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Chemistry 205 Report

Electrochemistry

Name:

Date: 22 - 03 - 2013

Partner:

Purpose:
1. To study the relative reactivity (tendency to be oxidized) of the metals Zn, Cu, Fe and hydrogen
2. To get acquainted with the use of a voltmeter
3. To measure the cell potentials of different galvanic cells
4. To compare the observed cell potentials with the literature values and expected ones from Nernst equation
5. To construct a concentration cell and measure its potential
6. To identify the reactions occuring at the anode and cathode during the electrolysis of aqueous potassium iodide solution

Part I: Redox Reactions and the Relative Activity of Metals

Table 1. **Redox Reactions and the Relative Activity of Metals**

Metal	Ion	Reaction yes or no?	Ionic Equation	Relative Reactivity
Zn	Fe^{2+}	YES	$Zn + Fe^{2+} \rightarrow Zn^{2+} + Fe$	Zn more reactive than Fe^{2+}
Zn	Cu^{2+}	YES	$Zn + Cu^{2+} \rightarrow Zn^{2+} + Cu$	Zn more reactive than Cu^{2+}
Zn	H^+	YES	$Zn + 2H^+ \rightarrow Zn^{2+} + H_2$	Zn more reactive than H^+
Fe	Zn^{2+}	NO		
Fe	Cu^{2+}	YES	$Fe + Cu^{2+} \rightarrow Cu + Fe^{2+}$	Fe is more reactive than Cu^{2+}
Fe	H^+	YES	$Fe + 2H^+ \rightarrow Fe^{2+} + H_2$	Fe more reactive than H^+
Cu	Fe^{2+}	No		
Cu	Zn^{2+}	NO		
Cu	H^+	NO		

Arrange the above metals in order of decreasing reactivity (tendency to be oxidized):

$Zn = -0.76$ V

$Fe = -0.44$

$Cu = 0.34$

$H^+ = 0$

$\Rightarrow Zn > Fe > H_2 > Cu$

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$E_{cell} = E^{\circ}_{cathode} - E^{\circ}_{anode}$

$E = E^{\circ}_{cell} - \frac{0,0257}{n} \ln \left(\frac{anode}{cathode}\right)$

Part II: Galvanic Cells

for (3b) (−1)

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Table 2. Emf of Galvanic Cells @ T=

#	Cell Diagram	E_{cell} measured	E_{cell} theoretical	Error
1	$Zn(s)|ZnSO_4(0.100M)||CuSO_4(0.100M)|Cu(s)$	1,07 V	1,1.	2,7%
2	$Zn(s)|ZnSO_4(0.100M)||Pb(NO_3)_2(0.100M)|Pb(s)$	0,61 V	0,63	3,17%
3	$Zn(s)|Zn(NO_3)_2(0.100M)||AgNO_3(0.100M)|Ag(s)$	1,23 V	1,56	21,15%
4	$Cu(s)|Cu(NO_3)_2(0.100M)||AgNO_3(0.100M)|Ag(s)$	0.38 V	0,46	17,4%

Table 3. Comparison of Measured Emf and Calculated Values from Nernst Equation @ T=

$E = E^{\circ}_{cathode} - E^{\circ}_{anode}$
$= 0,8 + 0,76$
$= 1,56$

#	Cell Diagram	E_{cell} measured	E_{cell} calculated	Error
1a	$Zn(s)|ZnSO_4(0.100M)||CuSO_4(0.0010M)|Cu(s)$	0,97	1,1 − 0,059 = 1,04 V	6,7%
1b	$Zn(s)|ZnSO_4(0.0010M)||CuSO_4(0.100M)|Cu(s)$	1,00	1,1 − (−0,059) = 1,15 V	13%
2a	$Zn(s)|ZnSO_4(0.100M)||Pb(NO_3)_2(0.0010M)|Pb(s)$ −0,76 −0,13	0.53	0,63 − 0,059 = 0,57 V	7,02%
2b	$Zn(s)|ZnSO_4(0.0010M)||Pb(NO_3)_2(0.100M)|Pb(s)$	0,54	0,63 + 0,059 = 0,68 V	20,6%
3a	$Zn(s)|Zn(NO_3)_2(0.100M)||AgNO_3(0.0010M)|Ag(s)$ −0,76 0,8	0,96	1,56 − 1.41	31,9%
3b	$Zn(s)|Zn(NO_3)_2(0.0010M)||AgNO_3(0.100M)|Ag(s)$	1,16	1,56 − 1.50	22,6%
4a	$Cu(s)|Cu(NO_3)_2(0.100M)||AgNO_3(0.0010M)|Ag(s)$ 0,34 0,8	0,87	0,46 − 0,31 V	12,9%
4b	$Cu(s)|Cu(NO_3)_2(0.0010M)||AgNO_3(0.100M)|Ag(s)$	0,47	0,46 0,49 V	4,08%

$E = E^{\circ} - \frac{0,0257}{2} \times \ln \frac{(0,1)}{(0,001)^2}$

Part III: Concentration Cells

Table 4: Emf of Concentration Cell

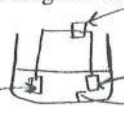
#	Cell Diagram	E_{cell} measured	E_{cell} calculated	Error
	$Cu(s)||Cu(NO_3)_2(0,001M)||Cu(NO_3)_2(0,1M)|Cu$	0,06	+0,06	0%

$E = E^{\circ} - \frac{0,0257}{2} \ln \left(\frac{0,001}{0,1}\right)$

$= 0 - (-0,06) = 0,06\ V$

Part IV: Electrolytic Cells

1. Draw a diagram for the electrolysis set-up labeling all parts.



2. Write your observations and conclusions.

At Anode. $2I^- \longrightarrow I_{2(s)} + 2e^-$
Blue color observed at anode $\Rightarrow$ (oxidation)

At Cathode $2H_2O + 2e^- \longrightarrow H_2(g) + 2OH^-$
Pink color observed at cathode (reduction)

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3. Write the electrode reactions :
<u>Anode:</u> $2I^- \longrightarrow I_{2(s)} + 2e^-$

<u>Cathode:</u> $2H_2O + 2e^- \longrightarrow H_{2(g)} + 2OH^-$

Questions

1. Write the cell reaction that occurs spontaneously in a cell that uses Ag/Ag^+ and Al/Al^{3+} half cell reactions under standard state conditions. Calculate the standard emf of the cell; then write the cell notation.

$Ag^+ + 1e^- \longrightarrow Ag \quad E^0 = 0,80V$ (reduction) $\rightarrow Ag^+ + 1e^- \rightarrow Ag$

$Al^{3+} + 3e^- \longrightarrow Al \quad E^0 = -1,66V$ (oxidation) $\rightarrow Al \longrightarrow Al^{3+} + 3e^-$

$E_{cell} = E^0_{cathode} - E^0_{anode}$

$= E^0 Ag/Ag^+ - E^0_{anode}(Al/Al^{3+})$

$= 0,80 - (-1,66) = \boxed{2,46V}$

$Al/Al^{3+} || Ag^+/Ag$
Anode Cathode

2. What is the potential of a cell made up of Zn/Zn^{2+} and Cu/Cu^{2+} half cells at 25ºC if $[Zn^{2+}] = 0.25$ M and $[Cu^{2+}] = 0.15$ M?

$Zn/Zn^{2+} \quad E^0 = -0,76V$

$Cu/Cu^{2+} \quad E^0 = 0,34V$

$E_{cell} = E^0_{cathode} - E^0_{anode}$

$= 0,34 - (-0,76)$

$= \boxed{1,1V}$

$E = E^0 - \frac{0,0257}{n} \ln \frac{[Zn^{2+}]}{[Cu^{2+}]}$

$= 1,1 - \frac{0,0257}{2} \ln\left(\frac{0,25}{0,15}\right)$

$= 1,1 - 6,56 \times 10^{-3} = \boxed{1,09V}$

3. Calculate the emf of the following concentration cell:

$Mg(s)|Mg^{2+}(0.24M)|| Mg^{2+}(0.53M)|Mg(s)$

Anode. $Mg \longrightarrow Mg^{2+} + 2e^-$

Cathode $Mg^{2+} + 2e^- \longrightarrow Mg$

$E = E^0 - \frac{0,0257}{2} \ln\left(\frac{0,24}{0,53}\right)$

$= 0 - (-0,01) = \boxed{0,01V}$