CH103 General Chemistry II

2018 Fall semester Quiz 2

Date: October 1. (Mon), ***Time: 19:00~19:45***

Professor Name	Class	Student I.D. Number	Name

1. A galvanic cell is constructed in which a $Br_2|Br^+$ half-cell is connected to a $Co^{2+}|Co$ half-cell. (total 8 pts)

$$Br_2(l) + 2\,e^- \rightarrow 2\,Br^-,\ E° = 1.065V$$

$$Co^{2+} + 2\,e^- \rightarrow Co(s),\ E° = -\,0.28V$$

a) Please write down balanced half equations, (make sure which is anode and cathode), and overall cell reaction. (6 pts)

$$\begin{aligned} &\text{anode:} && Co(s) \longrightarrow Co^{2+}(aq) + 2\,e^- \\ &\text{cathode:} && Br_2(l) + 2\,e^- \longrightarrow 2\,Br^-(aq) \\ &\text{overall:} && Co(s) + Br_2(l) \longrightarrow Co^{2+}(aq) + 2\,Br^-(aq) \end{aligned}$$

Each equation is 2 pts.

-1 pt for every unbalanced equations.

No points for wrong anode and cathode.

b) Calculate the cell potential, assuming that all reactants and products are in their standard states. (2 pts)

$$\mathcal{E}^\circ_{\text{cell}} = \mathcal{E}^\circ_{\text{cathode}} - \mathcal{E}^\circ_{\text{anode}} = 1.065 - (-0.28) = \boxed{1.34\ \text{V}}$$

No partial points.

2. The standard reduction potential of an MnO_4^-, Mn^{2+} half-cell is $E°_{cell} = 1.49V$. This half-cell is combined with a $Zn^{2+}|$ Zn half-cell ($E°_{cell} = -0.76V$) in a galvanic cell, given that $[Zn^{2+}] = [MnO_4^-] = [Mn^{2+}] = [H_3O^+] = 1M$. (total 15 pts)

a) Write down balanced half equations and the overall cell reaction. (5 pts)

$MnO_4^- + 8\,H_3O^+ + 5\,e^- \rightarrow Mn^{2+} + 12\,H_2O$ (2 pts)

$Zn \rightarrow Zn^{2+} + 2\,e^-$ (1 pt)

2 MnO^-_4 + 16 H_3O^+ + 5 Zn → 2 Mn^{2+} + 24 H_2O + 5Zn^{2+} (2 pts)

b) Calculate the cell potential, assuming that all reactants and products are in their standard states. (2 pts)

E° = 1.49 – (-0.76) = 2.25V

No partial points

c) Suppose that the same reaction as above is operated at pH 2.00 with [MnO_4^-] = 0.35M, [Mn^{2+}] = 0.0020 M and [Zn^{2+}] = 0.015M. Calculate the cell potential E_{cell} at 25°C. (5 pts)

From the equation $E_{cell} = E°_{cell} – (0.0592V /n) \log Q$

$E(MnO_4^-, Mn^{2+}) = 1.49\ V – (0.0592\ V / 10) \log (0.002)^2/[(0.35)^2(10^{-2})^{16}] = 1.32V$

+2 pts for right answer

If a student got n value right (shown from the work), but wrong answer, give 1 pt

$E(Zn|Zn^{2+}) = - 0.76\ V - (0.0592\ V / 2) \log (1)/(0.015) = -0.814\ V$

+2 pts for right answer

If a student got n value right (shown from the work), but wrong answer, give 1 pt

E_{cell} = 1.33 – (- 0.814) = 2.144V

+1 pt for Final answer right

d) Calculate the equilibrium constant for the redox reaction at 25°C using the cell potential calculated in part (b), and explain what the K value (answer) implies. (3 pts)

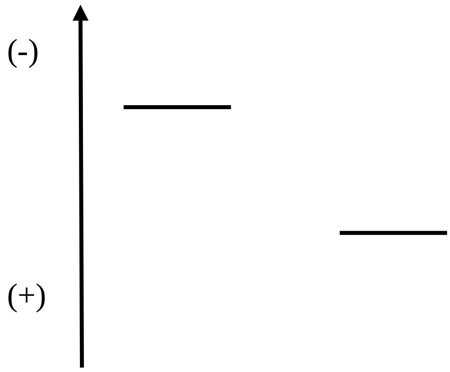
Log K = (n/0.0592V) $E°_{cell}$ = (10/0.0592) (2.25V) = 380

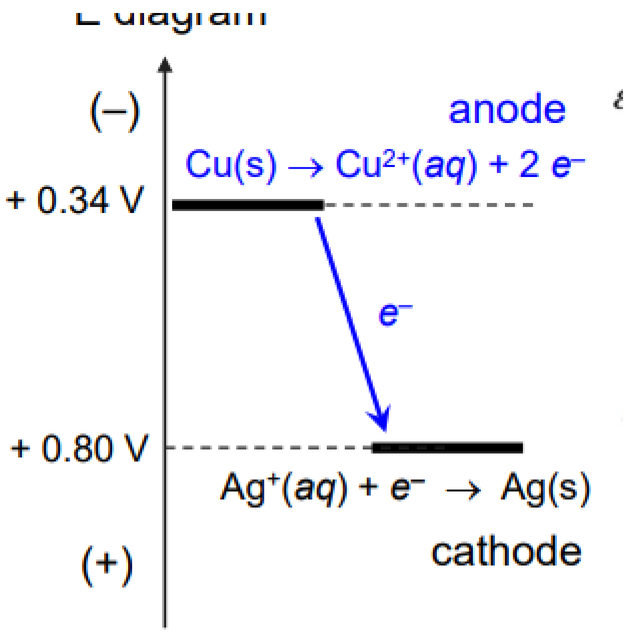
K = 10^{380} (2 pt)

Large equilibrium constant implies that the strength of permanganate ion as an oxidizing agent and that of zinc is as reducing agent. (1 pt)

3. Complete the energy level diagram for the Cu|Cu^{2+} || Ag^+|Ag galvanic cell. (total 7 pts)

a) On the diagram, 1) indicate which is the anode and cathode. 2) write down the half cell equations next to anode and cathode position. 3) draw in which direction electron flows. (5 pts)





2 pts for anode and cathode

2 pts for right half cell equations

1 pt for direction of electron flow.

b) This reaction is spontaneous. If this spontaneous reaction ends, explain what will happen to the energy levels of the anode and cathode. (2 pts)

Electrons will continue to flow from Cu to Ag+ until the energies of the electrons in both half-reactions have become equal to one another and there is no longer a driving force for the reaction to occur

4. Corrosion takes place in real life situations, for example, iron gets rusted, which is the sign of corrosion.

   a) **Sacrifical anode** is a way of preventing iron corrosion. Explain what it is. (3 pts)

By using another metal that is harder to reduce than iron, that sacrificial metal will be oxidized in preference to iron. Then iron can be protected as other metal gets oxidized instead.

   b) Choose two possible sacrificial anode, referring to the standard reduction potentials table give below. Explain Why. (3 pts)

Table 11.1 Standard Reduction Potentials at 25°C (298 K) for Many Common Half-reactions

Half-reaction	$\mathcal{E}°$ (V)	Half-reaction	$\mathcal{E}°$ (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$IO_4^- + 2H^+ + 2e^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	−0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	−1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	−2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	−2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	−2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	−2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	−3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

If a student chose any two from Cr, Zn, Mn, Al, Mg, La, Na, Ca, Ba, K and Li, give 1 pt each.

From comparing Standard reduction potential, one that gets more readily oxidized than $Fe^{2+} + 2e^- \rightarrow Fe$ (E^0 <-0.41V) can be used as a sacrificial anode. (this kind of explanation will get 1pt)

5. The following reduction potentials are measured at pH 0 in aqueous solution: (total 4 pts)

$IO_3^- + 12\ H^+ + 10\ e^- \rightarrow I_2\ (l) + 6\ H_2O$ $\quad E^0 = 1.195$ V

$I_2\ (l) + 2\ e^- \rightarrow 2\ I^-$ $\quad E^0 = 0.5355$ V

a) Write chemical equation of iodine disproportionation. (2 pts)

$6I_2 + 6H_2O \rightarrow 10\ I^- + 2IO_3^- + 12\ H^+$

b) Calculate the standard potential difference. Does iodine disproportionate spontaneously in acidic solution? (2 pts)

E = 0.5355 V – 1.195 V = -0.6595 V : non spontaneous reaction.