A.P. Chemistry Practice Test – Ch. 17: Electochemistry

MULTIPLE CHOICE. Choose the one alternative that best completes the statement or answers the question.

1) The gain of electrons by an element is called _____.
   A) oxidation
   B) sublimation
   C) reduction
   D) disproportionation
   E) fractionation

2) _____ is reduced in the following reaction:

$$Cr_2O_7^{2-} + 6S_2O_3^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 3S_4O_6^{2-} + 7H_2O$$

   A) $Cr_2O_7^{2-}$ B) $S_2O_3^{2-}$ C) H^+ D) $S_4O_6^{2-}$ E) Cr^{3+}

3) _____ is the oxidizing agent in the reaction below.

$$Cr_2O_7^{2-} + 6S_2O_3^{2-} + 14H^+ \rightarrow 2Cr^{3+} + 3S_4O_6^{2-} + 7H_2O$$

   A) Cr^{3+} B) H^+ C) $S_2O_3^{2-}$ D) $Cr_2O_7^{2-}$ E) $S_4O_6^{2-}$

4) Which of the following reactions is a redox reaction?

   (a) $K_2CrO_4 + BaCl_2 \rightarrow BaCrO_4 + 2KCl$
   (b) $Pb_2^{2+} + 2Br^- \rightarrow PbBr$
   (c) $Cu + S \rightarrow CuS$

   A) (a) only B) (b) only C) (c) only D) (a) and (c) E) (b) and (c)

5) Which substance is the reducing agent in the reaction below?

$$Pb + PbO_2 + 2H_2SO_4 \rightarrow 2PbSO_4 + 2H_2O$$

   A) H_2O B) Pb C) PbO_2 D) $PbSO_4$ E) H_2SO_4

6) What is the oxidation number of chromium in the dichromate ion?
   A) +7 B) +3 C) +14 D) +12 E) +6

7) What is the oxidation number of potassium in potassium permanganate?
   A) 0 B) +3 C) +2 D) +1 E) –1

8) What is the oxidation number of manganese in the permanganate ion, MnO_4^-?
   A) +2 B) +5 C) +1 D) +7 E) +4

9) What is the coefficient of the bromide ion when the following redox equation is balanced?

$BrO^- + Fe(OH)_2 \rightarrow Br^- + Fe(OH)_3$ (basic solution)

A) 4 B) 3 C) 2 D) 1 E) 5

10) What is the coefficient of Fe^{3+} when the following equation is balanced?

$CN^- + Fe^{3+} \rightarrow CNO^- + Fe^{2+}$ (basic solution)

A) 1 B) 2 C) 3 D) 4 E) 5

11) The balanced half-reaction in which sulfate ion is reduced to sulfite ion is a _____ process.

A) three-electron
B) four-electron
C) two-electron
D) one-electron
E) six-electron

12) The half-reaction occurring at the anode in the balanced reaction shown below is _____.

$3MnO_4^-(aq) + 24H^+(aq) + 5Fe(s) \rightarrow 3Mn^{2+}(aq) + 5Fe^{3+}(aq) + 12H_2O(l)$

A) $Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$
B) $MnO_4^-(aq) + 8H^+(aq) + 5e^- \rightarrow Mn^{2+}(aq) + 4H_2O(l)$
C) $Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$
D) $2MnO_4^-(aq) + 12H^+(aq) + 6e^- \rightarrow 2Mn^{2+}(aq) + 3H_2O(l)$
E) $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$

13) The purpose of the salt bridge in an electrochemical cell is to _____.

A) provide a source of ions to react at the anode and cathode.
B) provide oxygen to facilitate oxidation at the anode.
C) provide a means for electrons to travel from the anode to the cathode.
D) provide a means for electrons to travel from the cathode to the anode.
E) maintain electrical neutrality in the half-cells via migration of ions.

14) In a voltaic cell, electrons flow from the _____ to the _____.

A) salt bride, anode
B) anode, cathode
C) anode, salt bridge
D) salt bridge, cathode
E) cathode, anode

15) 1V = _____.

A) 96485 C B) 1 J/C C) 1 C/J D) 1 J/s E) 1 amp • s

Table 20.1

Half Reaction	E°(V)
$F_2(g) + 2e^- \rightarrow 2F^-(aq)$	+2.87
$Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$	+1.359
$Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$	+1.065
$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	+1.23
$Ag^+ + e^- \rightarrow Ag(s)$	+0.799
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$	+0.771
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	+0/536
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	+0.34
$2H^+ + 2e^- \rightarrow H_2(g)$	0
$Pb^{2+} + 2e^- \rightarrow Pb(s)$	–0.126
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	–0.28
$Li^+ + e^- \rightarrow Li(s)$	–3.05

16) Which of the halogens in Table 20.1 is the strongest oxidizing agent?

A) Br_2

B) I_2

C) Cl_2

D) F_2

E) All of the halogens have equal strength as oxidizing agents.

17) Which one of the following types of elements is most likely to be a good oxidizing agent?

A) transition elements

B) alkaline earth elements

C) lanthanides

D) alkali metals

E) halogens

Table 20.1

Half Reaction	E°(V)
F_2 (g) + 2e⁻ → 2F⁻ (aq)	+2.87
Cl_2 (g) + 2e⁻ → 2Cl⁻ (aq)	+1.359
Br_2 (l) + 2e⁻ → 2Br⁻ (aq)	+1.065
O_2 (g) + 4H^+ (aq) + 4e⁻ → 2H_2O (l)	+1.23
Ag^+ + e⁻ → Ag (s)	+0.799
Fe^{3+} (aq) + e⁻ → Fe^{2+} (aq)	+0.771
I_2 (s) + 2e⁻ → 2I⁻ (aq)	+0/536
Cu^{2+} + 2e⁻ → Cu (s)	+0.34
2H^+ + 2e⁻ → H_2 (g)	0
Pb^{2+} + 2e⁻ → Pb (s)	–0.126
Ni^{2+} + 2e⁻ → Ni (s)	–0.28
Li^+ + e⁻ → Li (s)	–3.05

18) Using Table 20.1, which substance can be oxidized by O_2 (g) in acidic aqueous solution?

A) Ni^{2+} (aq) B) Br_2 (l) C) Ag (s) D) Cu^{2+} (aq) E) Br^- (aq)

Table 20.2

Half–reaction	E° (V)
Cr^{3+} (aq) + 3e⁻ → Cr (s)	–0.74
Fe^{2+} (aq) + 2e⁻ → Fe (s)	–0.440
Fe^{3+} (aq) + e⁻ → Fe^{2+} (s)	+0.771
Sn^{4+} (aq) + 2e⁻ → Sn^{2+} (aq)	+0.154

19) The standard cell potential ($E°_{cell}$) for the voltaic cell based on the reaction below is _____ V.

$$Sn^{2+}\,(aq) + 2Fe^{3+}\,(aq) \rightarrow 2Fe^{2+}\,(aq) + Sn^{4+}\,(aq)$$

A) +1.39 B) +0.46 C) +0.617 D) +1.21 E) –0.46

20) The standard cell potential ($E°_{cell}$) for the voltaic cell based on the reaction below is _____ V.

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

A) +0.89 B) +1.94 C) –0.59 D) +2.53 E) –1.02

21) The relationship between the change in Gibbs free energy and the emf of an electrochemical cell is given by _____.

A) $\Delta G = \frac{-E}{nF}$ B) $\Delta G = -nRTF$ C) $\Delta G = \frac{-nF}{E}$ D) $\Delta G = \frac{-nF}{ERT}$ E) $\Delta G = -nFE$

Table 20.2

Half-reaction	E° (V)
$Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$	–0.74
$Fe^{2+}(aq) + 2e^- \rightarrow Fe(s)$	–0.440
$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(s)$	+0.771
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	+0.154

22) Which of the following reactions will occur spontaneously as written?

A) $Sn^{4+}(aq) + Fe^{2+}(aq) \rightarrow Sn^{2+}(aq) + Fe(s)$

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

C) $3Fe(s) + 2Cr^{3+}(aq) \rightarrow 2Cr(s) + 3Fe^{2+}(aq)$

D) $Sn^{4+}(aq) + Fe^{3+}(aq) \rightarrow Sn^{2+}(aq) + Fe^{2+}(aq)$

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

23) The standard cell potential (E°_{cell}) for the reaction below is +0.63 V. The cell potential for this reaction is _____ V when $[Zn^{2+}] = 1.0$ M and $[Pb^{2+}] = 2.0 \times 10^{-4}$ M.

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

A) 0.74 B) 0.52 C) 0.63 D) 0.85 E) 0.41

24) Consider an electrochemical cell based on the reaction:

$$2H^+(aq) + Sn(s) \rightarrow Sn^{2+}(aq) + H_2(g)$$

Which of the following actions would change the measured cell potential?

A) lowering the pH in the cathode compartment

B) increasing the pressure of hydrogen gas in the cathode compartment

C) increasing the $[Sn^{2+}]$ in the anode compartment

D) increasing the pH in the cathode compartment

E) Any of the above will change the measure cell potential.

25) One of the differences between a voltaic cell and an electrolytic cell is that in an electrolytic cell _____.

A) electrons flow toward the anode

B) a nonspontaneous reaction is forced to occur

C) an electric current is produced by a chemical reaction

D) O_2 gas is produced at the cathode

E) oxidation occurs at the cathode

26) How many minutes will it take to plate out 2.19 g of chromium metal from a solution of Cr^{3+} using a current of 35.2 amps in an electrolyte cell?

A) 1.92 B) 5.77 C) 346 D) 115 E) 17.3

27) How many grams of Ca metal are produced by the electrolysis of molten $CaBr_2$ using a current of 30.0 amp for 10.0 hours?

A) 112 B) 448 C) 0.0622 D) 224 E) 22.4

28) How many grams of copper will be plated out by a current of 2.3 A applied for 25 minutes to a 0.50–M solution of copper(II) sulfate?

A) 0.019 B) 1.1 C) 0.036 D) 2.2 E) 1.8×10^{-2}

29) A voltaic cell is constructed with two silver–silver chloride electrodes, where the half–reaction is

$$AgCl\ (s) + e^- \rightarrow Ag\ (s) + Cl^-\ (aq) \qquad E^\circ = +0.222\ V$$

The concentrations of chloride ion in the two compartments are 0.0222 M and 2.22 M, respectively. The cell emf is _____ V.

A) 0.00222 B) 0.232 C) 0.118 D) 0.212 E) 22.2

30) How many seconds are required to produce 4.00 g of aluminum metal from the electrolysis of molten $AlCl_3$ with an electrical current of 12.0 A?

A) 2.90×10^5 B) 3.57×10^3 C) 27.0 D) 9.00 E) 1.19×10^3

Advanced Placement Chemistry: 1996 Free Response Questions

7) $Sr(s) + Mg^{2+} <===> Sr^{2+} + Mg(s)$

Consider the reaction represented above that occurs at 25°C. All reactants and products are in their standard states. The value of the equilibrium constant, K_{eq}, for the reaction is 4.2×10^{17} at 25°C.

(a) Predict the sign of the standard cell potential, E°, for a cell based on the reaction. Explain your prediction.

(b) Identify the oxidizing agent for the spontaneous reaction.

(c) If the reaction were carried out at 60°C instead of 25°C, how would the cell potential change? Justify your answer.

(d) How would the cell potential change if the reaction were carried out at 25°C with a 1.0-molar solution of $Mg(NO_3)_2$ and a 0.10-molar solution of $Sr(NO_3)_2$? Explain.

(e) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

Advanced Placement Chemistry: 1996 Free Response Answers

- Question 1 is question 4 in previous years, question 2 is question 1 in previous years and questions 3&4 are questions 2&3 in previous years.
- students are now allowed 10 minutes to answer question 1, after which they must seal that portion of the test.
- [delta] is used to indicate the capital Greek letter.
- [square root] applies to the numbers enclosed in parenthesis immediately following
- All simplifying assumptions are justified within 5%.
- One point deduction for a significant figure or math error, applied only once per problem.
- No credit earned for numerical answer without justification.

7)
(a) two points

The sign of the cell potential will be positive
because (any one is sufficient):

K is greater than 1
the reaction is spontaneous (occurs)
E° for Sr^{2+} is more positive
Standard reduction potential for Sr more negative
E° = + 0.52 V

Note: only 1 point earned for just E° positive because K_{eq} positive.

(b) one point
The oxidizing agent is Mg^{2+}

(c) two point
The cell potential would increase
Since all ions are at 1 M, Q for the system is 1 and E° = (RT/nF) ln K
so as T increases, so should E°

Note: no credit lost if student recognizes K_{eq} dependence on T. For temperature change in this problem, decrease in ln K term is small relative to the term RT/nF

OR
No change, because in the Nernst equation $E_{cell} = E° - (RT/nF) \ln Q$
$\ln Q = 0$, and $E_{cell} = E°$
Note: this second approach earns 1 point only

(d) two points
E_{cell} will increase
In the equation $E_{cell} = E° - (0.0592 / n) \log Q$
$Q = 0.1$ therefore log Q is negative therefore term after E° is positive therefore E_{cell} increases

OR
with the concentration of Mg^{2+} larger than that of Sr^{2+}, Le Chatelier's principle predicts the reaction will have a larger driving force to the right and a more positive E_{cell}

(e) one point
At equilibrium, $E_{cell} = 0$
Note: "balanced", "neutral", or "no net reaction" not accepted

Answer Key
Testname: CH_17_PRAC_TEST_ELECTROCHEMISTRY.TST

MULTIPLE CHOICE. Choose the one alternative that best completes the statement or answers the question.

1) C
ID: chem9b 20.1–1

2) A
ID: chem9b 20.1–2

3) C
ID: chem9b 20.1–3

4) C
ID: chem9b 20.1–5

5) B
ID: chem9b 20.1–9

6) E
ID: chem9b 20.1–10

7) D
ID: chem9b 20.1–11

8) D
ID: chem9b 20.1–12

9) D
ID: chem9b 20.1–15

10) B
ID: chem9b 20.1–19

11) C
ID: chem9b 20.1–23

12) C
ID: chem9b 20.1–25

13) E
ID: chem9b 20.1–27

14) B
ID: chem9b 20.1–28

15) B
ID: chem9b 20.1–32

16) D
ID: chem9b 20.1–33

17) E
ID: chem9b 20.1–35

18) B
ID: chem9b 20.1–37

19) C
ID: chem9b 20.1–40

20) A
ID: chem9b 20.1–44

21) E
ID: chem9b 20.1–45

22) B
ID: chem9b 20.1–47

23) B
ID: chem9b 20.1–49

24) E
ID: chem9b 20.1–51

25) B
ID: chem9b 20.1–59

26) B
ID: chem9b 20.1–60

27) D
ID: chem9b 20.1–62

28) B
ID: chem9b 20.1–65

29) C
ID: chem9b 20.2–2

30) B
ID: chem9b 20.2–10

A.P. Chemistry
Practice Test - Ch. 17: Electrochemistry

(1) LEO goes GER
(C)

(2) method one: $Cr_2O_7^{2-} \Rightarrow Cr_2O_7^{2-} \Rightarrow Cr_2O_7^{2-}$

(+12 / −14) → (+6 / −2)

when $Cr_2O_7^{2-} \rightarrow Cr^{3+}$, the oxidation # goes down from +6 to +3, which is reduction. (A)

method two: $Cr_2O_7^{2-} \rightarrow Cr^{3+}$

↳ $6e^- + 14H^+ + Cr_2O_7^{2-} \rightarrow 2Cr^{3+} + 7H_2O$

Gain electrons = reduction

(3) oxidizing agent = species which is reduced = thing which causes something else to be oxidized

since $Cr_2O_7^{2-}$ is reduced, it is the oxidizing agent (D)

($S_2O_3^{2-}$ is oxidized to $S_4O_6^{2-}$)

$S_2O_3^{2-}$: (+2 / −2) → +4 / −6

$S_4O_6^{2-}$: (+2.5 / −2) → +10 / −12

(4) $\overset{0}{Cu} + \overset{0}{S} \rightarrow \overset{+2}{Cu}\overset{-2}{S}$ redox (C)

none of the other reactions has a change in oxidation #'s

(5) reducing agent = thing which is oxidized Pb = choice (B)

$\overset{0}{Pb} \rightarrow \overset{+2}{PbSO_4}$

PAGE TWO

(6) $Cr_2O_7^{2-}$ (+6, -2) → +12, -14 → +6 (E)

(7) $KMnO_4 \rightarrow K^+ + MnO_4^-$ (+1) (D)

(8) MnO_4^- (+7, -2) → +7, -8 (D)

(9) $2e^- + 2H^+ + BrO^- \rightarrow Br^- + H_2O$

$2e^- + 2H^+ + 2OH^- + BrO^- \rightarrow Br^- + H_2O + 2OH^-$

$2e^- + 2H_2O + BrO^- \rightarrow Br^- + H_2O + 2OH^-$

$2e^- + H_2O + BrO^- \rightarrow Br^- + 2OH^-$

$OH^- + Fe(OH)_2 \rightarrow Fe(OH)_3 + e^-$

$2OH^- + 2Fe(OH)_2 \rightarrow 2Fe(OH)_3 + 2e^-$

$2e^- + H_2O + BrO^- \rightarrow Br^- + 2OH^-$

$2OH^- + 2Fe(OH)_2 \rightarrow Fe(OH)_3 + 2e^-$

(D)

$H_2O + BrO^- + 2Fe(OH)_2 \rightarrow Br^- + Fe(OH)_3$

(10) $H_2O + CN^- \rightarrow CNO^- + 2H^+ + 2e^-$

$2e^- + 2Fe^{3+} \rightarrow 2Fe^{2+}$

$H_2O + CN^- + 2OH^- \rightarrow CNO^- + 2H^+ + 2OH^- + 2e^-$

$2e^- + 2Fe^{3+} \rightarrow 2Fe^{2+}$

$H_2O + CN^- + 2OH^- \rightarrow CNO^- + H_2O$

$2Fe^{3+} \rightarrow 2Fe^{2+}$

(B)

$CN^- + 2OH^- + 2Fe^{3+} \rightarrow 2Fe^{2+} + CNO^-$

(11) $2e^- + 2H^+ + SO_4^{2-} \rightarrow SO_3^{2-} + H_2O$ (C)

or,

$SO_4^{2-} \rightarrow SO_3^{2-}$

(+6) (+4)

(12) Anode = oxidation

(C) $Fe \rightarrow Fe^{3+} + 3e^-$

(13) E

(14) OER
gain electrons reduction

LEO
lose electrons oxidation
Anode = oxidation

Cathode = reduction

Voltaic = galvanic = spontaneous
so e⁻'s go from anode → cathode (B)

(15) B

(16) strongest O.A. = gets reduced most easily, since $F_2 + 2e^- \rightarrow 2F^-$ has largest $\mathcal{E}^\circ$, it gets reduced most easily (in comparison with the standard hydrogen electrode). (D)

(17) E

(18) $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O \quad +1.23$

(A) I don't see any data about Ni^{3+} (or higher)

(B) I don't see any species of Br with an ox # > 0

(C) $Ag \rightarrow Ag^+ + e^- \quad -0.799\,V$ looks good to me

(d) I don't know of any ox #'s for Cu greater than +2

(E) $2Br^- \rightarrow Br_2 + 2e^- \quad -1.065\,V$ also looks good to me

Ag and Br^- should both be oxidized by O_2 in aqueous solution because the overall $\mathcal{E}^\circ_{cell}$ will be > 1

(19) $e^- + Fe^{3+} \rightarrow Fe^{2+} \quad \mathcal{E}^\circ = +0.771\,V$

$Sn^{+2} \rightarrow Sn^{4+} + 2e^- \quad \mathcal{E}^\circ = -0.154\,V$

$+0.617\,V$ (C)

(20) $2e^- + Sn^{4+} \rightarrow Sn^{2+} \quad \mathcal{E}^\circ = +0.154\,V$

$Cr \rightarrow Cr^{3+} + 3e^- \quad \mathcal{E}^\circ = +0.74\,V$

$+0.89\,V$ (A)

(21) $\Delta G = -nF\mathcal{E}$ (E)

(22) (B) $Sn^{4+} + 2e^- \rightarrow Sn^{2+} \quad +0.154\,V$

$Cr \rightarrow Cr^{3+} + 3e^- \quad +0.74$

$+0.89\,V$

(23) $\mathcal{E} = \mathcal{E}^\circ - \frac{0.0592}{n} \log Q = +0.63\,V - \frac{0.0592}{2} \log\left(\frac{1.0}{2\times10^{-4}}\right)$

$= 0.63 - 0.109$

$= 0.52\,V$

$Q = \frac{[Zn^{2+}]}{[Pb^{2+}]} = \frac{1.0}{2\times10^{-4}}$

(B)

$Pb^{2+} + 2e^- \rightarrow Pb$

$Zn \rightarrow Zn^{2+} + 2e^-$ $\rightarrow n = 2$

(24) $2H^+ + 2e^- \rightarrow H_2$ reduction = cathode

$Sn \rightarrow Sn^{2+} + 2e^-$ oxidation = anode

(A) Yes - lower pH = higher $[H^+]$. If $[H^+]$ increases, Q decreases, rxn shifts forward, $\mathcal{E}$ increases.

(B) Yes - higher $[H_2]$ means greater Q, rxn shifts left, $\mathcal{E}$ decreases (same as higher P_{H_2})

(C) yes (E) ✓

(D) yes

(25) (B) (voltaic = galvanic)

(26) $2.19\text{ g Cr} \times \frac{1\text{ mol Cr}}{52\text{ g Cr}} \times \frac{3\text{ mole}^-}{1\text{ mol Cr}} \times \frac{96485\text{ C}}{1\text{ mole}^-} \times \frac{1\text{ S}}{35.2\text{ C}} \times \frac{1\text{ min}}{60\text{ s}} = \underline{5.77}\text{ min}$ (B)

$Cr^{3+} + 3e^- \rightarrow Cr$

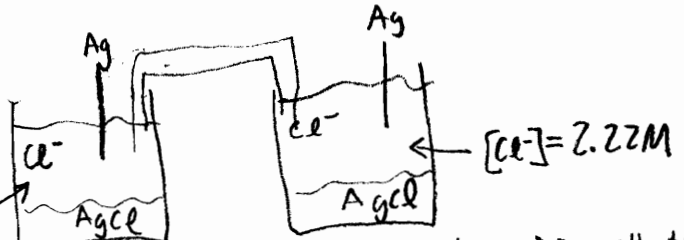
(27) $Ca^{2+} + 2e^- \rightarrow Ca$

$10.0\text{ hrs} \times \frac{60\text{ min}}{1\text{ hr}} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{30.0\text{ C}}{1\text{ S}} \times \frac{1\text{ mole}^-}{96485\text{ C}} \times \frac{1\text{ mol Ca}}{2\text{ mole}^-} \times \frac{40.08\text{ g}}{1\text{ mol Ca}} = \underline{224}\text{ g Ca}$ (D)

(28) $Cu^{2+} + 2e^- \rightarrow Cu$

$25\text{ min} \times \frac{60\text{ s}}{1\text{ min}} \times \frac{2.3\text{ C}}{1\text{ s}} \times \frac{1\text{ mole}^-}{96485\text{ C}} \times \frac{1\text{ mol Cu}}{2\text{ mole}^-} \times \frac{63.55\text{ g}}{1\text{ mol Cu}} = \underline{1.1}\text{ g}$ (B)

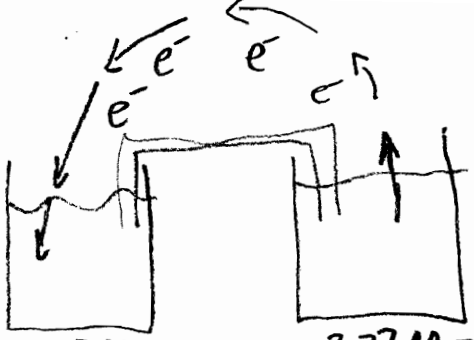
(29) $\mathcal{E} = \mathcal{E}^\circ - \frac{0.0592}{n}\log Q$



continued →

which way will electrons flow spontaneously (voltaic = galvanic = spontaneous)? well, the conc's of Cl^- aren't equal, so compartment on right must decrease $[Cl^-]$ and compartment on left must increase $[Cl^-]$. When AgCl takes an e^-, Cl^- is produced, causing an increase in $[Cl^-]$.

(29) (cont'd)



$[Cl^-] = 0.0222M$
gaining electrons
$[Cl^-]$ increasing
$AgCl_{(s)} + e^- \rightarrow Ag_{(s)} + Cl^-_{(aq)}$

$2.22M = [Cl^-]$
losing electrons
$[Cl^-]$ decreasing
$Ag_{(s)} + Cl^- \rightarrow AgCl_{(s)} + e^-$

		$\mathcal{E}^\circ$
reduction:	$AgCl_{(s)} + e^- \rightarrow Ag_{(s)} + \underset{0.0222M}{Cl^-_{(aq)}}$	0.222 V
oxidation:	$Ag_{(s)} + \underset{2.22M}{Cl^-_{(aq)}} \rightarrow AgCl_{(s)} + e^-$	-0.222 V

$$\underset{2.22M}{Cl^-_{(aq)}} \rightarrow \underset{0.0222M}{Cl^-} \qquad 0.0\ V$$

$$\mathcal{E} = \mathcal{E}^\circ - \frac{.0592}{n}\log Q$$

$$= 0 - \frac{.0592}{1}\log\frac{.0222}{2.22} = 0 - (.0592)(\log 0.01)$$

$$= -.0592(-2)$$

$$= 0.118\ V \quad (C)$$

(30) $4.00\,g \times \frac{1\ mol\ Al}{26.98\,g\ Al} \times \frac{3\ mole^-}{1\ mol\ Al} \times \frac{96485\,C}{1\ mole^-} \times \frac{1\,s}{12.0\,C} = \frac{3580\,s}{3.58\times10^3\,s}$

$Al^{3+} + 3e^- \rightarrow Al$

(B)