

Chem_162_Lab_2_Electrochemistry_Craig_Spring_2017

Name: Aaron Ong
ID Number: 1471258

Quiz Section: ac
Lab Partner: Stephanie Park

Chem 162 Experiment #2: Electrochemistry

Note:
All sections of this report, except the structural formulas, must be typed.

Total Points = 60 (10 notebook, 50 report)

By signing below, you certify that you have not falsified data, that you have not plagiarized any part of this lab report, and that all calculations and responses other than the reporting of raw data are your own independent work. Failure to sign this declaration will result in 5 points being deducted from your report score.

Signature:

Data and Calculations

Part I: Standard Potentials

Standard Potentials vs. Silver/Silver Chloride Reference Electrode

Cell Number	Potential, V vs. Ag/AgCl (spontaneous rxn)	Cathode Electrode	Anode Electrode
1 (Ag)	0.556	Ag	AgCl
2 (Zn)	0.785	AgCl	Zn
3 (Pb)	0.362	AgCl	Pb
4 (Cu)	0.138	Cu	AgCl

Spontaneous Reaction

(between each half cell and Ag/AgCl reference)

Cell #

- 1
- 2
- 3
- 4

1	$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
2	$2AgCl(s) + Zn(s) \rightarrow Zn^{2+}(aq) + 2Cl^-(aq) + 2Ag(s)$
3	$2AgCl(s) + Pb(s) \rightarrow Pb^{2+}(aq) + 2Cl^-(aq) + 2Ag(s)$
4	$2Ag(s) + 2Cl^-(aq) + Cu^{2+}(aq) \rightarrow Cu(s) + 2AgCl(s)$

8 pts

Using the data from Table 2.1 in the lab manual, type the calculation for the expected E° value for the $Cu^{2+}/Cu - Ag/AgCl$ cell. (3 pts)

$E^{\circ}(\text{standard}) Cu^{2+} = 0.34 V$
 $E^{\circ}(\text{standard}) Ag = 0.22 V$
 $E^{\circ}(\text{cell standard}) = E^{\circ}(\text{cathode}) - E^{\circ}(\text{anode}) = 0.34 V - 0.22 V = 0.12 V$

Based on your measured potential for this cell and the literature value for the standard reduction potential for the Ag/AgCl reference electrode, what would you expect the overall potential to be for the spontaneous reaction between your Cu^{2+}/Cu electrode and a standard hydrogen electrode? Type your calculation for the expected standard reduction potential vs the SHE as well as the % error between this value and the literature value. (3 pts)

Measured cell potential from lab = $E(\text{cell}) + E(\text{anode})$
 $E(\text{anode}) = 0.220 V$
 $E(\text{cell}) = 0.138 V$
 $E(\text{cell}) + E(\text{anode}) = 0.358 V$
 $(0.358 - 0.340) / 0.340 \times 100 = 5.3\% \text{ error}$

Zinc (#2) and Lead (#3) E_{cell} , V

Expected E_{cell} (V) based on YOUR E° values

0.423

Actual value of your measurement (V)

0.748

Oxidation Half-reaction:
Reduction Half-reaction:

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
 $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

4 pts

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