Periodic Table of the Elements
001

Predict the standard emf of the given cell
C(gr) | Sn^{4+}(aq), Sn^{2+}(aq) || Pb^{4+}(aq), Pb^{2+} | Pt(s)

1. +1.52 V correct
2. +0.75 V
3. +0.37 V
4. +0.52 V
5. +2.14 V
6. +1.34 V

Explanation:
Identify the cathode (right-side) and anode (left-side) reactions and potentials from the cell diagram.
At the cathode,
Pb^{4+}(aq) + 2 e^- \rightarrow Pb^{2+}(aq) \quad E^\circ = +1.67 \text{ V}
At the anode,
Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2 e^- -E^\circ = -0.15 \text{ V}

\[ E^\circ_{\text{cell}} = E^\circ_{\text{cathode}} - E^\circ_{\text{anode}} = +1.67 \text{ V} - (-0.15 \text{ V}) = +1.52 \text{ V} \]

003

Consider the cell
Zn(s) | Zn^{2+}(aq) || Fe^{2+}(aq) | Fe(s)
at standard conditions.
Calculate the value of \( \Delta G^\circ \) for the reaction that occurs when current is drawn from this cell.

1. – 62 kJ \cdot \text{mol}^{-1} correct
2. – 230 kJ \cdot \text{mol}^{-1}
3. + 62 kJ \cdot \text{mol}^{-1}
4. + 230 kJ \cdot \text{mol}^{-1}
5. – 31 kJ \cdot \text{mol}^{-1}

Explanation:
Of the species listed, the strongest oxidizing agent is
1. Sn^{2+} correct
2. Mn^{2+}
3. Mn
4. Sn
5. Ga^{3+}
As the battery discharges, electrons flow from the ? terminal to the ? terminal through the external circuit and ? reaction occurs at the positive terminal.

1. positive; negative; a reduction
2. positive; negative; an oxidation
3. negative; positive; a reduction correct
4. negative; positive; an oxidation
5. positive; negative; an acid/base

**Explanation:**
In a voltaic cell electrons flow from the negative to the positive terminals. Reduction occurs at the positive terminal.

Which of the following batteries could not be recharged?

1. dry cell correct
2. lead storage battery
3. nickel-cadmium battery

**Explanation:**

Calculate the potential for the cell indicated:

| Fe | Fe^{2+} (10^{-3} M) || Pb^{2+} (10^{-5} M) | Pb |
| Pb^{2+} + 2 e^- → Pb | E^0 = -0.126 V |
| Fe^{2+} + 2 e^- → Fe | E^0 = -0.440 V |

1. 0.255 V correct
2. 0.432 V
3. 0.373 V
4. 0.196 V

**Explanation:**

The overall reaction is

Fe + Pb^{2+} → Fe^{2+} + Pb

Please notice that since the concentrations are not 1 M, the Nernst equation must be used.
In this cell notation, the anode is located on the left of the salt bridge || and the cathode on the right. So first calculate

\[ E_{\text{cell}}^0 = E_{\text{cathode}} - E_{\text{anode}}^0 \]

\[ = -0.126 \text{ V} - (-0.440) \text{ V} = 0.314 \text{ V} \]

Using the Nernst Equation

\[ E_{\text{cell}} = E_{\text{cell}}^0 - \frac{0.05916}{n} \log Q \]

\[ = 0.314 \text{ V} - \frac{0.05916}{2} \log \left( \frac{[\text{Fe}^{2+}]}{[\text{Pb}^{2+}]} \right) \]

\[ = 0.314 \text{ V} - \frac{0.05916}{2} \log \left( \frac{10^{-3}}{10^{-5}} \right) \]

\[ = 0.25484 \text{ V} \]

What weight of Cl₂ gas will be produced by electrolysis of molten NaCl when a current of 4.35 amps flows through it for 15.0 hours? (Cl = 35.457 g/mol)

1. 86.3 g correct
2. 19.8 g
3. 0.0250 g
4. 1.44 g
5. 43.2 g

**Explanation:**
Using the smallest possible integer coefficients to balance the redox equation

\[ \text{MnO}_4^- + C_2O_4^{2-} \rightarrow \text{Mn}^{2+} + \text{CO}_2 \]  

(acidic solution), the coefficient for \( C_2O_4^{2-} \) is

1. 5. **Correct**

2. 2.

3. 4.

4. 7.

5. The correct coefficient is not given.

**Explanation:**

The oxidation number of C changes from +3 to +4, so C is oxidized. The oxidation number of Mn changes from +7 to +2, so Mn is reduced. We set up oxidation and reduction half-reactions:

Red: \( \text{MnO}_4^- \rightarrow \text{Mn}^{2+} \)

Oxid: \( C_2O_4^{2-} \rightarrow \text{CO}_2 \)

Mn atoms are balanced. We need 2 \( \text{CO}_2 \) molecules to balance C:

Oxid: \( C_2O_4^{2-} \rightarrow 2\text{CO}_2 \)

Since this is an acidic solution, we use \( \text{H}_2\text{O} \) and \( \text{H}^+ \) to balance O and H atoms, adding the \( \text{H}_2\text{O} \) to the side needing oxygen:

Red: \( 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \)

Oxid: \( C_2O_4^{2-} \rightarrow 2\text{CO}_2 \)

We balance the total charge in each half-reaction by adding electrons. In the preceding reduction reaction there is a total charge of +7 on the left and +2 on the right. Five electrons are added to the left:

Red: \( 5e^- + 8\text{H}^+ + \text{MnO}_4^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O} \)

Oxid: \( C_2O_4^{2-} \rightarrow 2\text{CO}_2 + 2e^- \)

The number of electrons gained by Mn must equal the number of electrons lost by C. We multiply the reduction reaction by 2 and the oxidation reaction by 5 to balance the electrons:

Red: \( 10e^- + 16\text{H}^+ + 2\text{MnO}_4^- \rightarrow 2\text{Mn}^{2+} + 8\text{H}_2\text{O} \)

Oxid: \( 5C_2O_4^{2-} \rightarrow 10\text{CO}_2 + 10e^- \)

Adding the half-reactions gives the overall balanced equation:

\[ 5C_2O_4^{2-} + 16\text{H}^+ + 2\text{MnO}_4^- \rightarrow 10\text{CO}_2 + 2\text{Mn}^{2+} + 8\text{H}_2\text{O} \]