mg = -2.42 V

2. 100 mL of a neutral solution containing 0.2 g of copper was electrolysed till the whole of copper was deposited. The current strength was maintained at 1.2 and the volume of solution was maintained at 100 mL. Assuming 100% efficiency, find out the time taken for deposition of copper. [At.wt of copper =63.58]

Ans: The quantity of electricity passed (100% efficiency) = Q(C)
=I(A)×t(s)
= 1.2 A × t(s)
= **1.2t C.**

The atomic weight of copper = 63.58 g/mol.
Moles of electrons passed = Q / 96500 = 1.2t / 96500
The mass of copper deposited = 0.2 g.
Hence,
0.2 g = atomic weight of copper × mole ratio × moles of electrons passed
0.2 g = 63.58 × 1/2 × 1.2t / 96500
t=506s.

3. The reduction potentials of and electrode are 0.34 V and 0.80 V respectively. Construct a galvanic cell using these values. For what concentration of Ag+ ions will e.m.f of the cell at 25 0C be zero if conc. of Cu2+ is 0.01 M. Given, ECu2+/Cu∘=0.34 volt and EAg+/Ag∘=0.80 volt.

Ans: Given, $E^0 Cu^{2+}/Cu = 0.34$ V and $E^0 Ag^+/Ag = 0.80$ V.
The standard emf will be positive if Cu/Cu^{2+} is anode and Ag^+/Ag is cathode. The cell can be represented as:
$Cu \mid Cu^{2+} \parallel Ag^+ \mid Ag$

The cell reaction is,
$Cu + 2Ag^+ \rightarrow Cu^{2+} + 2Ag$

E^0cell = Oxid. potential of anode + Red. potential of cathode
$= -0.34 + 0.80$
$= 0.46$ volt
Applying the Nernst equation,
$Ecell = E^0cell - 0.0591/2 \times \log [Cu^{2+}] / [Ag^+]^2$

When, Ecell=0
$E^0cell = 0.0591/2 \times \log [Cu^{2+}] / [Ag^+]^2$
or $\log [Cu^{2+}] / [Ag^+]^2 = 0.462 \times 2 / 0.0591 = 15.6345$

$[Cu^{2+}] / [Ag^+]^2 = 4.3102 \times 10^{15}$
$[Ag^+]^2 = 0.01 / 4.3102 \times 10^{15}$
$= 0.2320 \times 10^{-17}$
$= 2.320 \times 10^{-18}$
$[Ag^+] =$ **1.523×10^{-9}M.**

4. Calculate the maximum work that can be obtained from the Daniel cell . Given that and are -0.76 and +0.34 V respectively.

Ans: According to given cell reaction
$E^0anode(Zn^{+2}/Zn) = -0.76V$
$E^0Cathode(Cu^{+2}/Cu) = +0.34V$
n=2

Now, W electrical $= -nF\ E^0cell$
$= -2 \times 96500 \times [0.34-(-0.76)]$
$= -2 \times 96599 \times 1.1$
= **−212.3K Joule.**

5. For Daniel Cell involving the cell reaction, ,the standard free energies of Zn (s), Cu (s), Cu2+ (aq) and Zn2+ (aq) are 0, 0, 64.4 KJmol-1 and -154.0 KJmol-1 respectively. Calculate the standard EMF of the cell.

Ans.: $\because \Delta G0=(GZn2+0-GCu2+0)=(-154-64.4)KJ/Mole$

$\therefore(-2\times E0\times F)=(-218.4\times 103)$

$\therefore$ **E0=1.1316volt**

Result:

1)

1. The EMF of the cell is **3.412V**

2.The Gibb's free energy change of the cell reaction, ΔG = **-658.516 kJ**

3.The Equilibrium constant of the cell reaction, K = **1.32**

4. The spontaneity of the cell reaction = **Spontaneous**

2)

1. The EMF of the cell is **-1.031V**

2. The Gibb's free energy change of the cell reaction, ΔG = **198.983 kJ**

3. The Equilibrium constant of the cell reaction, K = **0.92**

4. The spontaneity of the cell reaction = **Non-Spontaneous**

Conclusion:

Thus, we have measured the EMF of a cell and predicted the spontaniety of the cell reaction.