AP WORKSHEET 09GHIJ: Electrochemistry Problems

NOTE: To help solve these problems, use the table of Standard Electrode Potentials found on page 5 of this worksheet. This table used to be provided with the AP exam but is no longer given, rather now the relevant SERP's will be given within the body of any question that requires them.

Questions 2 and 4 require the use of the Nernst equation that is not examined *quantitatively* on the AP exam. However, it is given on the equations & constants sheet, and may still be useful to work the problems to help you understand the electrochemistry concepts, but you will NOT be asked to perform Nernst equation *calculations* on the AP exam.

$$E_{cell} = E° - \left(\frac{RT}{nF}\right) \ln Q \qquad \text{or} \qquad E_{cell} = E° - \left(\frac{0.0592}{n}\right) \log Q$$

Where, R is the gas constant (8.314 J/K mol), T = Kelvin temp, n = # of electrons transferred, F = Faraday constant, E° = voltage generated **IF** the conditions **WERE** standard, ln = natural log and Q = the reaction quotient (remember to raise to stoichiometric numbers and omit pure solids & liquids).

1. When silver metal reacts with bromine at 298 K, a spontaneous REDOX reaction takes place.

   (a) Calculate the following for this reaction;

      (i) The E°. (2)

      (ii) The equilibrium constant. (2)

(b) Considering both the information in the question, AND the values that you have calculated in (a) above, WITHOUT doing any further calculations, comment on the likely value for ∆G°. (2)

2. A silver electrode is placed in a saturated solution of AgBr. This half-cell is connected to a standard hydrogen electrode and the voltage is found to be + 0.437 V. Calculate Ksp for AgBr. (6)

3. Consider the following disproportionation reaction at 298 K.

$$3Fe^{2+}_{(aq)} \rightarrow 2Fe^{3+}_{(aq)} + Fe_{(s)}$$

(a) Calculate the standard potential, E°. (2)

(b) What does the value that you have calculated in (a) suggest about the thermodynamic favorability of the reaction? Explain you answer in terms of $\Delta G°$. (2)

4. Consider the following cell that is operating at 298 K.

$$Al(s) \mid Al^{3+} (0.03\ M) \parallel Fe^{2+} (0.3\ M) \mid Fe(s)$$

(a) Calculate the voltage that is generated when the cell first starts to work. (2)

(b) After a period of time has elapsed, the $[Al^{3+}]$ has increased from 0.03 M to 0.06 M. Calculate the cell potential at this point in time. (4)

STANDARD REDUCTION POTENTIALS IN AQUEOUS SOLUTION AT 25°C

Half-reaction			E^0(V)
$F_2\ (g) + 2e^-$	→	$2F^-$	+ 2.87
$Co^{3+} + e^-$	→	Co^{2+}	+ 1.82
$Au^{3+} + 3e^-$	→	$Au(s)$	+ 1.50
$Cl_2\ (g) + 2e^-$	→	$2Cl^-$	+ 1.36
$O_2\ (g) + 4H^+ + 4e^-$	→	$2H_2O(l)$	+ 1.23
$Br_2\ (l) + 2e^-$	→	$2Br^-$	+ 1.07
$2Hg^{2+} + 2e^-$	→	Hg_2^{2+}	+ 0.92
$Hg^{2+} + 2e^-$	→	$Hg(l)$	+ 0.85
Ag+ $+ e^-$	→	$Ag(s)$	+ 0.80
$Hg_2^{2+} + 2e^-$	→	$2Hg(l)$	+ 0.79
$Fe^{3+} + e^-$	→	Fe^{2+}	+ 0.77
$I_2\ (s) + 2e^-$	→	$2I^-$	+ 0.53
$Cu^+ + e^-$	→	$Cu(s)$	+ 0.52
$Cu^{2+} + 2e^-$	→	$Cu(s)$	+ 0.34
$Cu^{2+} + e^-$	→	Cu^+	+ 0.15
$Sn^{4+} + 2e^-$	→	Sn^{2+}	+ 0.15
$S\ (s) + 2H^+ + 2e^-$	→	$H_2S(g)$	+ 0.14
$2H^+ + 2e^-$	**→**	**$H_2(g)$**	**0.00**
$Pb^{2+} + 2e^-$	→	$Pb(s)$	- 0.13
$Sn^{2+} + 2e^-$	→	$Sn(s)$	- 0.14
$Ni^{2+} + 2e^-$	→	$Ni(s)$	- 0.25
$Co^{2+} + 2e^-$	→	$Co(s)$	- 0.28
$Tl^+ + e^-$	→	$Tl(s)$	- 0.34
$Cd^{2+} + 2e^-$	→	$Cd(s)$	- 0.40
$Cr^{3+} + e^-$	→	Cr^{2+}	- 0.41
$Fe^{2+} + 2e^-$	→	$Fe(s)$	- 0.44
$Cr^{3+} + 3e^-$	→	$Cr(s)$	- 0.74
$Zn^{2+} + 2e^-$	→	$Zn(s)$	- 0.76
$Mn^{2+} + 2e^-$	→	$Mn(s)$	- 1.18
$Al^{3+} + 3e^-$	→	$Al(s)$	- 1.66
$Be^{2+} + 2e^-$	→	$Be(s)$	- 1.70
$Mg^{2+} + 2e^-$	→	$Mg(s)$	- 2.37
$Na^+ + e^-$	→	$Na(s)$	- 2.71
$Ca^{2+} + 2e^-$	→	$Ca(s)$	- 2.87
$Sr^{2+} + 2e^-$	→	$Sr(s)$	- 2.89
$Ba^{2+} + 2e^-$	→	$Ba(s)$	- 2.90
$Rb^+ + e^-$	→	$Rb(s)$	- 2.92
$K^+ + e^-$	→	$K(s)$	- 2.92
$Cs^+ + e^-$	→	$Cs(s)$	- 2.92
$Li^+ + e^-$	→	$Li(s)$	- 3.05