This print-out should have 25 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

How many moles of $Cl_2(g)$ are produced by the electrolysis of concentrated sodium chloride if 2.00 A are passed through the solution for 4.00 hours? The equation for this process (the "chloralkali" process) is

$$2\,NaCl(aq) + 2\,H_2O(\ell) \rightarrow 2\,NaOH(aq) + H_2(g) + Cl_2(g)$$

1. 0.149 mol **correct**

2. 0.447 mol

3. 0.298 mol

4. 0.0745 mol

5. 0.00248 mol

Explanation:

002 10.0 points

A steel surface has been electroplated with 5.10 g of vanadium (V, molar mass = 51 g/mol). If 2.90×10^4 C of charge were used, what was the original oxidation number of V?

1. +4

2. +3 **correct**

3. +5

4. +6

5. +1

6. +2

Explanation:

$$\text{moles V} = \frac{5.1\text{ g}}{51\text{ g/mol}} = 0.1\text{ mol}$$

$$\text{moles } e^- = \frac{2.90 \times 10^4\text{ C}}{9.65 \times 10^4\text{ C/mol}} = 0.3\text{ mol}$$

$$\text{ox. number} = \frac{0.3\text{ mol } e^-}{0.1\text{ mol V}} = +3$$

003 10.0 points

How long will it take to deposit 0.00235 moles of gold by the electrolysis of $KAuCl_4(aq)$ using a current of 0.214 amperes?

1. 26.5 min

2. 17.7 min

3. 106 min

4. 70.7 min

5. 53.0 min **correct**

Explanation:

004 10.0 points

Consider 3 electrolysis experiments:

The first: 1 Faraday of electricity is passed through a solution of $AgNO_3$.

The second: 2 Faradays of electricity are passed through a solution of $Zn(NO_3)_2$.

The third: 3 Faradays of electricity are passed through a solution of $Bi(NO_3)_3$.

1. Equal numbers of moles of all three metals are produced. **correct**

2. Twice as many moles of metallic zinc are produced than metallic silver.

3. The reaction producing the smallest mass of metal is that of the silver solution.

4. Equal masses of all three metals are produced.

Explanation:

1 F = 1 mol e^-.

The relevant half-reactions are

$$Ag^+ + 1\,e^- \rightarrow Ag$$
$$Zn^{2+} + 2\,e^- \rightarrow Zn$$
$$Bi^{3+} + 3\,e^- \rightarrow Bi$$

Using the given amounts of electricity,

Exp.	metal produced	grams produced
1	1 mol Ag	$(1 \text{ mol Ag}) \dfrac{107.87 \text{ g Ag}}{\text{mol Ag}}$
2	1 mol Zn	$(1 \text{ mol Zn}) \dfrac{65.39 \text{ g Zn}}{\text{mol Zn}}$
3	1 mol Bi	$(1 \text{ mol Bi}) \dfrac{208.98 \text{ g Bi}}{\text{mol Bi}}$

Of the answer choices, the true statement says that equal moles of metals are produced.

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005 10.0 points

Sodium is produced by electrolysis of molten sodium chloride. What are the products at the anode and cathode, respectively?

1. Cl_2(g) and $Na_2O(\ell)$

2. Cl_2(g) and $Na(\ell)$ **correct**

3. $Na(\ell)$ and O_2(g)

4. Cl^-(aq) and $Na_2O(\ell)$

5. O_2(g) and $Na(\ell)$

Explanation:

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006 10.0 points

What is the standard cell potential of the strongest battery that could be made using these half reactions?

$Br_2 + 2\,e^- \longrightarrow 2\,Br^-$ $\quad E^\circ = +1.07$

$Fe^{3+} + 3\,e^- \longrightarrow Fe$ $\quad E^\circ = -0.04$

$Co^{3+} + e^- \longrightarrow Co^{2+}$ $\quad E^\circ = +1.80$

$Zn^{2+} + 2\,e^- \longrightarrow Zn$ $\quad E^\circ = -0.76$

1. 1.11

2. 1.84

3. 1.83

4. 1.03

5. 1.04

6. 2.56 **correct**

Explanation:

The strongest reducing agent is Co^{3+} and the strongest oxidizing agent is Zn. The standard cell potential of a battery built from these species would be:

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= 1.80 - (-0.76) \\ &= 2.56 \end{aligned}$$

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007 10.0 points

What would be the E° cell of an electrolytic cell made from the half reactions

$AgCl(s) + e^- \longrightarrow Ag(s) + Cl^-(aq)$ $\quad E^\circ = +0.22$ V

$Al^{3+}(aq) + 3\,e^- \longrightarrow Al(s)$ $\quad E^\circ = -1.66$ V

1. −1.44

2. 1.44

3. −1.88 **correct**

4. 1.88

Explanation:

$$\begin{aligned} E^\circ &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= -1.66 - 0.22 = -1.88 \end{aligned}$$

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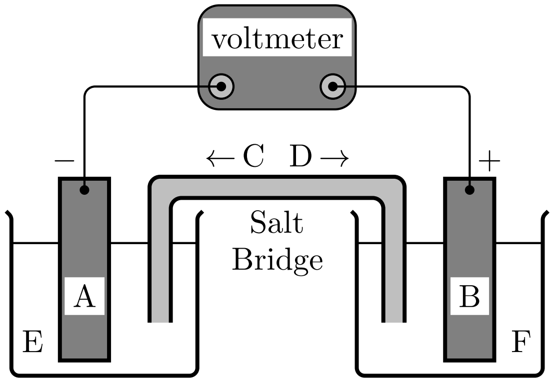
008 (part 1 of 3) 10.0 points

The galvanic cell below uses the standard half-cells $Mg^{2+}\,|\,Mg$ and $Zn^{2+}\,|\,Zn$, and a salt bridge containing KCl(aq). The voltmeter gives a positive voltage reading.



Identify A and write the half-reaction that occurs in that compartment.

1. Zn(s); $Zn^{2+}(aq) + 2\,e^- \rightarrow Zn(s)$

2. Mg(s); $Mg^{2+}(aq) + 2\,e^- \rightarrow Mg(s)$

3. Zn(s); $Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$

4. Mg(s); $Mg(s) \rightarrow Mg^{2+}(aq) + 2\,e^-$ **correct**

Explanation:

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009 (part 2 of 3) 10.0 points

What happens to the size of the electrode A during the operation of the cell?

1. increases

2. No change

3. decreases **correct**

Explanation:

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010 (part 3 of 3) 10.0 points

What is the voltmeter reading?

1. +4.30 V

2. +3.40 V

3. +2.50 V

4. +0.50 V

5. +1.60 V **correct**

Explanation:

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011 10.0 points

The electrolysis of an aqueous sodium chloride solution using inert electrodes produces gaseous chlorine at one electrode. At the other electrode gaseous hydrogen is produced and the solution becomes basic around the electrode. What is the equation for the cathode half-reaction in the electrolytic cell?

1. $2\,Cl^- \rightarrow Cl_2 + 2\,e^-$

2. $2\,H_2O + 2\,e^- \rightarrow H_2 + 2\,OH^-$ **correct**

3. $Cl_2 + 2\,e^- \rightarrow 2\,Cl^-$

4. $H_2 + 2\,OH^- \rightarrow 2\,H_2O + 2\,e^-$

5. None of the other answers listed is correct.

Explanation:

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012 10.0 points

What is the E°_{cell} of

$$Zn(s) \mid Zn^{2+}(aq) \parallel Ce^{4+}(aq) \mid Ce^{3+}(aq)$$

$Zn^{2+} + 2\,e^- \rightarrow Zn$ $\quad E^\circ_{red} = -0.76$
$Ce^{4+} + e^- \rightarrow Ce^{3+}$ $\quad E^\circ_{red} = +1.61$

1. −0.85

2. +1.61

3. −2.37

4. +2.37 **correct**

5. +0.85

Explanation:

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013 10.0 points

Calculate the cell potential for a cell based on the reaction

$$Cu(s) + 2\,Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2\,Ag(s)$$

when the concentrations are as follows: $[Ag^+] = 0.51$ M, $[Cu^{2+}] = 0.9$ M. (The temperature is 25°C and $E^0 = 0.4624$ V.)

Correct answer: 0.446443 V.

Explanation:

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014 10.0 points

Standard reduction potentials are established by comparison to the potential of which half reaction?

1. $F_2 + 2\,e^- \longrightarrow 2\,F^-$

2. $Li^+ + e^- \longrightarrow Li$

3. $2\,H^+ + 2\,e^- \longrightarrow H_2$ **correct**

4. $2\,H_2O + 2\,e^- \longrightarrow H_2 + 2\,OH^-$

5. $Na^+ + e^- \longrightarrow Na$

Explanation:

The hydrogen electrode is the standard reference electrode

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015 10.0 points

Consider the cell

$Pb(s)\,|\,PbSO_4(s)\,|\,SO_4^{2-}(aq\,, 0.60\text{ M})\,||$
$H^+(aq\,, 0.70\text{ M})\,|\,H_2(g\,, 192.5\text{ kPa})\,|\,Pt\,.$

If E° for the cell is 0.36 V at 25°C, write the Nernst equation for the cell at this temperature.

1. $E = 0.36 - 0.01285\ \ln\left[\dfrac{1.90}{(0.70)^2(0.60)}\right]$ **correct**

2. $E = 0.36 - 0.02569\ \ln\left[\dfrac{192.5}{(0.70)^2(0.60)}\right]$

3. $E = 0.36 - 0.01285\ \ln\left[\dfrac{1.90}{(0.70)(0.60)}\right]$

4. $E = 0.36 + 0.01285\ \ln\left[\dfrac{1.9}{(0.70)^2(0.60)}\right]$

5. $E = 0.36 + 0.01285\ \ln\left[\dfrac{192.5}{(0.70)^2(0.60)}\right]$

Explanation:

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016 10.0 points

A concentration cell consists of the same redox couples at the anode and the cathode, with different concentrations of the ions in the respective compartments. Find the unknown concentration for the following cell.

$Pb(s)\,|\,Pb^{2+}(aq,\,?)\,||$
$Pb^{2+}(aq,\,0.1\text{ M})\,|\,Pb(s) \quad E = 0.025\text{ V}$

Correct answer: 0.0142836 M.

Explanation:

$E_{cell} = 0.025\text{ V}$ $\quad M = 0.1\text{ M}$
$E^\circ_{cell} = 0\text{ V}$ $\quad F = 96485\text{ C/mol}$
$\dfrac{R\,T}{F} = 0.025693\text{ V}$

The cell reaction is

$Pb^{2+}(aq,\,0.1\text{ M}) \to Pb^{2+}(aq, x)$
$n = 2, \quad E_{cell} = 0.025\text{ V}$

Using the Nernst equation,

$$E_{cell} = E^\circ_{cell} - \frac{R\,T}{n\,F}\ln Q$$

$$\frac{R\,T}{n\,F}\ln\left(\frac{x}{M}\right) = E^\circ_{cell} - E_{cell}$$

$$\ln\left(\frac{x}{M}\right) = \frac{n\,F}{R\,T}(E^\circ_{cell} - E_{cell})$$

$$x = M\exp\left[\frac{n\,F}{R\,T}(E^\circ_{cell} - E_{cell})\right]$$

$$x = (0.1\text{ M})\exp\left[\frac{2(0\text{ V} - 0.025\text{ V})}{0.025693\text{ V}}\right]$$

$$= 0.0142836\text{ M}\,.$$

Thus $[Pb^{2+}] = 0.0142836\text{ M}\,.$

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017 10.0 points

What is ratio of $[Co^{2+}]\ |\ [Ni^{2+}]$ when a battery built from the two half reactions

$Ni^{2+} \longrightarrow Ni \qquad E^\circ = -0.25\text{ V}$
$Co^{2+} \longrightarrow Co \qquad E^\circ = -0.28\text{ V}$

reaches equilibrium?

1. 3.20

2. 0.10

3. 0.31

4. 10.23 **correct**

Explanation:

$E^\circ_{cell} = +0.03\text{ V}$

$$E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{0.05916}{N_e}\log Q$$
$$0 = 0.03 - \frac{0.05916}{2}\log\frac{[\mathrm{Co}^{2+}]}{[\mathrm{Ni}^{2+}]}$$
$$\log\frac{[\mathrm{Co}^{2+}]}{[\mathrm{Ni}^{2+}]} = 1.01$$
$$\frac{[\mathrm{Co}^{2+}]}{[\mathrm{Ni}^{2+}]} = 10^{1.01} = 10.23$$

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018 10.0 points

If E° for the disproportionation of $Cu^+(aq)$ to $Cu^{2+}(aq)$ and Cu(s) is +0.37 V at 25°C, calculate the equilibrium constant for the reaction.

1. 1.3×10^3

2. 3.2×10^{12}

3. 2.4×10^2

4. 5.7×10^{18}

5. 1.8×10^6 **correct**

Explanation:

The disproportionation is

$2\,Cu^+ \rightarrow Cu^{2+} + Cu$

This corresponds to a ONE electron transfer from one Cu^+ ion to another Cu^+ ion. Therefore $n = 1$ for this reaction.

$$nFE = RT\ln K$$
$$K = e^{nFE/RT} = e^{1(0.37)/0.257} = 1.8 \times 10^6$$

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019 10.0 points

You turn on a flashlight containing brand new NiCad batteries and keep it lit for a minute or two. Which of the following can be considered TRUE regarding the chemical state of these batteries?

I. ΔG for the battery reaction is negative.
II. $E_{\text{cell}} > 0$.
III. The batteries are at equilibrium.
IV. E_{cell} is substantially decreasing during this time.

1. All

2. II and IV only

3. III only

4. All but IV

5. All but III

6. Maybe IV and I

7. I and II only **correct**

Explanation:

(I and II: TRUE) ΔG must be negative because the light IS on and therefore a spontaneous change is occuring. Same logic goes for the E_{cell} except opposite in sign (positive).

(III: FALSE) The batteries are headed toward equilibrium but are not there yet. True equilibrium would mean a potential of zero volts - the ultimate DEAD battery.

(IV: FALSE) The reason batteries work as well as they do is because they hold a very constant potential throughout most of the life of the battery. Only near the battery's end does the potential start to fall off from the brand new potential. Why? Most batteries rely on solids in their equilibria which means constant voltage as long as the solid reactants (and products) are present (Q in the Nernst equation will equal 1). Once the solids are depleted, you see the great change in potential and usually the failure of the battery to work in its intended device.

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020 10.0 points

What is ΔG° for the half reaction below?

Reaction	E°
$ClO_3^- + 6\,H^+(aq) \longrightarrow \frac{1}{2}\,Cl_2(g) + 3\,H_2O(\ell)$	+1.47

1. $194,000\ \mathrm{kJ \cdot mol^{-1}}$

2. $-1,418\ \mathrm{kJ \cdot mol^{-1}}$

3. $-709\ \mathrm{kJ \cdot mol^{-1}}$ **correct**

4. $194\ \mathrm{kJ \cdot mol^{-1}}$

Explanation:

$$\begin{aligned}\Delta G &= -n_e FE \\ &= -5(96,485)(1.47) \\ &= -709,165\ \text{J}\cdot\text{mol}^{-1} \\ &= -709\ \text{kJ}\cdot\text{mol}^{-1}\end{aligned}$$

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021 10.0 points

For the reduction of Cu^{2+} by Zn, $\Delta G^\circ = -\ 212$ kJ/mol and $E^\circ = +1.10$ V. If the coefficients in the chemical equation for this reaction are multiplied by 2, $\Delta G^\circ = -\ 424$ kJ/mol. This means $E^\circ = +2.20$ V.

1. False **correct**

2. True

Explanation:

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022 10.0 points

Consider the cell

$$\text{Zn(s)}\,|\,\text{Zn}^{2+}\text{(aq)}\,||\,\text{Fe}^{2+}\text{(aq)}\,|\,\text{Fe(s)}$$

at standard conditions. Calculate the value of ΔG°_r for the reaction that occurs when current is drawn from this cell.

1. $-\ 62\ \text{kJ}\cdot\text{mol}^{-1}$ **correct**

2. $+\ 230\ \text{kJ}\cdot\text{mol}^{-1}$

3. $-\ 31\ \text{kJ}\cdot\text{mol}^{-1}$

4. $+\ 62\ \text{kJ}\cdot\text{mol}^{-1}$

5. $-\ 230\ \text{kJ}\cdot\text{mol}^{-1}$

Explanation:

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023 10.0 points

The standard potential of the cell

$$\text{Pb(s)}\,|\,\text{PbSO}_4\text{(s)}\,|\,\text{SO}_4^{2-}\text{(aq)}\,||\,\text{Pb}^{2+}\text{(aq)}\,|\,\text{Pb(s)}$$

is +0.23 V at 25°C. Calculate the equilibrium constant for the reaction of 1 M Pb^{2+}(aq) with 1 M SO_4^{2-}(aq).

1. 3.7×10^{16}

2. 1.7×10^{-8}

3. 8.0×10^{17}

4. 6.0×10^{7} **correct**

5. 7.7×10^{3}

Explanation:

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024 10.0 points

The standard voltage of the cell

$$\text{Ag(s)}\,|\,\text{AgBr(s)}\,|\,\text{Br}^-\text{(aq)}\,||\,\text{Ag}^+\text{(aq)}\,|\,\text{Ag(s)}$$

is +0.73 V at 25°C. Calculate the equilibrium constant for the cell reaction.

1. 2.2×10^{12} **correct**

2. 4.6×10^{-13}

3. 5.1×10^{14}

4. 2.0×10^{-15}

5. 3.9×10^{-29}

Explanation:

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025 10.0 points

The equilibrium constant for the reaction

$$2\,\text{Hg}(\ell) + 2\,\text{Cl}^-\text{(aq)} + \text{Ni}^{2+}\text{(aq)} \rightarrow \text{Ni(s)} + \text{Hg}_2\text{Cl}_2\text{(s)}$$

is 5.6×10^{-20} at 25°C. Calculate the value of E° for a cell utilizing this reaction.

1. + 1.14 V

2. − 0.57 V **correct**

3. + 0.57 V

4. − 1.14 V

5. − 0.25 V

Explanation: