This print-out should have 21 questions. Multiple-choice questions may continue on the next column or page – find all choices before answering.

001 10.0 points

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

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

What is the proper cell diagram for this reaction?

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

2. $\mathrm{Mn(s)\,|\,Mn^{2+}(aq)\,||\,Ti^{2+}(aq)\,|\,Ti(s)}$ **correct**

3. $\mathrm{Mn^{2+}(aq)\,|\,Mn(s)\,||\,Ti(s)\,|\,Ti^{2+}(aq)}$

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

Explanation:

The two half-reactions, written as reductions, are

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

Equate e^-:

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

Add the balanced half reactions:

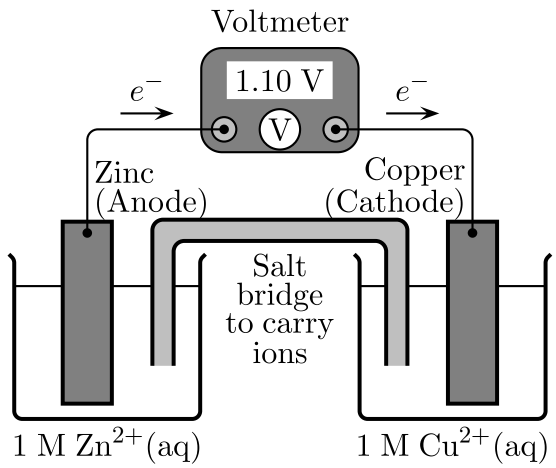
$$\mathrm{Mn(s)} + \mathrm{Ti^{2+}(aq)} \rightarrow \mathrm{Mn^{2+}(aq)} + \mathrm{Ti(s)}$$

To write the cell diagram we place the species involved with the reduction reaction (at the cathode) on the right of the salt bridge and those involved with the oxidation reaction (at the anode) on the left of the bridge. Electrons are not shown, and the extreme left and right species listed are the electrode materials the anode and cathode respectively, are made of. As neither reaction involves solely gases or aqueous ions, we can use the elemental metals involved in the half reactions.

The cell diagram is

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

002 10.0 points



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

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

2. $\mathrm{Cu^{2+}(aq)} + 2\,e^- \rightarrow \mathrm{Cu(s)}$ **correct**

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

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

Explanation:

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

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

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

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

003 10.0 points

What is the standard cell potential of a battery made from the half reactions

$2\,\mathrm{H^+} + 2\,e^- \longrightarrow \mathrm{H_2}$ $\quad E^\circ = 0.00$ V

$\mathrm{O_2} + 4\,\mathrm{H^+} + 4\,e^- \longrightarrow 2\,\mathrm{H_2O}$ $\quad E^\circ = +1.23$ V

1. 2.46

2. 1.23 **correct**

3. -1.23

4. -2.46

Explanation:

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= 1.23 - 0.00 = 1.23 \end{aligned}$$

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004 10.0 points

In a galvanic cell,

1. oxidation takes place at the cathode.

2. electrolytes are added to carry electrons between electrodes.

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

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

Explanation:

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

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005 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 oxidizing agent.

2. is reduced and is the reducing agent.

3. is reduced and is the oxidizing agent.

4. is oxidized and is the reducing agent. **correct**

Explanation:

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

$$Ag^+ + Cu \rightarrow Cu^+ + Ag$$

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.

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006 10.0 points

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

$$\text{Fe} \mid \text{Fe}^{2+}(1.0\text{ M}) \parallel \text{Zn}^{2+}(1.0\text{ M}) \mid \text{Zn}$$

$E^0(\text{Zn}) = -0.76$ $E^0(\text{Fe}) = -0.44$

1. -1.20

2. -0.32 **correct**

3. $+0.32$

4. $+1.20$

Explanation:

$$\begin{aligned} E_{\text{cell}} &= E^0_{(\text{cathode})} - E^0_{(\text{anode})} \\ &= -0.76 - (-0.44) \\ &= -0.32 \end{aligned}$$

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007 10.0 points

Which of the metals in the list below will react with 1 M H_2SO_4 to produce hydrogen gas?

	E^0 (volts)
$Na^+ + 1\,e^- \rightarrow Na$	-2.714
$Cd^{2+} + 2\,e^- \rightarrow Cd$	-0.403
$Pb^{2+} + 2\,e^- \rightarrow Pb$	-0.126
$Cu^{2+} + 2\,e^- \rightarrow Cu$	$+0.337$

1. Na, Cd, and Pb only **correct**

2. Na, Cd, Pb, and Cu

3. Na and Cd only

4. some other combination than those listed

5. Na only

Explanation:

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008 10.0 points

Consider the voltaic cell:

Pt | Sn^{2+} (0.10 M), Sn^{4+} (0.0010 M)
|| Ag^{+} (0.010 M) | Ag

$Sn^{4+} + 2\,e^{-} \rightarrow Sn^{2+}$ $E^0 = +0.15$ V

$Ag^{+} + 1\,e^{-} \rightarrow Ag(s)$ $E^0 = +0.80$ V

The electrons flow in the external circuit from

1. Ag to Pt.

2. Ag to Sn^{4+}.

3. Sn^{2+} to Ag^{+}.

4. Sn to Ag.

5. Pt to Ag. **correct**

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.

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009 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. -3 V

2. 6 V **correct**

3. -6 V

4. 3 V

Explanation:

Using the tables, the largest values of E°_{red} are about + 3 V and - 3 V.

The species with the positive value would be reduced, the other would be oxidised so:

E°_{cell} = + 3 V - (- 3 V) = + 6 V

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010 10.0 points

Consider the cell

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

Calculate E°.

1. + 0.54 V

2. + 1.20 V

3. + 0.98 V **correct**

4. $-$ 0.54 V

5. $-$ 1.20 V

Explanation:

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011 10.0 points

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

1. Zn^{2+}

2. Pb^{4+}

3. V^{3+} **correct**

4. Fe^{2+}

5. O_3 in acidic medium

Explanation:

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012 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. Ag

2. Ag^{+}

3. Fe^{2+}

4. Cu

5. Fe **correct**

Explanation:

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013 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:

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014 10.0 points

For the cell diagram

$$\mathrm{Cd(s)\,|\,CdSO_4(aq)\,||\,Hg_2SO_4\,|\,Hg(\ell)}$$

what reaction occurs at the cathode?

1. $\mathrm{Hg_2SO_4(s)} + 2\,e^- \rightarrow 2\,\mathrm{Hg}(\ell) + \mathrm{SO_4^{2-}(aq)}$ **correct**

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

Explanation:

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015 10.0 points

What is the cathode in

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

$\mathrm{Mg^{2+}} + 2\,e^- \rightarrow \mathrm{Mg}$ $\quad \mathcal{E}^\circ_{\rm red} = -2.36$

$\mathrm{Au^+} + e^- \rightarrow \mathrm{Au}$ $\quad \mathcal{E}^\circ_{\rm red} = +1.69$

and what is the cell type?

1. Not enough information is provided.

2. Mg(s); an electrolytic cell

3. Mg(s); a voltaic cell

4. Au(s); an electrolytic cell

5. Au(s); a voltaic cell **correct**

Explanation:

The diagram A | B || C | 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 $\mathrm{Au^+(aq)\ |\ Au(s)}$.

To determine the cell type, calculate $\mathcal{E}^\circ$ cell:

$$\mathrm{Mg(s)} \rightarrow \mathrm{Mg^{2+}} + 2\,e^- \qquad \mathcal{E}^\circ_{\rm anode} = +2.36\ \mathrm{V}$$

$$2\,\mathrm{Au^+} + 2\,e^- \rightarrow 2\,\mathrm{Au(s)} \qquad \mathcal{E}^\circ_{\rm cathode} = +1.69\ \mathrm{V}$$

$$2\,\mathrm{Au^+} + \mathrm{Mg(s)} \rightarrow 2\,\mathrm{Au(s)} + \mathrm{Mg^{2+}} \qquad \mathcal{E}^0_{\rm cell} = 4.05\,V$$

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

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016 10.0 points

Consider the half-reactions

$\mathrm{Mn^{2+}} + 2\,e^- \rightarrow \mathrm{Mn}$ $\quad E^0 = -1.029$ V

$\mathrm{Ga^{3+}} + 3\,e^- \rightarrow \mathrm{Ga}$ $\quad E^0 = -0.560$ V

$\mathrm{Fe^{2+}} + 2\,e^- \rightarrow \mathrm{Fe}$ $\quad E^0 = -0.409$ V

$\mathrm{Sn^{2+}} + 2\,e^- \rightarrow \mathrm{Sn}$ $\quad E^0 = -0.136$ V

Using the redox couples to establish a voltaic cell, which reaction would be nonspontaneous?

1. $\mathrm{Fe^{2+} + Mn \rightarrow Mn^{2+} + Fe}$

2. $\mathrm{2\,Ga^{3+} + 3\,Fe \rightarrow 2\,Ga + 3\,Fe^{2+}}$ **correct**

3. $\mathrm{Sn^{2+} + Fe \rightarrow Sn + Fe^{2+}}$

4. $\mathrm{2\,Ga + 3\,Sn^{2+} \rightarrow 2\,Ga^{3+} + 3\,Sn}$

5. $\mathrm{Sn^{2+} + Mn \rightarrow Sn + Mn^{2+}}$

Explanation:

Only $\mathrm{2\,Ga^{3+} + 3\,Fe \rightarrow 2\,Ga + 3\,Fe^{2+}}$ has a negative $E^0_{\rm cell}$, so it is the only reaction here that is nonspontaneous.

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017 10.0 points

Find the standard emf of the given cell

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

1. -0.91 V

2. -1.35 V

3. $+1.35$ V **correct**

4. -2.03 V

5. $+2.03$ V

6. $+0.91$ 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,

$Cu(s) \rightarrow Cu^{2+}(aq) + 2\,e^- \qquad -E^\circ = -0.34$ V

$$\begin{aligned} E^\circ_{\text{cell}} &= E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} \\ &= +1.69 \text{ V} - (+0.34 \text{ V}) \\ &= +1.35 \text{ V} \end{aligned}$$

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018 10.0 points

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

1. Pt

2. Cr

3. Au

4. H_2 **correct**

5. V

Explanation:

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019 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 positive to most negative

2. From most common to least common

3. From most spontaneous to least spontaneous

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.

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020 10.0 points

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

1. Li^+

2. F_2

3. F^- **correct**

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).

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021 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°
$Au^{3+}(aq) + 3\,e^- \longrightarrow Au(s)$	$+1.50$
$I_2(s) + 2\,e^- \longrightarrow 2\,I^-(aq)$	$+0.53$

1. $I^-(aq)$

2. Au(s) **correct**

3. $I_2(s)$

4. $Au^{3+}(aq)$

Explanation:

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