Culver City H.S. • AP Chemistry

Name _____
Period ___ Date ___/___/___

18 • Electrochemistry

PRACTICE TEST

anode (ox.) || cathode (red.)

1. Which of the following is the correct cell notation for the reaction
$Hg_2^{2+} + Cd(s) \rightarrow Cd^{2+} + 2Hg(l)$
a) $Cd^{2+} | Cd || Hg_2^{2+} | Hg$
b) $Cd^{2+} | Hg_2^{2+} || Cd | Hg$
c) $Cd | Cd^{2+} || Hg_2^{2+} | Hg$
d) $Cd^{2+} | Hg || Hg_2^{2+} | Cd$
e) $Hg | Cd || Hg_2^{2+} | Cd^{2+}$

2. Consider an electrochemical cell where the following reaction takes place:
$3Sn^{2+}(aq) + 2Al(s) \rightarrow 3Sn(s) + 2Al^{3+}(aq)$
Which of the following is the correct cell notation for this cell?
a) $Al | Al^{3+} || Sn^{2+} | Sn$
b) $Al^{3+} | Al || Sn | Sn^{2+}$
c) $Sn | Sn^{2+} || Al^{3+} | Al$
d) $Sn | Al^{3+} || Al | Sn^{2+}$
e) $Al | Sn^{2+} || Sn | Al^{3+}$

$2Al \rightarrow 2Al^{3+}$ oxd
$3Sn^{2+} \rightarrow 3Sn$ red.

3. An early method of producing aluminum metal was the reaction of aluminum salts with sodium metal: $Al^{3+} \rightarrow Al$
$Al^{3+} + 3Na(s) \rightleftharpoons Al(s) + 3Na^+ \quad E° = +1.05\ V$
What is $\Delta G°$ for this reaction
a) -304 kJ
b) -101 kJ
c) +101 kJ
d) +202 kJ
e) +304 kJ

$\Delta G° = -nFE°$
$= -(3\,e^-)(96{,}485\ \frac{coul}{mole})(1.05\ V)$
$= 303928\ coul\ V$
$= 303928\ J$
$= 304\ kJ$

4. Calculate ΔG for the following reaction:
cathode, anode
$I_2(s) + 2Br^-(aq) \rightarrow 2I^-(aq) + Br_2(l)$
a) +105 kJ
b) -105 kJ
c) +312 kJ
d) +52 kJ
e) -312 kJ

$E°_{cell} = .535\ V - (1.08) = -.545\ V$
$\Delta G = -nFE = -(2)(96485)(-.545) =$

5. If ΔG of the following reaction is -203 kJ, what is $E°$? $2Ag^+(aq) + Ni(s) \rightarrow 2Ag(s) + Ni^{2+}(aq)$
a) -1.05 V
b) +2.10 V
c) +0.0011 V
d) -0.011 V
e) +1.05 V

+105169 J

$\Delta G° = -nFE°$
$E° = \frac{-\Delta G°}{nF} = \frac{-(-203000\ J)}{(2\ mole^-)(96485\ \frac{coul}{mole^-})}$
$= +1.05\ V$

6. Given the two half reactions and their potentials, which net reaction is spontaneous?
$Mg^{2+}(aq) + 2e^- \rightarrow Mg(s) \quad E° = -2.37\ V$
$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s) \quad E° = -0.25\ V$ ← cathode
a) $Ni(s) + Mg^{2+}(aq) \rightarrow Mg(s) + Ni^{2+}(aq)$
b) $Ni^{2+}(aq) + Mg(s) \rightarrow Mg^{2+}(aq) + Ni(s)$
c) $Ni(s) + Mg(s) \rightarrow Mg^{2+}(aq) + Ni^{2+}(aq)$
d) $Mg^{2+}(aq) + Ni^{2+}(aq) \rightarrow Mg(s) + Ni(s)$
e) $Mg^{2+}(aq) + Mg(s) \rightarrow Ni(s) + Ni^{2+}(aq)$

7. Calculate $E°$ for the following reaction:
cathode
$Sn^{4+}(aq) + 2K(s) \rightarrow Sn^{2+}(aq) + 2K^+(aq)$
a) +6.00 V
b) -3.08 V
c) +3.08 V
d) +2.78 V
e) -2.78 V

$Sn^{4+} + 2e^- \rightarrow Sn^{2+} \quad E° = .15\ V$ ← cathode
$K^+ + e^- \rightarrow K \quad E° = -2.93\ V$

8. Calculate $E°$ for the following reaction:
cathode
$2Al^{3+}(aq) + 3Cd(s) \rightarrow 2Al(s) + 3Cd^{2+}(aq)$
a) -2.06 V
b) +4.52 V
c) +2.06 V
d) -4.52 V
e) -1.26 V (electrolytic)

$Al^{3+} + 3e^- \rightarrow Al \quad E° = -1.66\ V$
$Cd^{2+} + 2e^- \rightarrow Cd \quad E° = -0.40\ V$

9. Using data from the reduction potential table and the reaction
anode, cathode
$2Ag(s) + Pt^{2+}(aq) \rightarrow Pt(s) + 2Ag^+(aq) \quad E° = 0.38\ V$
calculate the standard reduction potential of the half-reaction
$Pt^{2+}(aq) + 2e^- \rightarrow Pt(s)$
a) -1.18 V
b) -0.40 V
c) 0.40 V
d) 1.18 V
e) 2.00 V

$E°_{cell} = E°_{cathode} - E°_{anode}$
$0.38\ V = x - (.80)$
$x = 1.18\ V$

10. An electrochemical cell of notation $Pd | Pd^{2+} || Cu^{2+} | Cu$ has an $E° = -0.65$ V. If we know that the standard reduction potential of Cu^{2+}/Cu is $E° = 0.34$ V, what is the standard reduction potential for Pd^{2+}/Pd?
a) -0.99 V
b) -0.31 V
c) +0.31 V
d) 0.62 V
e) +0.99 V

$E°_{cell} = E°(cathode) - E°(anode)$
$-0.65 = 0.34 - x$
$x = .99\ V$

11. What is the equilibrium constant for the following reaction at 298 K?

$2Ag^+(aq) + 2I^-(aq) \rightarrow I_2(s) + 2Ag(s)$ $E° = +0.265$ V

a) 2.99×10^4 d) 87.9
b) 9.04×10^8 e) 1.60×10^7
c) 7.73×10^3

Use $\Delta G = -nFE$ Then $\Delta G = -RT\ln K$] solve for K

12. What is the equilibrium constant for the following reaction at 20C? 293K

$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$ $E° = +0.78$ V

a) 2.3×10^{26} d) 1.8×10^{28}
b) 6.9×10^{26} e) 1.2×10^{-21}
c) 1.4×10^{27}

13. What is the cell potential for

$3Sn^{4+}(aq) + 2Al(s) \rightarrow 3Sn^{2+}(aq) + 2Al^{3+}(aq)$

$E° = 1.81$ V when $[Sn^{4+}] = 1.0$, $[Sn^{2+}] = 1.0 \times 10^{-2}$, and $[Al^{3+}] = 1.5 \times 10^{-3}$ at 298 K.

$6e^- + 3Sn^{4+} \rightarrow 3Sn^{2+}$ $2Al \rightarrow 2Al^{3+} + 6e^-$

$Q = \frac{[Sn^{2+}]^3[Al^{3+}]^2}{[Sn^{4+}]^3} = 2.25 \times 10^{-12}$

a) 1.70 V d) 1.86 V
b) 1.76 V e) 1.93 V
c) 1.81 V

Nernst → $E_{cell} = 1.81V - \frac{0.0592}{6}\log(2.25 \times 10)^{-12} = 1.92V$

14. If the potential cell is +1.32 V at Q = 0.0969 with n = 2, what is the standard potential of the cell? Nernst Eq.

a) +1.35 V d) +1.34 V
b) +1.48 V e) +1.29 V
c) +1.31 V

$1.32V = E° - \frac{0.0592}{2}\log(.0969)$

$E° = 1.29V$

15. Predict the product at the anode when electric current is passed through a solution of KI.

a) $I_2(s)$ d) K(s)
b) $K^+(aq)$ e) $O_2(g)$
c) $H_2(g)$

e⁻ are lost at (+) electrode

$E° = -1.23V$ $2H_2O \rightarrow O_2 + 4H^+ + 4e^-$

$K^+ + e^- \rightarrow K$ $E° = -2.93V$

$I_2 + 2e^- \rightarrow 2I^-$ $E° = .535V$ ← higher $E°_{ox}$

16. If electric current is passed through aqueous LiBr, the product at the cathode would be H_2 and the product at the anode would be Br_2.

a) $H_2O(l)$, $Li^+(aq)$ d) $Br_2(l)$, $H_2(g)$
b) $Br_2(l)$, Li(s) e) $H_2(g)$, $Br_2(l)$
c) Li(s), $Br_2(l)$

$Br_2 + 2e^- \rightarrow 2Br^-$ $E° = 1.08$ V ← higher $E°_{ox}$

$Li^+ + e^- \rightarrow Li$ $E° = -3.045$ V

$2H_2O \rightarrow O_2 + 4H^+ + 4e^-$ $E° = -1.23$ V ← higher $E°_{red}$

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ $E° = -.828$

17. How long would it take to deposit 1.36 g of copper from an aqueous solution of copper(II) sulfate by passing a current of two amperes through the solution? $Cu^{2+} + 2e^- \rightarrow Cu$

a) 2070 sec d) 736 sec
b) 1.11×10^{-5} sec e) 1030 sec
c) 2570 sec

$1.36g\,Cu \times \frac{1mol}{63.55g} \times \frac{2mole^-}{1mol} \times \frac{96485Coul}{mole^-} \times \frac{1s}{2Coul} =$

18. If a current of 6.0 amps is passed through a solution of Ag^+ for 1.5 hours, how many grams of silver are produced? $Ag^+ + e^- \rightarrow Ag$

a) 0.60 g d) 3.0 g
b) 36 g e) 1.0 g
c) 0.34 g

$1.5hr \times \frac{3600s}{1hr} \times \frac{6.0Coul}{s} \times \frac{1mole^-}{96485Coul} \times \frac{1mol\,Ag^+}{1mole^-} \times \frac{107.87g}{1mol}$

19. How many kilowatt hours of electrical energy are required to plate 2.00 grams of silver from an aqueous solution of silver nitrate on to a necklace using 3.00V? (1 joule = 1 volt-coulomb and 1 kwh = 3.60×10^6 J)

a) 0.00135 kwh d) 0.00149 kwh
b) 0.000165 kwh e) 2.07 kwh
c) 32.4 kwh

$(1789Coul)(3.00V) = 5367J$ $\frac{5367J}{3.60 \times 10^6 J/kwh} =$

$2.00g\,Ag \times \frac{1mol}{107.87g} \times \frac{1mole^-}{1mol} \times \frac{96485Coul}{1mole^-} = 1789Coul$

20. How is aluminum currently produced in industry?

a) by reduction of Al^{3+} in Al_2O_3 with Na(s)
b) electrochemical reduction of pure Al_2O_3 to give Al and O_2
c) electrolysis of AlF_3 to give Al and F_2
d) electrolysis of a mixture of Al_2O_3 and Na_3AlF_6 to give Al and O_2 Hall-Heroult process
e) by reduction of Al^{3+} in Al_2O_3 with CO(g)

21. How was aluminum originally made?

a) the Hall-Heroult process
b) Al_2O_3 mixed with cryolite is electrolyzed
c) electrolysis of molten Al_2O_3
d) mining and purifying directly
e) reducing $AlCl_3$ with sodium ← very expensive process

22. Using data from the reduction potential table, predict which of the following is the best oxidizing agent. most likely Reduced (top of chart)

a) F_2 d) Ag^+
b) Ag e) Al^{3+}
c) Sn^{4+}

23. Under acidic conditions the bromate ion is reduced to the bromide ion. Write the balanced half-reaction for this process. $BrO_3^- \Rightarrow Br^-$

a) $BrO_3^- + 6H^+ + 6e \rightarrow Br^- + 3H_2O$
b) $2BrO_3^- + 6H^+ \rightarrow Br_2^- + 6H_2O + 3e$
c) $Bro_3^- + 6H_2O + 10e \rightarrow Br_2^- + 12H^+ + 3\ O_2$
d) $2BrO_3^- + 6H_2O \rightarrow 2Br^- + 12H^+ + 6\ O_2 + 8e$
e) $2BrO_3^- + 6H^+ \rightarrow Br_2^- + 3H_2O + 3e$

24. Balance the following redox equation which occurs in acidic solution.

$N_2H_4(g) + BrO_3^-(aq) \rightarrow Br^-(aq) + N_2(g)$

a) $3N_2H_4 + BrO_3^- \rightarrow 3N_2 + Br^- + 3H_2O + 6H^+$
b) $N_2H_4 + BrO_3^- + 2H^+ \rightarrow 2Br^- + N_2 + 3H_2O$
c) $3N_2H_4 + 2BrO_3^- + 12H^+ \rightarrow 3N_2 + 2Br^- + 6H_2O + 12H^+$
d) $N_2H_4 + 2BrO_3^- + 8H^+ \rightarrow 2Br^- + N_2 + 6H_2O$
e) $3N_2H_4 + 2BrO_3^- \rightarrow 3N_2 + 2Br^- + 6H_2O$

25. Which of the following reactions is NOT a redox reaction?

a) $2HgO(s) \rightarrow 2\ Hg(l) + O_2(g)$
b) $H_2(g) + Br_2(g) \rightarrow 2HBr(g)$
c) $2HCl(aq) + Zn(s) \rightarrow H_2(g) + ZnCl_2(aq)$
d) $\overset{+1}{H_2}\overset{+4}{C}\overset{-2}{O_3}(aq) \rightarrow \overset{+1}{H_2}\overset{-2}{O}(l) + \overset{+4}{C}\overset{-2}{O_2}(g)$ no change in oxidation #'s
e) $2KClO_3 \rightarrow 2KCl(s) + 3\ O_2(g)$

Standard Reduction Potentials at 25°C E° (volts)

$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87
$Au^{3+} + 3e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.36
$O_2(g) + 4H_3O^+(aq) + 4e^- \rightarrow 6H_2O(l)$	+1.23
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.08
$Ag^+(aq) + e^- \rightarrow Ag(s)$	+0.80
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0.535
$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$	+0.337
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	+0.15
$Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$	-0.14
$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$	-0.40
$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$	-0.763
$2H_2O(l) + 2e^- \rightarrow H_2(g) + 2OH^-(aq)$	-0.828
$Al^{3+}(aq) + 3e^- \rightarrow Al(s)$	-1.66
$K^+(aq) + e^- \rightarrow K(s)$	-2.93
$Li^+(aq) + e^- \rightarrow Li(s)$	-3.045

Answers:

1.	C	11.	B	21.	E
2.	A	12.	B	22.	A
3.	A	13.	E	23.	A
4.	A	14.	E	24.	E
5.	E	15.	A	25.	D
6.	B	16.	E		
7.	C	17.	A		
8.	E	18.	B		
9.	D	19.	D		
10.	E	20.	D		

#11) $\Delta G = -nFE = (-2\ \frac{mole^-}{})(96485\ \frac{Coul}{mole^-})(.265\ V)$

$= -51137\ Coul \cdot V = -51137\ J$

$\Delta G = -RT\ln K = -51137\ J$

$\ln K = \frac{-51137\ J}{-(8.314\ J/K\,mol)(298K)} = 20.64$

$K = e^{20.64} = 9.2 \times 10^8$ (B)

#12) $-nFE = -RT\ln K$ ← ΔG

$\ln K = \frac{nFE}{RT} = \frac{2(96485)(.78)}{(8.314)(293K)} = 61.788$

$K = e^{61.788} = 6.83 \times 10^{26}$ (B)

#24) $3 \times (N_2H_4 \rightarrow N_2 + 4H^+ + 4e^-)$

$2 \times (6e^- + 6H^+ + BrO_3^- \rightarrow Br^- + 3H_2O)$

$3N_2H_4 + 12H^+ + 2BrO_3^- \rightarrow 3N_2 + 12H^+ + 2Br^- + 6H_2O$