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:
   \[ \text{Cr}_2\text{O}_7^{2-} + 6\text{S}_2\text{O}_3^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{S}_4\text{O}_6^{2-} + 7\text{H}_2\text{O} \]
   A) \text{Cr}_2\text{O}_7^{2-}  
   B) \text{S}_2\text{O}_3^{2-}  
   C) \text{H}^+  
   D) \text{S}_4\text{O}_6^{2-}  
   E) \text{Cr}^{3+}

3) __________ is the oxidizing agent in the reaction below.
   \[ \text{Cr}_2\text{O}_7^{2-} + 6\text{S}_2\text{O}_3^{2-} + 14\text{H}^+ \rightarrow 2\text{Cr}^{3+} + 3\text{S}_4\text{O}_6^{2-} + 7\text{H}_2\text{O} \]
   A) \text{Cr}^{3+}  
   B) \text{H}^+  
   C) \text{S}_2\text{O}_3^{2-}  
   D) \text{Cr}_2\text{O}_7^{2-}  
   E) \text{S}_4\text{O}_6^{2-}

4) Which of the following reactions is a redox reaction?
   (a) \[ \text{K}_2\text{CrO}_4 + \text{BaCl}_2 \rightarrow \text{BaCrO}_4 + 2\text{KCl} \]
   (b) \[ \text{Pb}^{2+} + 2\text{Br}^- \rightarrow \text{PbBr} \]
   (c) \[ \text{Cu} + \text{S} \rightarrow \text{CuS} \]
   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?
   \[ \text{Pb} + \text{PbO}_2 + 2\text{H}_2\text{SO}_4 \rightarrow 2\text{PbSO}_4 + 2\text{H}_2\text{O} \]
   A) \text{H}_2\text{O}  
   B) \text{Pb}  
   C) \text{PbO}_2  
   D) \text{PbSO}_4  
   E) \text{H}_2\text{SO}_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, \text{MnO}_4^{-2}?
   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?

\[ \text{BrO}^- + \text{Fe(OH)}_2 \rightarrow \text{Br}^- + \text{Fe(OH)}_3 \] (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?

\[ \text{CN}^- + \text{Fe}^{3+} \rightarrow \text{CNO}^- + \text{Fe}^{2+} \] (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 ________.

\[ 3\text{MnO}_4^- (aq) + 24\text{H}^+ (aq) + 5\text{Fe} (s) \rightarrow 3\text{Mn}^{2+} (aq) + 5\text{Fe}^{3+} (aq) + 12\text{H}_2\text{O} (l) \]

A) $\text{Fe}^{2+} (aq) \rightarrow \text{Fe}^{3+} (aq) + e^-$
B) $\text{MnO}_4^- (aq) + 8\text{H}^+ (aq) + 5e^- \rightarrow \text{Mn}^{2+} (aq) + 4\text{H}_2\text{O} (l)$
C) $\text{Fe} (s) \rightarrow \text{Fe}^{3+} (aq) + 3e^-$
D) $2\text{MnO}_4^- (aq) + 12\text{H}^+ (aq) + 6e^- \rightarrow 2\text{Mn}^{2+} (aq) + 3\text{H}_2\text{O} (l)$
E) $\text{Fe} (s) \rightarrow \text{Fe}^{2+} (aq) + 2e^-$

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

15) $1\text{V} = \text{______}$.  
   A) 96485 C  B) 1 J/C  C) 1 C/J  D) 1 J/s  E) 1 amp $\cdot$ s
<table>
<thead>
<tr>
<th>Half Reaction</th>
<th>E'(V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>F₂(g) + 2e⁻ → 2F⁻(aq)</td>
<td>+2.87</td>
</tr>
<tr>
<td>Cl₂(g) + 2e⁻ → 2Cl⁻(aq)</td>
<td>+1.359</td>
</tr>
<tr>
<td>Br₂(l) + 2e⁻ → 2Br⁻(aq)</td>
<td>+1.065</td>
</tr>
<tr>
<td>O₂(g) + 4H⁺(aq) + 4e⁻ → 2H₂O(l)</td>
<td>+1.23</td>
</tr>
<tr>
<td>Ag⁺ + e⁻ → Ag(s)</td>
<td>+0.799</td>
</tr>
<tr>
<td>Fe³⁺(aq) + e⁻ → Fe²⁺(aq)</td>
<td>+0.771</td>
</tr>
<tr>
<td>I₂(s) + 2e⁻ → 2I⁻(aq)</td>
<td>+0.536</td>
</tr>
<tr>
<td>Cu²⁺ + 2e⁻ → Cu(s)</td>
<td>+0.34</td>
</tr>
<tr>
<td>2H⁺ + 2e⁻ → H₂(g)</td>
<td>0</td>
</tr>
<tr>
<td>Pb²⁺ + 2e⁻ → Pb(s)</td>
<td>−0.126</td>
</tr>
<tr>
<td>Ni²⁺ + 2e⁻ → Ni(s)</td>
<td>−0.28</td>
</tr>
<tr>
<td>Li⁺ + e⁻ → Li(s)</td>
<td>−3.05</td>
</tr>
</tbody>
</table>

16) Which of the halogens in Table 20.1 is the strongest oxidizing agent?
   A) Br₂
   B) I₂
   C) Cl₂
   D) F₂
   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

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

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

<table>
<thead>
<tr>
<th>Half-reaction</th>
<th>$E'\ (V)$</th>
</tr>
</thead>
<tbody>
<tr>
<td>$Cr^{3+}\ (aq) + 3e^- \rightarrow Cr\ (s)$</td>
<td>-0.74</td>
</tr>
<tr>
<td>$Fe^{2+}\ (aq) + 2e^- \rightarrow Fe\ (s)$</td>
<td>-0.440</td>
</tr>
<tr>
<td>$Fe^{3+}\ (aq) + e^- \rightarrow Fe^{2+}\ (s)$</td>
<td>+0.771</td>
</tr>
<tr>
<td>$Sn^{4+}\ (aq) + 2e^- \rightarrow Sn^{2+}\ (aq)$</td>
<td>+0.154</td>
</tr>
</tbody>
</table>

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

<table>
<thead>
<tr>
<th>Half-reaction</th>
<th>( E^\circ ) (V)</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{Cr}^{3+} (aq) + 3e^- \rightarrow \text{Cr} (s) )</td>
<td>-0.74</td>
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<td>+0.771</td>
</tr>
<tr>
<td>( \text{Sn}^{4+} (aq) + 2e^- \rightarrow \text{Sn}^{2+} (aq) )</td>
<td>+0.154</td>
</tr>
</tbody>
</table>

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

A) \( \text{Sn}^{4+} (aq) + \text{Fe}^{2+} (aq) \rightarrow \text{Sn}^{2+} (aq) + \text{Fe} (s) \)
B) \( 3\text{Sn}^{4+} (aq) + 2\text{Cr} (s) \rightarrow 2\text{Cr}^{3+} (aq) + 3\text{Sn}^{2+} (aq) \)
C) \( 3\text{Fe} (s) + 2\text{Cr}^{3+} (aq) \rightarrow 2\text{Cr} (s) + 3\text{Fe}^{2+} (aq) \)
D) \( \text{Sn}^{4+} (aq) + \text{Fe}^{3+} (aq) \rightarrow \text{Sn}^{2+} (aq) + \text{Fe}^{2+} (aq) \)
E) \( 3\text{Fe}^{2+} (aq) \rightarrow \text{Fe} (s) + 2\text{Fe}^{3+} (aq) \)

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

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

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

24) Consider an electrochemical cell based on the reaction:

\( 2\text{H}^+ (aq) + \text{Sn} (s) \rightarrow \text{Sn}^{2+} (aq) + \text{H}_2 (g) \)

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 \([\text{Sn}^{2+}]\) in the anode compartment
D) increasing the pH in the cathode compartment
E) Any of the above will change the measured 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) \( \text{O}_2 \) 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 \( \text{Cr}^{3+} \) 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 CaBr₂ 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 x 10⁻²

29) A voltaic cell is constructed with two silver-silver chloride electrodes, where the half-reaction is
   \[ \text{AgCl (s) + e}^- \rightarrow \text{Ag (s) + Cl}^- \text{ (aq)} \quad E^\circ = +0.222 \text{ V} \]
   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 AlCl₃ with an electrical current of 12.0 A?
   A) 2.90 × 10⁵  B) 3.57 × 10³  C) 27.0  D) 9.00  E) 1.19 × 10³
Advanced Placement Chemistry: 1996 Free Response Questions

7) \[ \text{Sr}(s) + \text{Mg}^{2+} \rightleftharpoons \text{Sr}^{2+} + \text{Mg}(s) \]

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 \( 4.2 \times 10^{17} \) at 25°C.

(a) Predict the sign of the standard cell potential, \( E^\circ \), 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 \( \text{Mg(NO}_3\text{)}_2 \) and a 0.10-molar solution of \( \text{Sr(NO}_3\text{)}_2 \)? Explain.

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

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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.
- \([\text{square root}]\) 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)
- \( E^\circ \) for \( \text{Sr}^{2+} \) is more positive
- Standard reduction potential for \( \text{Sr} \) more negative
- \( E^\circ = +0.52 \text{ V} \)

Note: only 1 point earned for just \( E^\circ \) positive because \( K_{eq} \) positive.

(b) one point

The oxidizing agent is \( \text{Mg}^{2+} \).

(c) two point

The cell potential would increase

Since all ions are at 1 M, \( Q \) for the system is 1 and \( E^\circ = (RT/nF) \ln K \) so as \( T \) increases, so should \( E^\circ \)
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 $E_{\text{cell}} = E^\circ - (RT/nF) \ln Q$
$\ln Q = 0$, and $E_{\text{cell}} = E^\circ$
Note: this second approach earns 1 point only

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

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

(e) one point
At equilibrium, $E_{\text{cell}} = 0$
Note: "balanced", "neutral", or "no net reaction" not accepted
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
Answer Key
Testname: CH_17_PRAC_TEST_ELECTROCHEMISTRY.TST

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
1. LEO goes GER

2. Method one: \[ \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}_2\text{O}_4^{2-} \rightarrow \text{Cr}_2\text{O}_7^{4+} \]
   
   When \( \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+} \), the oxidation # goes down from +6 to +3, which is reduction. (A)

   Method two:
   
   \[ \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{Cr}^{3+} \]
   
   \[ 6e^- + 14H^+ + \text{Cr}_2\text{O}_7^{2-} \rightarrow 2\text{Cr}^{3+} + 7H_2O \] (Gain electrons = Reduction)

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

   Since \( \text{Cr}_2\text{O}_7^{2-} \) is reduced, it is the oxidizing agent.

   \( \text{S}_2\text{O}_3^{2-} \) is oxidized to \( \text{S}_4\text{O}_6^{2-} \)

   \[ \begin{array}{c}
   +1 \quad \text{S}_2\text{O}_3^{2-} \\
   +1 \quad \text{S}_4\text{O}_6^{2-} \\
   +1 \quad \text{S}_2 \quad -6
   \end{array} \]

4. \( \text{Cu} + \text{O} \rightarrow \text{Cu}^{2+} \)

   None of the other reactions has a change in oxidation #. (C)

5. Reducing agent = thing which is oxidized

   \( \text{Pb} = \text{chlorine} \)

   \( \text{Pb}^{2+} \rightarrow \text{PbSO}_4 \)
9. \[ 2e^- + 2H^+ + BrO^- \rightarrow Br^- + H_2O \]
   \[ 2e^- + 2H^+ + 2OH^- + BrO^- \rightarrow Br^- + H_2O + 2OH^- \]
   \[ 2e^- + 2H_2O + BrO^- \rightarrow Br^- + H_2O + 2OH^- \]
   \[ 2e^- + H_2O + BrO^- \rightarrow Br^- + 2OH^- \]

10. \[ OH^- + Fe(OH)_2 \rightarrow Fe(OH)_3 + e^- \]
    \[ 2OH^- + 2Fe(OH)_2 \rightarrow 2Fe(OH)_3 + 2e^- \]

   \[ H_2O + BrO^- + 2Fe(OH)_2 \rightarrow Br^- + Fe(OH)_3 \]
11. \[ 2e^- + 2H^+ + SO_4^{2-} \rightarrow SO_3^{2-} + H_2O \]

12. Anode = oxidation
   \[ Fe \rightarrow Fe^{3+} + 3e^- \]

13. \[ \text{OER} \]
   \[ \text{gain electrons reduction} \]
   Cathode = Reduction

14. \[ \text{LCO} \]
   \[ \text{lose electrons oxidation} \]
   Anode = oxidation

15. \[ B \]

16. Strongest O.A. = gets reduced most easily,
   since \( F_2 + 2e^- \rightarrow 2F^- \) has largest \( E^0 \), it gets reduced most easily (in comparison with the standard hydrogen electrode).

17. \[ E \]
18. $O_2 + 4H^+ + 4e^- \rightarrow 2H_2O \quad E^0 = 1.23$

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. $Ag \rightarrow Ag^+ + e^- \quad E^0 = -0.799V$

D. I don't know of any ox #'s for Cu greater than +2.
E. $2Br^- \rightarrow Br_2 + 2e^- \quad E^0 = -1.065V$

Ag and Br should both be oxidized by $O_2$ in aqueous solution because the overall $E^0_{cell}$ will be $> 1$.

19. $e^- + Fe^{3+} \rightarrow Fe^{2+} \quad E^0 = +0.771V$
$Sn^{+2} \rightarrow Sn^{4+} + 2e^- \quad E^0 = -0.154V$

20. $2e^- + Sn^{4+} \rightarrow Sn^{2+} \quad E^0 = +0.154V$
$Cr \rightarrow Cr^{3+} + 3e^- \quad E^0 = +0.74V$

21. $\Delta G = -nF\Delta E$

22. B. $Sn^{4+} + 2e^- \rightarrow Sn^{2+} + 0.154V$
Cr $\rightarrow Cr^{3+} + 3e^- \quad +0.74V$

23. $E = E^0 - \frac{0.0592}{n} \log Q = +0.63V - \frac{0.0592}{2} \log \left(\frac{1.0}{2} \times 10^{-4}\right)$

$Q = \frac{[Sn^{2+}]}{[Pb^{2+}]} = \frac{10}{2 \times 10^{-4}}$

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

$Zn \rightarrow Zn^{2+} + 2e^- \quad n = 2$
29) \[ 2H^+ + 2e^- \rightarrow H_2 \] reduction = cathode
\[ Sn \rightarrow Sn^{2+} + 2e^- \] oxidation = anode

A) Yes - lower pH = higher [H⁺]. If [H⁺] increases, Q decreases, reaction shifts forward, \( E \) increases.
B) Yes - higher [H₂] means greater Q, reaction shifts left, C decreases
C) Yes
D) Yes

25) (voltaic = galvanic)

26) \[ 2.19 \text{g} \text{Cr} \times \frac{1 \text{mol} \text{Cr}}{52.99 \text{g} \text{Cr}} \times \frac{3 \text{mol}^-}{1 \text{mol} \text{Cr}} \times \frac{96485 \text{C}}{1 \text{mol}^-} \times \frac{1 \text{S}}{35.2 \text{C}} \times \frac{1 \text{min}}{60 \text{S}} = \frac{5.77 \text{min}}{ \text{B} } \]

\[ \text{Cr}^{3+} + 3e^- \rightarrow \text{Cr} \]

27) \[ \text{Ca}^{2+} + 2e^- \rightarrow \text{Ca} \]

\[ 10.0 \text{hrs} \times \frac{60 \text{min}}{1 \text{hr}} \times \frac{60 \text{S}}{1 \text{min}} \times \frac{30.0 \text{C}}{1 \text{S}} \times \frac{1 \text{mol}^-}{96485 \text{C}} \times \frac{1 \text{mol} \text{Ca}}{2 \text{mol}^-} \times \frac{40.085 \text{g}}{1 \text{mol} \text{Ca}} = \frac{224 \text{g Ca}}{ \text{D} } \]

28) \[ \text{Cu}^{2+} + 2e^- \rightarrow \text{Cu} \]

\[ 25 \text{min} \times \frac{60 \text{S}}{1 \text{min}} \times \frac{2.3 \text{C}}{1 \text{S}} \times \frac{1 \text{mol}^-}{96485 \text{C}} \times \frac{1 \text{mol} \text{Cu}}{2 \text{mol}^-} \times \frac{63.55 \text{g}}{1 \text{mol} \text{Cu}} = \frac{1.1 \text{g Cu}}{ \text{B} } \]

29) \[ E = E^0 - \frac{0.0592}{n} \log Q \]

\[ [\text{Cl}^-] = 0.022 \text{M} \]

[continued]

Which way will electrons flow spontaneously (voltaic = galvanic = spontaneous)? Well, the conc's of \( \text{Cl}^- + \text{e}^- \) are equal, so compartment on right must decrease [e⁻] and compartment on left must increase [Cl⁻]. When AgCl takes an e⁻, Cl⁻ is produced, causing an increase in [Cl⁻].
\[
\begin{align*}
[\text{Ce}^-] &= 0.0222 M \\
\text{gaining electrons} \\
[\text{Ce}^-] \text{ increasing} \\
\text{AgCl(s)} + e^- &\rightarrow \text{Ag(s)} + \text{Cl}^- (\text{aq}) \\
\text{2.22M} &= [\text{Ce}^-] \\
\text{losing electrons} \\
\text{[Ce] decreasing} \\
\text{Ag(s)} + \text{Cl}^- &\rightarrow \text{AgCl(s)} + e^- \\
\frac{E^o}{0.222 V} \\
\text{reduction: } \text{AgCl(s)} + e^- &\rightarrow \text{Ag(s)} + \text{Cl}^- (\text{aq}) \\
\text{0.0222 M} \\
\text{oxidation: } \text{Ag(s)} + \text{Cl}^- (\text{aq}) &\rightarrow \text{AgCl(s)} + e^- \\
\text{0.222V} \\
\text{Cl}^- (\text{aq}) &\rightarrow \text{Cl}^- \\
\text{2.22M} \\
\text{0.0222 M} \\
E &= E^o - \frac{.0592}{n} \log Q \\
&= 0 - \frac{.0592}{1} \log \frac{0.222}{2.22} = 0 - (0.0592)(\log 0.01) \\
&= -0.0592(-2) \\
&= 0.118 V \, (C)
\end{align*}
\]

30. \[
4.00 \times \frac{1 \text{ mole Al}}{26.98 g \text{ Al}} \times \frac{3 \text{ mole}^-}{1 \text{ mole Al}} \times \frac{96485 \text{ C}}{1 \text{ mole}^-} \times \frac{1 \text{ s}}{12.0 \text{ C}} = \frac{3.580 \times 10^3 \text{ s}}{3.58 \times 10^3 \text{ s}}
\]

\[
\text{Al}^{3+} + 3e^- \rightarrow \text{Al}
\]