This print-out should have 27 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 (part 1 of 2) 10.0 points Using oxidation and reduction half-reactions,

balance the skeletal equation

$$N_2H_4(g) + ClO_3^-(aq) \rightarrow NO(g) + Cl^-(aq)$$

of hydrazine with chlorate ions. The reaction takes place in basic solution. What is the smallest possible integer coefficient of NO in the balanced equation?

Correct answer: 6.

Explanation: Reduction half reaction: Balance O: $N_2H_4(g) + 2H_2O(\ell) \rightarrow 2NO(g)$ Balance H: $N_2H_4(g) + 2H_2O(\ell) + 8OH^-(aq) \rightarrow$ $2 \operatorname{NO}(g) + 8 \operatorname{H}_2 O(\ell)$ Balance charge: $N_2H_4(g) + 8 OH^-(aq) \rightarrow$ $2 \text{ NO(g)} + 6 \text{ H}_2 \text{O}(\ell) + 8 e^{-1}$ Oxidation half reaction: Balance O: $\mathrm{ClO}_3^-(\mathrm{aq}) \rightarrow \mathrm{Cl}^-(\mathrm{aq}) + 3 \,\mathrm{H_2O}(\ell)$ Balance H: $\text{ClO}_3^-(\text{aq}) + 6 \text{H}_2\text{O}(\ell) \rightarrow$ $Cl^{-}(aq) + 3H_2O(\ell) + 6OH^{-}(aq)$ Balance charge: $ClO_3^-(aq) + 3H_2O(\ell) + 6e^- \rightarrow Cl^-(aq) + 6OH^-(aq)$ Balance e^- : $3 [N_2H_4(g) + 8 OH^-(aq) \rightarrow$ $2 \text{ NO(g)} + 6 \text{ H}_2 \text{O}(\ell) + \frac{8 e^{-}}{2}$ $4 \left[\text{ClO}_3^-(\text{aq}) + 3 \operatorname{H}_2 \operatorname{O}(\ell) + \frac{6 e^-}{2} \right] \rightarrow$ $Cl^{-}(aq) + 6 OH^{-}(aq)$ Add the half reactions: $3 N_2 H_4(g) + 4 ClO_3^-(aq) \rightarrow$ $6 \text{ NO}(g) + 6 \text{ H}_2 O(\ell) + 4 \text{ Cl}^{-}(aq)$

002 (part 2 of 2) 10.0 points

Identify the oxidizing agent in the reaction.

 $\mathbf{2.}~\mathrm{Cl}^{-}$

3. ClO_3^- correct

4. N_2H_4

Explanation:

003 (part 1 of 2) 10.0 points In the reaction

$$\operatorname{Cl}_2(g) + \operatorname{S}_2\operatorname{O}_3^{2-}(aq) \to \operatorname{Cl}^-(aq) + \operatorname{SO}_4^{2-}(aq)$$

of the thiosulfate ion with chlorine gas in an acidic solution, identify the oxidizing agent.

1. SO_4^{2-}

2. Cl⁻

3. Cl_2 correct

4. $S_2O_3^{2-}$

Explanation:

004 (part 2 of 2) 10.0 points

Balance the equation using oxidation and reduction half-reactions. What is the smallest possible integer coefficient of Cl⁻ in the combined balanced equation?

Correct answer: 8.

Explanation:

Consider the half-reactions:

$$\operatorname{Cl}_2(g) + 2 e^- \to 2 \operatorname{Cl}^-(\operatorname{aq})$$

$$S_2O_3^{2-}(aq) + 5 H_2O(\ell) \rightarrow 2 SO_4^{2-}(aq) + 10 H^+(aq) + 8 e^{-1}$$

Balance the e^{-1} :

$$4\left[\operatorname{Cl}_2(g) + \frac{2e^-}{2e^-} \to 2\operatorname{Cl}^-(\operatorname{aq})\right]$$

 $\begin{array}{l} \mathrm{S}_{2}\mathrm{O}_{3}^{2-}(\mathrm{aq}) + 5\,\mathrm{H}_{2}\mathrm{O}(\ell) \rightarrow \\ & 2\,\mathrm{SO}_{4}^{2-}(\mathrm{aq}) + 10\,\mathrm{H}^{+}(\mathrm{aq}) + 8\,e^{-} \\ \mathrm{Add\ the\ half-reactions:} \end{array}$

 $\begin{array}{l} 4 \operatorname{Cl}_2(g) + \operatorname{S}_2 \operatorname{O}_3^{2-}(aq) + 5 \operatorname{H}_2 \operatorname{O}(\ell) \to \\ 8 \operatorname{Cl}^-(aq) + 2 \operatorname{SO}_4^{2-}(aq) + 10 \operatorname{H}^+(aq) \end{array}$

005 10.0 points

Consider the half-reactions and the balanced equation for the cell reaction represented by the skeletal equation

$$Mn(s) + Ti^{2+}(aq) \rightarrow Mn^{2+}(aq) + Ti(s).$$

What is the proper cell diagram for this reaction?

1. $Ti^{2+}(aq) | Ti(s) || Mn(s) | Mn^{2+}(aq)$

2.
$$Ti(s) | Ti^{2+}(aq) || Mn^{2+}(aq) | Mn(s)$$

3. $Mn(s) | Mn^{2+}(aq) || Ti^{2+}(aq) | Ti(s)$ correct

4.
$$Mn^{2+}(aq) | Mn(s) || Ti(s) | Ti^{2+}(aq)$$

Explanation:

The two half-reactions, written as reductions, are

$$Mn^{2+}(aq) + 2 e^{-} \rightarrow Mn(s)$$
$$Ti^{2+}(aq) + 2 e^{-} \rightarrow Ti(s)$$

Equate e^- :

$$\begin{aligned} \mathrm{Ti}^{2+}(\mathrm{aq}) + \frac{2e^{-}}{2e^{-}} &\to \mathrm{Ti}(\mathrm{s}) \\ \mathrm{Mn}(\mathrm{s}) &\to \mathrm{Mn}^{2+}(\mathrm{aq}) + \frac{2e^{-}}{2e^{-}} \end{aligned}$$

Add the balanced half reactions:

$$\mathrm{Mn}(s) + \mathrm{Ti}^{2+}(\mathrm{aq}) \to \mathrm{Mn}^{2+}(\mathrm{aq}) + \mathrm{Ti}(s)$$

The cell diagram is

$$Mn(s) | Mn^{2+}(aq) || Ti^{2+}(aq) | Ti(s)$$



In this electrochemical cell, what is the reduction half reaction?

1. $\operatorname{Zn}^{2+}(\operatorname{aq}) + 2 e^{-} \to \operatorname{Zn}(\operatorname{s})$ 2. $\operatorname{Zn}(\operatorname{s}) \to \operatorname{Zn}^{2+}(\operatorname{aq}) + 2 e^{-}$ 3. $\operatorname{Cu}^{2+}(\operatorname{aq}) + 2 e^{-} \to \operatorname{Cu}(\operatorname{s})$ correct 4. $\operatorname{Cu}(\operatorname{s}) \to \operatorname{Cu}^{2+}(\operatorname{aq}) + 2 e^{-}$ Explanation:

$$\operatorname{Zn}(s) + \operatorname{Cu}^{2+}(aq) \to \operatorname{Zn}^{2+}(aq) + \operatorname{Cu}(s)$$

Reduction occurs at the cathode. In this cell the reduction half reaction is

$$\operatorname{Cu}^{2+}(\operatorname{aq}) + 2e \to \operatorname{Cu}(s)$$

 Cu^{2+} cations are attracted to the solid Cu electrode where they are reduced to Cu(s).

007 10.0 points

What is the standard cell potential of a battery made from the half reactions $2 H^+ + 2 e^- \longrightarrow H_2$ $E^\circ = 0.00 V$ $O_2 + 4 H^+ + 4 e^- \longrightarrow 2 H_2 O$ $E^\circ = +1.23 V$ **1.** 1.23 correct **2.** 2.46 **3.** -2.46 **4.** -1.23 Explanation:

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$
$$= 1.23 - 0.00 = 1.23$$

008 10.0 points

In an electrolytic cell, the negative terminal is the (cathode/anode) and is the site of the (oxidation/reduction) half reaction.

1. cathode, oxidation

- 2. cathode, reduction correct
- 3. anode, reduction
- 4. anode, oxidation

Explanation:

In an electrolytic cell (analogous to charging a battery) the cathode is always attributed a negative sign. By definition, the cathode is the site of reduction, regardless of the type of electrochemical cell and whether it is charging or discharging.

009 10.0 points

In a galvanic cell,

1. electrolytes are added to carry electrons between electrodes.

2. oxidation and reduction take place at the same time but at different electrodes. correct

3. oxidation takes place at the cathode.

4. electrical energy is used to reverse spontaneous chemical reactions.

Explanation:

The reaction in a galvanic cell is spontaneous and does not need to be driven by a battery. Both reduction and oxidation occur.

010 10.0 points

Write the half reactions and the balanced equation for the galvanic cell

 $Ag(s) | AgCl(s) | Cl^{-}(aq) | |$

 $Cl^{-}(aq) | Hg_2Cl_2(s) | Hg(\ell)$ What is the smallest possible integer coefficient of $Hg_2Cl_2(s)$ in the combined balanced equation?

Correct answer: 1.

Explanation:

The two half-reactions, written as reductions, are

$$\begin{split} \operatorname{AgCl}(s) + e^- &\to \operatorname{Ag}(s) + \operatorname{Cl}^-(\operatorname{aq}) \\ & E^\circ = 0.22 \text{ V} \\ \operatorname{Hg_2Cl_2}(s) + 2 \, e^- &\to 2 \operatorname{Hg}(\ell) + 2 \operatorname{Cl}^-(\operatorname{aq}) \\ & E^\circ = 0.27 \text{ V} \\ \operatorname{Equate the} e^-: \\ \operatorname{Hg_2Cl_2}(s) + \frac{2 \, e^-}{2 \, e^-} &\to 2 \operatorname{Hg}(\ell) + 2 \operatorname{Cl}^-(\operatorname{aq}) \\ 2 \left[\operatorname{Ag}(s) + \operatorname{Cl}^-(\operatorname{aq}) &\to \operatorname{AgCl}(s) + \frac{e^-}{2}\right] \\ \operatorname{Add the balanced half reactions:} \\ \operatorname{Hg_2Cl_2}(s) + 2 \operatorname{Ag}(s) &\to 2 \operatorname{AgCl}(s) + 2 \operatorname{Hg}(\ell) \\ & E^\circ_{\operatorname{cell}} = 0.27 \text{ V} - 0.22 \text{ V} = 0.05 \text{ V} \end{split}$$

011 10.0 points

Silver is plated on copper by immersing a piece of copper into a solution containing silver(I) ions. In the plating reaction, copper

1. is oxidized and is the reducing agent. correct

2. is reduced and is the reducing agent.

3. is reduced and is the oxidizing agent.

4. is oxidized and is the oxidizing agent.

Explanation:

This is a displacement reaction. The copper displaces the silver ions from solution, giving the reaction

 $\mathrm{Ag^{+}} + \mathrm{Cu} \rightarrow \mathrm{Cu^{+}} + \mathrm{Ag}$

Copper loses electrons to form the ion, so copper is oxidized. Since copper gives electrons to silver, it causes silver to be reduced. Therefore, copper is the reducing agent.

012 10.0 points

What is the E^0 for the following electrochemical cell where Zn is the cathode?

Fe | Fe²⁺(1.0 M) || $Zn^{2+}(1.0 M)$ | Zn

 $E^{0}(\text{Zn}) = -0.76$ $E^{0}(\text{Fe}) = -0.44$ **1.** +1.20 **2.** -1.20

3. +0.32

4. -0.32 **correct**

Explanation:

$$E_{\text{cell}} = E_{(\text{cathode})}^0 - E_{(\text{anode})}^0$$

= -0.76 - (-0.44)
= -0.32

013 10.0 points

Which of the metals in the list below will react with 1 M H₂SO₄ to produce hydrogen gas?

E^{o} (volts)
-2.714
-0.403
-0.126
+0.337

1. Na and Cd only

2. Na only

- 3. Na, Cd, and Pb only correct
- 4. Na, Cd, Pb, and Cu

5. some other combination than those listed

Explanation:

 $\begin{array}{c|cccc} 014 & 10.0 \text{ points} \\ \text{Consider the voltaic cell:} \\ \text{Pt} & \left| \operatorname{Sn}^{2+} (0.10 \text{ M}), \operatorname{Sn}^{4+} (0.0010 \text{ M}) \right| \\ & & \left| \right| \operatorname{Ag}^+ (0.010 \text{ M}) \left| \operatorname{Ag} \right| \\ \text{Sn}^{4+} + 2 \, e^- \to \operatorname{Sn}^{2+} & E^0 = +0.15 \text{ V} \\ \operatorname{Ag}^+ + 1 \, e^- \to \operatorname{Ag}(\text{s}) & E^0 = +0.80 \text{ V} \\ \text{The electrons flow in the external circuit} \\ \text{from} \end{array}$

1. Ag to Pt.

2. Pt to Ag. correct

3. Ag to Sn^{4+} .

4. Sn to Ag.

5.
$$Sn^{2+}$$
 to Ag^+ .

Explanation:

In the cell notation the anode reaction shows on the left of the salt bridge and the cathode on the right side. As this is a voltaic cell electrons flow from the anode via the Pt electrode through the external circuit to the Ag cathode.

015 10.0 points

Using the standard potential tables, what is the largest approximate E^0 value that can be achieved when two half cell reactions are combined to form a battery?

1. −6 V
2. −3 V

3. 3 V

4.6 V correct

Explanation:

Using the tables, the largest values of $E_{\rm red}^{\circ}$ are about + 3 V and - 3 V.

The species with the positive value would be reduced, the other would be oxidised so: $E_{\text{cell}}^{\circ} = +3 \text{ V} - (-3 \text{ V}) = +6 \text{ V}$

016 10.0 points r the cell

Consider the cell

 $\operatorname{Zn}(s) | \operatorname{Zn}^{2+}(aq) || \operatorname{Cl}^{-}(aq) | \operatorname{AgCl}(s) | \operatorname{Ag}(s)$

Calculate E° .

1. - 0.54 V
2. - 1.20 V
3. + 0.54 V
4. + 0.98 V correct

5. + 1.20 V

Explanation:

017 10.0 points

Which species will oxidize Cr^{2+} but not Mn^{2+} ?

1. O_3 in acidic medium

2. V^{3+} correct

3. Fe²⁺

4. Zn²⁺

5. Pb⁴⁺

Explanation:

018 10.0 points

If the standard potentials for the couples $Cu^{2+} | Cu, Ag^+ | Ag$, and $Fe^{2+} | Fe$ are +0.34, +0.80, and - 0.44 V, respectively, which is the strongest reducing agent?

1. Fe²⁺

2. Ag⁺

3. Ag

4. Cu

5. Fe correct

Explanation:

019 10.0 points

In a working electrochemical cell (+ cell voltage), the cations in the salt bridge move toward the cathode.

1. True correct

2. False

Explanation:

020 10.0 points For the cell diagram $Cd(s) | CdSO_4(aq) || Hg_2SO_4 | Hg(\ell)$

what reaction occurs at the cathode?

1.
$$CdSO_4(s) + 2e^- \rightarrow 2Cd(\ell) + SO_4^{2-}(aq)$$

2. $\operatorname{Hg}_2\operatorname{SO}_4(s) + 2e^- \rightarrow 2\operatorname{Hg}(\ell) + \operatorname{SO}_4^{2-}(\operatorname{aq})$ correct

Explanation:

021 10.0 points What is the cathode in

 $Mg(s) \mid Mg^{2+}(aq) \mid \mid Au^+(aq) \mid Au(s)$

 $Mg^{2+} + 2e^{-} \rightarrow Mg \qquad \qquad \mathcal{E}^{\circ}_{red} = -2.36$ Au⁺ + e⁻ \rightarrow Au $\qquad \qquad \mathcal{E}^{\circ}_{red} = +1.69$ and what is the cell type?

- **1.** Mg(s); an electrolytic cell
- **2.** Not enough information is provided.
- **3.** Au(s); a voltaic cell **correct**
- **4.** Mg(s); a voltaic cell
- **5.** Au(s) an electrolytic cell

Explanation:

The diagram A \mid B $\mid\mid$ C \mid D is read as follows:

 $A \rightarrow B + n e^-$ (oxidation)

 $C + m e^- \rightarrow D$ (reduction)

Since reduction occurs at the cathode, the cathode is $Au^+(aq) \mid Au(s)$.

To determine the cell type, calculate \mathcal{E}° cell:

$$\begin{split} \mathrm{Mg}(\mathrm{s}) &\to \mathrm{Mg}^{2+} + 2\,e^{-} \\ &\mathcal{E}_{\mathrm{anode}}^{\circ} = +2.36~\mathrm{V} \\ 2\,\mathrm{Au}^{+} + 2\,e^{-} &\to 2\,\mathrm{Au}(\mathrm{s}) \\ &\mathcal{E}_{\mathrm{cathode}}^{\circ} = +1.69~\mathrm{V} \\ 2\,\mathrm{Au}^{+} + \mathrm{Mg}(\mathrm{s}) &\to 2\,\mathrm{Au}(\mathrm{s}) + \mathrm{Mg}^{2+} \\ &\mathcal{E}_{\mathrm{call}}^{0} = 4.05\,V \end{split}$$

Since \mathcal{E}_{cell}^0 is positive, the reaction is spontaneous, so this is a voltaic cell (a battery).

022 10.0 points

Consider the half-reactions	
$\mathrm{Mn}^{2+} + 2 \ e^- \to \mathrm{Mn}$	$E^0 = -1.029 \text{ V}$
$Ga^{3+} + 3 e^- \rightarrow Ga$	$E^0 = -0.560 \text{ V}$
$\mathrm{Fe}^{2+} + 2 \ e^- \to \mathrm{Fe}$	$E^0 = -0.409 \text{ V}$
$\operatorname{Sn}^{2+} + 2e^- \to \operatorname{Sn}$	$E^0 = -0.136 \text{ V}$
Using the redox couples	to establish a

voltaic cell, which reaction would be non-spontaneous?

1. $2 \operatorname{Ga} + 3 \operatorname{Sn}^{2+} \rightarrow 2 \operatorname{Ga}^{3+} + 3 \operatorname{Sn}$

2.
$$\operatorname{Fe}^{2+} + \operatorname{Mn} \to \operatorname{Mn}^{2+} + \operatorname{Fe}$$

3.
$$\operatorname{Sn}^{2+} + \operatorname{Mn} \to \operatorname{Sn} + \operatorname{Mn}^{2+}$$

4. $2 \operatorname{Ga}^{3+} + 3 \operatorname{Fe} \rightarrow 2 \operatorname{Ga} + 3 \operatorname{Fe}^{2+}$ correct

5. $\operatorname{Sn}^{2+} + \operatorname{Fe} \to \operatorname{Sn} + \operatorname{Fe}^{2+}$

Explanation:

Only $2 \operatorname{Ga}^{3+} + 3 \operatorname{Fe} \to 2 \operatorname{Ga} + 3 \operatorname{Fe}^{2+}$ has a negative E_{cell}^0 , so it is the only reaction here that is nonspontaneous.

023 10.0 points

Find the standard emf of the given cell

$$\operatorname{Cu}(s) | \operatorname{Cu}^{2+}(\operatorname{aq}) || \operatorname{Au}^{+}(\operatorname{aq}) | \operatorname{Au}(s)$$

1.-2.03 V

- $\mathbf{2.} + 2.03 \text{ V correct}$
- **3.** −1.35 V
- **4.** -0.91 V

5. +0.91 V

6. +1.35 V

Explanation:

Identify the cathode (right-side) and anode (left-side) reactions and potentials from the cell diagram.

At the cathode, $Au^+(aq) + e^- \rightarrow Au(s) \qquad E^\circ = +1.69 V$ At the anode, $\operatorname{Cu}(\mathrm{s}) \to \operatorname{Cu}^{2+}(\mathrm{aq}) + 2 e^{-} \qquad -E^{\circ} = -0.34 \text{ V}$

$$E_{\text{cell}}^{\circ} = E_{\text{cathode}}^{\circ} - E_{\text{anode}}^{\circ}$$
$$= +1.69 \text{ V} - (+0.34 \text{ V})$$
$$= +1.35 \text{ V}$$

024 10.0 points

Which species will reduce Ag^+ but not Fe^{2+} ?

H₂ correct
V
Au
Cr
Pt
Explanation:

025 10.0 points

If the table of standard reduction potentials is ordered with the strongest reducing agents at the top, how are the reduction potentials ordered (from top to bottom)?

1. From most common to least common

2. From most spontaneous to least spontaneous

3. From most positive to most negative

4. From most negative to most positive correct

Explanation:

Good reducing agents have positive (spontaneous) oxidation potentials. The reduction potentials is the opposite of the oxidation potential, so the oxidized form of a good reducing agent will have a negative (nonspontaneous) reduction potential.

026 10.0 points

Which specie is the weakest reducing agent in the table of half reactions?

1. F^- correct

2. Li

3. F₂

4. Li^+

Explanation:

For a molecule to be a good reducing agent, it must have a large, positive oxidation potential (*i.e.*, it is spontaneously oxidized).

027 10.0 points

If the two half reactions below were used to make an electrolytic cell, what species would be consumed at the anode?

Half reaction	E°
$\operatorname{Au}^{3+}(\operatorname{aq}) + 3 e^{-} \longrightarrow \operatorname{Au}(\operatorname{s})$	+1.50
$I_2(s) + 2 e^- \longrightarrow 2 I^-(aq)$	+0.53

1. Au³⁺(aq)

2. I⁻(aq)

3. Au(s) correct

4. $I_2(s)$

Explanation:

An electrolytic cell must have a negative standard cell potential and therefore the anodic reaction must consume Au(s).