1. At the cathode (negative electrode) in electrolysis:

   ☐ positive ions gain electrons (reduction)

   ☐ water is made

   ☐ negative ions lose electrons

   ☐ nothing happens

2. The standard hydrogen electrode is defined as having a potential of:

   ☐ 0.00 V

   ☐ 1.00 V

   ☐ 0.059 V

   ☐ −1.00 V

3. A redox reaction is feasible when the cell potential is:

   ☐ equal to 0.059

   ☐ negative

   ☐ zero only

   ☐ positive

4. Electrolysing an aqueous solution of a reactive metal salt gives, at the cathode:

   ☐ the halogen

   ☐ hydrogen

   ☐ oxygen

   ☐ the reactive metal

5. The standard cell potential is:

   ☐ E°(more positive) − E°(less positive)

   ☐ the sum of the two E° values

   ☐ always zero

   ☐ E°(less positive) − E°(more positive)

6. A reaction is feasible if the cell potential (Ecell) is positive.

   ☐ True ☐ False

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8. A more positive standard electrode potential means a stronger reducing agent.

   ☐ True ☐ False

9. $\mathrm{Ag^+ + e^- \rightleftharpoons Ag}$, $E^\ominus = +0.80$ V. Use $E = E^\ominus + \dfrac{0.059}{z}\log[\mathrm{Ag^+}]$ to find E, in V, when $[\mathrm{Ag^+}] = 0.010\ \dfrac{\mathrm{mol}}{\mathrm{dm^3}}$.

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10. A current of 1.00 A flows for 965 s during the electrolysis of dilute sulfuric acid: $\mathrm{2H^+ + 2e^- \rightarrow H_2}$. What volume of hydrogen forms, in $\mathrm{cm^3}$ at room conditions? (one mole of gas = 24 000 $\mathrm{cm^3}$; $F = 96\,500$)

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1. **positive ions gain electrons (reduction)**

   Positive ions are attracted to the cathode and gain electrons (reduction); the anode does oxidation.

2. **0.00 V**

   It is the reference half-cell, set at exactly 0.00 V; all other E° values are measured against it.

3. **positive**

   A positive E°cell means the reaction can happen (analogous to a negative ΔG).

4. **hydrogen**

   For a reactive metal, water is reduced to hydrogen at the cathode instead of the metal.

5. **E°(more positive) − E°(less positive)**

   Subtract the less positive electrode potential from the more positive one.

6. **True**

   Ecell = E(reduced) − E(oxidised).

7. **cathode**

   Reduction occurs at the cathode.

8. **False**

   False. A more negative E° means a stronger reducing agent, because that species loses electrons more readily. A more positive E° means a stronger oxidising agent.

9. **0.682**

   $z = 1$ and $\log 0.010 = -2$, so $E = 0.80 + 0.059 \times (-2) = +0.68$ V: diluting the oxidised form makes E less positive.

10. **120**

    $Q = 965$ C, so 0.0100 mol of electrons and 0.00500 mol of H_2 (two electrons each); $0.00500 \times 24\,000 = 120$ cm^3.

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