GENERAL CHEMISTRY LABORATORY

CH131

LAB REPORT #10:
ELECTROCHEMISTRY EXPERIMENT

Name:____

ID#:____

TF:____

SECTION/Day/ Time:____

Result summary: Your report for lab #10 should include ONLY page 1 and 5. For the lab report preparation use table 1 with the standard reduction potentials and the experimental observation data in table 2-5.

BOSTON UNIVERSITY

DETAILED PROTOCOL OF THE EXPERIMENT

Interaction of metal and metal ions

A. In a clean well plate four wells have been filled with 0.2M $Zn(NO_3)_2$ stock solutions. Then one strip of Zn, Cu, Pb or Ag metal has been placed in a separate well with $Zn(NO_3)_2$ solution. The appearance of a chemical reaction between each metal and the ionic $Zn(NO_3)_2$ solution has been recorded in Data Table 2 below. In case no reaction has been observed **"NR"** was written in Data Table 2.

B. After that the entire experiment has been repeated with three other stock solutions: 0.2M $Cu(NO_3)_2$, 0.2M $AgNO_3$ and 0.2M $Pb(NO_3)_2$, and the appearance of a chemical reaction between each metal and each ionic solution has been recorded in Data Tables 3-5. In case no reaction has been observed **"NR"** was written in Data Tables 3-5.

C. In Data Table 6 below write the balanced net ionic equations for each metal/solution combination tested in the experiment. Use the data from Table 1 to calculate the value of the standard voltage E^0_{rxn} and the sign of ΔG^0 for each reaction for each net reaction. **Compare the calculated sign of ΔG^0 value with the presence or absence of the reaction in the well recorded in tables 2-5. Write "NR" in Data Table 6 if no reaction has been observed according data tables 2-5.**

Table 1. Standard reduction potentials E^0_{red} for the ionic redox reactions at 25^{0}C

Reduction half-reaction at 25^{0}C	**Standard reduction potential, E^0_{red}, Volt**
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	– 0.76
$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$	– 0.13
$2H^+(aq) + 2e^- \rightarrow H_2(g)$	0
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+ 0.34
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+ 0.80

Here are several examples of calculations. Example #1: Zn(s)/Cu^{2+} (aq) system

In this case, you will see a black substance on the surface of the zinc plate. It is the layer of Cu(s), which is deposited on the zinc surface as the product of the following reactions:

Oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $\quad E^0_{ox} = + 0.76V$

Reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ $\quad E^0_{red} = + 0.34V$

Net reaction: $Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$ $\quad E^0_{rxn} = 1.10 V$

As you can see from the calculations above, there is very important conclusion: a standard voltage E^0_{rxn} of this net reaction is positive! The standard free energy change ΔG^0 of the redox reaction can be written as:

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

Here n- is the number of electron moles in the redox reaction and F is a Faraday number, the absolute value for the charge of one mole of electrons, which also has a positive value. **Therefore, the net redox reactions with $E^0_{rxn} > 0$ should have $\Delta G^0 < 0$. But the negative sign of ΔG^0 means the redox reaction is spontaneous! That is why the net reaction between Zn(s) and Cu^{2+} (aq) with E^0_{rxn} = 1.10 V manifests its spontaneity by plating out Cu(s) on the surface of the Zn plate.**

Actually, you can predict the sign of ΔG^0 for the net ionic reaction without any calculations: if the spontaneous reaction has been observed - the ΔG^0 value should be negative! And vice versa – if the standard voltage for the redox reaction has a positive value, the reaction should be spontaneous!

Example #2: $Cu(s)/ Zn^{2+}$ (aq) system. In that case:

Oxidation:	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$	$E^0_{ox} = -0.34V$
Reduction:	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	$E^0_{red} = -0.76V$
Net reaction:	$Cu(s) + Zn^{2+}(aq) \rightarrow Zn(s) + Cu^{2+}(aq)$	$E^0_{rxn} = -1.10$ V

As you can see from the above analysis, E^0_{rxn} <0 for the $Zn(s)/Cu^{2+}(aq)$ system. It leads to ΔG^0 >0 for the net reaction. That is why there is no plating of copper on the Zn plate in a Cu^{2+} (aq) solution in table 2!

Finally, for the system with the same metal and metal ion (for example, $Zn(s)/Zn^{2+}(aq)$ system), you can easily calculate that $E^0_{rxn}=0$ because in that case $E^0_{ox} = - E^0_{red}$. Therefore, $\Delta G^0 = 0$ for this kind of net ionic reaction, which means the chemical system is in the state of the equilibrium: the oxidation and reduction rates are the same and, of course, in that case you shouldn't expect any visible reaction in the system.

SPECIAL CASE OF Cu(s) immersed in $Ag^+(aq)$ SOLUTION

Reduction half-reaction at 25^0C	Standard reduction potential, E^0_{red}, Volt
$Ag^+ (aq) + e^- \rightarrow Ag(s)$	+ 0.80

However, the reduction reaction $2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$ has the same reduction potential $E^0_{red} = +0.80V$ but not the double value 1.6V.

Explanation:

Of course, the double charge 2Q produces the double energy 2W in that reaction.

However, by the definition:

Standard reduction voltage is Charge transfer energy per charge:

E^0_{red} = Energy for charge transfer /Charge =2W/2Q = W/Q

– and therefore still has the same value of 0.8V.

That is why:

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$	with $E^0_{ox} = -0.34V$
$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$	with $E^0_{red} = +0.80V$
$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$	with $E^0_{cell} = +0.46V$

Data Table 2. Experimental data for reaction of Zn, Cu, Pb and Ag metals in $Zn^{2+}(aq)$ ionic solution

Ionic solution used in the experiment	Metal strip used in the experiment			
	Zn(s)	Cu(s)	Pb(s)	Ag(s)
$Zn^{2+}(aq)$	**NR**	**NR**	**NR**	**NR**

Data Table 3. Experimental data for reaction of Zn, Cu, Pb and Ag metals in $Cu^{2+}(aq)$ ionic solution

Ionic solution used in the experiment	Metal strip used in the experiment			
	Zn(s)	Cu(s)	Pb(s)	Ag(s)
$Cu^{2+}(aq)$	**Black**	**NR**	**Dark grey**	**NR**

Data Table 4. Experimental data for reaction of Zn, Cu, Pb and Ag metals in $Ag^{+}(aq)$ ionic solution

Ionic solution used in the experiment	Metal strip used in the experiment			
	Zn(s)	Cu(s)	Pb(s)	Ag(s)
$Ag^{+}(aq)$	**Black**	**Dark grey**	**Black**	**NR**

Data Table 5. Experimental data for reaction of Zn, Cu, Pb and Ag metals in $Pb^{2+}(aq)$ ionic solution

Ionic solution used in the experiment	Metal strip used in the experiment			
	Zn(s)	Cu(s)	Pb(s)	Ag(s)
Pb^{2+} (aq)	**Black**	**NR**	**NR**	**NR**

D. For all the redox reactions in table 6 below write in the appropriate column: a) the net balanced ionic reaction (don't miss the charge for ions and the state [(s) or (aq)] for each substance; b) E^0_{rxn} value in V calculated with table 1 data; c) the sign of ΔG^0 (<, = , > 0); and d) the type of the reaction (Spontaneous, "Equilibrium, NR" or no reaction ("NR"). The first three reactions shown in table 6 as the example.

E. Compare the metals used in this experiment (Zn, Cu, Pb and Ag) and their ions ($Zn^{2+}(aq)$, $Cu^{2+}(aq)$, $Pb^{2+}(aq)$, $Ag^{+}(aq)$) by their relative reactivity as reducing (or oxidizing) agents and arrange them in the rows of oxidizing and reducing reactivity in Table 7 using table 1 and the observation data in Table 2-5. **Note:** a) **the highest oxidizing reactivity should be assigned to the metal ion, which ionic solution reacts with all other solid metals (oxidizing them to their metal ions); b) the highest reducing reactivity should be assigned to the metal which reduces all other metal ions to their neutral atoms**.

Data Table 6. Spontaneity of the oxidation/reduction reactions.

System	Net ionic reaction for the system in the well	E^0_{rxn}, Volts	Sign of ΔG^0 (<, =, or > 0)	Type of the reaction ("Spontaneous", "Equilibrium, NR", or "NR")
$Zn(s)/\ Cu^{2+}(aq)$	$Zn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Zn^{2+}(aq)$	1.1	< 0	Spontaneous
$Cu(s)/Zn^{2+}(aq)$	$Cu(s) + Zn^{2+}(aq) \rightarrow Cu^{2+}(aq) + Zn(s)$	– 1.1	> 0	NR
$Zn(s)/Zn^{2+}(aq)$	$Zn(s) + Zn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Zn(s)$	0	= 0	Equilibrium, NR
$Zn(s)/Ag^{+}(aq)$				
$Zn(s)/Pb^{2+}(aq)$				
$Cu(s)/Cu^{2+}(aq)$				
$Cu(s)/Ag^{+}(aq)$				
$Cu(s)/Pb^{2+}(aq)$				
$Pb(s)/Zn^{2+}(aq)$				
$Pb(s)/Cu^{2+}(aq)$				
$Pb(s)/Ag^{+}(aq)$				
$Pb(s)/Pb^{2+}(aq)$				
$Ag(s)/Zn^{2+}(aq)$				
$Ag(s)/Cu^{2+}(aq)$				
$Ag(s)/Ag^{+}(aq)$				
$Ag(s)/Pb^{2+}(aq)$				

Table 7. Relative reactivity of four metals and their ions. (Put the most reactive metal (or ion) at the left side of the row and the least reactive substance to the right side of the row).

Oxidizing relative reactivity row for Zn^{2+}, Cu^{2+}, Pb^{2+} and Ag^{+}	**(the most reactive ion) > > > (the least reactive ion)**
Reducing relative reactivity row for Zn, Cu, Pb and Ag	**(the most reactive metal) > > > (the least reactive metal)**