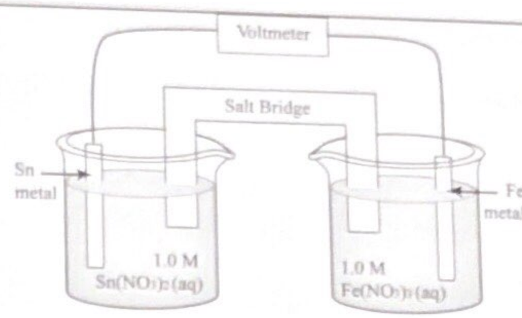
**SQ-2** A spontaneous electrochemical cell is set up as shown. Which statement is true?

| | |
|---|---|
| $Sn^{2+} + 2e^- \rightarrow Sn$ | $E° = -0.136$ V |
| $Fe^{3+} + 3e^- \rightarrow Fe$ | $E° = -0.036$ V |




**(A)** The tin electrode is the cathode and electrons move from left to right.

**(B)** The tin electrode is the cathode and electrons move from right to left.

**(C)** The tin electrode is the anode and electrons move from left to right.

**(D)** The tin electrode is the anode and electrons move from right to left.

***Knowledge Required:*** (1) Definition of spontaneous reaction. (2) Definition of anode and cathode.

***Thinking it Through:*** You are being asked to evaluate some statements about a spontaneous electrochemical cell. You begin by using the given cell potentials to determine the spontaneous reaction. Recall that for a reaction to be spontaneous the value of $E°_{cell}$ is positive because reactions with positive cell potentials have negative free energy changes, and are spontaneous. To get the spontaneous reaction you need to reverse the reaction involving Sn. You also have to multiply the half-reactions so that electrons cancel (here 6 electrons are being transferred as written). Remember that the oxidation half-reaction provides the electrons for the reduction half-reaction. You also remember to **NOT** multiply the values of $E°$ since these cell potentials are *intensive* properties and so their value does not depend on the number of times the half-reaction occurs.

| | | |
|---|---|---|
| $Sn \rightarrow Sn^{2+} + 2e^-$ | $E° = 0.136$ V | oxidation |
| $Fe^{3+} + 3e^- \rightarrow Fe$ | $E° = -0.036$ V | reduction |

| | |
|---|---|
| $3(Sn \rightarrow Sn^{2+} + 2e^-)$ | $E° = 0.136$ V |
| $2(Fe^{3+} + 3e^- \rightarrow Fe)$ | $E° = -0.036$ V |
| $3Sn + 2Fe^{3+} \rightarrow 3Sn^{2+} + 2Fe$ | $E° = 0.100$ V |

You also know that the oxidation reaction occurs at the anode and the reduction occurs at the cathode. For the spontaneous reaction, the anode is the Sn and the Fe is the cathode. You know that electrons will move from where they are produced (the oxidation half-reaction or the reaction at anode) to where they are consumed (the reduction half-reaction or the reaction at the cathode). The source of electrons for the reduction is the anode. Therefore, the correct choice is **(C)**.

Choices **(A)** and **(B)** are not correct because tin is the anode.

Choice **(D)** is not correct because the electrons move from the anode to the cathode.

***Practice Questions Related to This:*** **PQ-4, PQ-5, PQ-6**

**SQ-3.** Given the standard reduction potentials at 25 °C, what is the standard cell potential of a $Ni(s) \mid Ni^{2+}(aq) \parallel Ag^+(aq) \mid Ag(s)$ galvanic cell?

| Half Reaction | $E°$ |
|---|---|
| $Ag^+(aq) + e^- \rightarrow Ag(s)$ | 0.79 V |
| $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ | –0.23 V |

**(A)** 0.56 V **(B)** 1.02 V **(C)** 1.35 V **(D)** 1.81 V

*Knowledge Required:* (1) Ability to interpret electrochemical cell notation. (2) Definition of a galvanic cell.

*Thinking it Through:* You are told the cell is galvanic and you remember that galvanic cells are spontaneous reactions. For a reaction to be spontaneous under standard conditions the value of $\Delta G°$ must be negative. You also recall that when $\Delta G°$ is negative, then $E°_{cell}$ positive. You are given the reaction in electrochemical cell notation. You know that in this notation, the first pair represents the reaction at the anode (oxidation) and the second pair is the reaction at the cathode (reduction). The double lines ("||") represent the salt bridge.

oxidation half-reaction||reduction half-reaction

oxidation reactant | oxidation product || reduction reactant | reduction product

Converting the given notation $Ni(s) \mid Ni^{2+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$ to half-reactions you write:

$Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$ $E° = 0.23$ V (*Note* you changed the sign of $E°$ from the reduction in the table)

$Ag^+(aq) + e^- \rightarrow Ag(s)$ $E° = 0.79$ V

Multiplying the $Ag^+$ half-reaction by 2 to make the electrons cancel, and remembering to **NOT** multiply the values of the cell potentials, since they are intensive properties, you get.

$Ni(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Ni^{2+}(aq)$ $E° = 1.02$ V which is choice **(B)**.

Choice **(A)** is not correct because the half-reaction for Ni was not reversed. Choice **(C)** is not correct because the half-reaction reaction for silver was multiplied by 2 and the Ni reaction was not reversed. Choice **(D)** is not correct because the half-reaction reaction for silver was multiplied by 2 and the Ni reaction was reversed.

*Practice Questions Related to This:* **PQ-7, PQ-8, PQ-9, PQ-10, PQ-11**

**SQ-4.** An oxidation-reduction reaction in which 3 electrons are transferred has a $\Delta G° = 18.55$ kJ·mol$^{-1}$ at 25°C. What is the value of $E°$?

**(A)** 0.192 V **(B)** –0.064 V **(C)** –0.192 V **(D)** –0.577 V

*Knowledge Required:* (1) The relationship of free energy to cell potential.

*Thinking it Through:* You are given a value of $\Delta G°$ and are asked for the cell potential. You recall the relationship between $\Delta G°$ and $E°$: $\Delta G° = -nFE°$

[**Note**: It isn't needed in this problem but the value of $E°$ is also related to the equilibrium constant.]

$$\Delta G° = -RT\ln K \quad \text{so} \quad -RT\ln K = -nFE°$$

$$E° = (RT/nF)\ln K$$

where $n$ is the number of electrons transferred and $F$ is the Faraday constant. *Note:* The value of the Faraday constant will typically be given on the reference page of the exam with other constants.

Solving for $E°$ you get: $E° = -\Delta G°/nF$

Substituting values results in:

$$E° = \frac{-(18.55\text{ kJ})\left(\dfrac{1000\text{ J}}{1\text{ kJ}}\right)}{(3\text{ mol electrons})\left(\dfrac{96485\text{ C}}{1\text{ mol electrons}}\right)} = -0.0642\text{ J}\cdot\text{C}^{-1} = -0.0642\text{ V}$$

Where you have used the definition 1 V = 1 J·C$^{-1}$. The correct choice is **(B)**.

Choice **(A)** is not correct because it omitted the number of electrons and dropped the negative sign.

Choice **(C)** is not correct because it omitted the number of electrons.

Choice **(D)** is not correct because it multiplied by the number of electrons rather than dividing.

*Practice Questions Related to This:* **PQ-12, PQ-13**

**SQ-5.** Conceptual

Consider the reaction.

$$Cu^{2+}(aq) + Fe(s) \rightarrow Cu(s) + Fe^{2+}(aq) \qquad E° = 0.78\text{ V}$$

What is the value of $E$ when $[Cu^{2+}]$ is equal to 0.040 M and $[Fe^{2+}]$ is equal to 0.40 M?

**(A)** 0.72 V **(B)** 0.75 V **(C)** 0.81 V **(D)** 0.84 V

***Knowledge Required:*** (1) Use the Nernst equation to calculate cell potential under nonstandard conditions.

***Thinking it Through:*** You recall the relationship between $E$ and $E°$ is given by the Nernst equation. [**Note**: the equation will be provided either in the problem or on the reference sheet.] There are two common versions of the Nernst equation:

$$E = E° - \frac{RT}{nF}\ln Q \qquad E = E° - \frac{0.0592}{n}\log Q$$

The 0.0592 comes from using the values for $R$, $F$, and 25 °C (298 K) and the conversion between the natural and base-10 log. You then use the reaction to write an expression for $Q$; recall from chapter 11 that Q is an expression written exactly like an equilibrium expression, but the values substituted into the expression are not necessarily equilibrium values. Once the expression is written and the appropriate values are used you calculate Q.

$$Q = \frac{\text{Products}}{\text{Reactants}} = \frac{\left[Fe^{2+}\right]}{\left[Cu^{2+}\right]} = \frac{0.40\text{ M}}{0.040\text{ M}} = 10$$

Now, using the value of $n = 2$ because 2 mole of electrons are transferred for every mole of reaction:

$$E = 0.78\text{ V} - \frac{0.0592}{2}\log(10) = 0.75\text{ V} \quad \text{which is choice } \mathbf{(B)}.$$

Choice **(A)** is incorrect because it omits $n$. Choice **(C)** is incorrect because it added the nonstandard cell potential term, $(0.0592/2)\log Q$. Choice **(D)** is incorrect because it added the nonstandard cell potential term and didn't use $n = 2$.

***Practice Questions Related to This:*** **PQ-14, PQ-15, PQ-16, PQ-17, PQ-18, PQ-19, PQ-20**

**SQ-6.** What half-reaction occurs at the anode during the electrolysis of molten sodium iodide?

**(A)** $2I^- \rightarrow I_2 + 2e^-$ **(B)** $I_2 + 2e^- \rightarrow 2I^-$

**(C)** $Na \rightarrow Na^+ + e^-$ **(D)** $Na^+ + e^- \rightarrow Na$

***Knowledge Required:*** (1) The definition of electrolysis. (2) Definitions of anode and cathode in electrochemical cells.

***Thinking it Through:*** You are being asked to identify the reaction occurring at the anode of an electrolytic cell. You remember that electrolysis reactions are reactions that use electrical energy to drive a chemical process. Specifically, these reactions have a positive $\Delta G$ and can be driven by the application of electrical energy. This is different from the spontaneous reactions used in galvanic cells, which have negative $\Delta G$ values and the excess free energy can be used to do electrical work. You also remember that no matter if the cell is a galvanic or electrolytic cell, oxidation occurs at the anode and reduction occurs at the cathode.

Because the electrolysis involves molten sodium iodide, you do not have to consider the oxidation or reduction of water. Remember that in aqueous solutions, these reactions are often the ones that occur.

Recall that in molten salts, the species of interest are the ions. For molten sodium iodide the ions are: $Na^+$ and $I^-$. These are the only options for reactants for the half-reactions. Because the process at the anode will be oxidation, you eliminate $Na^+$ because it is in its highest oxidation state. Therefore, the only possible reaction is the oxidation of $I^-$: $2I^- \rightarrow I_2 + 2e^-$ which is choice **(A)**.

Choice **(B)** is not correct because this is the reduction of $I_2$. Choice **(C)** is not correct because it is the oxidation of Na, a species not present in the molten salt. Choice **(D)** is not correct because it is the reduction of $Na^+$.

***Practice Questions Related to This:*** **PQ-21, PQ-22, PQ-23**

**SQ-7.** How many minutes are required to plate 2.08 g of copper at a constant current flow of 1.26 A?

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

| Constants | |
|---|---|
| Molar mass, Cu | 63.5 g·mol$^{-1}$ |
| 1 faraday ($F$) | 96485 C |

**(A)** 41.8 min **(B)** 83.6 min

**(C)** 133 min **(D)** 5016 min

***Knowledge Required:*** (1) The stoichiometry of electrolytic processes. (2) The relationships among amperes, faradays, time, and moles of electrons.

***Thinking it Through:*** You are given a mass and current. You recall that an Ampere is defined as 1 C of charge per second: $1\text{ A} = \frac{1\text{ C}}{1\text{ s}}$. Therefore, the 1.26 A can be written as 1.26 C·s$^{-1}$ or as a conversion factor; 1.26 C = 1 s.

The stoichiometric relationship involving electrons and copper is: 2 mol e$^-$ = 1 mol Cu
The Faraday is also a conversation between coulomb and moles of electrons: 1 mol e$^-$ = 96485 C
You can use these relationships to find the time required. You start with the mass of copper desired:

$$2.08\text{ g Cu}\left(\frac{1\text{ mol Cu}}{63.5\text{ g Cu}}\right)\left(\frac{2\text{ mol e}^-}{1\text{ mol Cu}}\right)\left(\frac{96485\text{ C}}{1\text{ mol e}^-}\right)\left(\frac{1\text{ s}}{1.26\text{ C}}\right)\left(\frac{1\text{ min}}{60\text{ s}}\right) = 83.6\text{ min}$$ which is choice **(B)**.

Choice **(A)** is incorrect because it didn't use the 2 mol of electrons per 1 mol Cu relationship. Choice **(C)** is incorrect because it incorrectly multiplied by 1.26. Choice **(D)** is incorrect because it is the number of seconds, and did not convert the time to minutes.

***Practice Questions Related to This:*** **PQ-24, PQ-25, PQ-26, PQ-27, PQ-28**

**SQ-8.** For the concentration cell

$$Fe(s)\,|\,Fe^{3+}(aq),\,0.100\text{ M}\,||\,Fe^{3+}(aq),\,1.00\text{ M}\,|\,Fe(s)$$

what is $E_{cell}$ at 25 °C (in mV)?

**(A)** 9.87 mV **(B)** 19.7 mV **(C)** 29.6 mV **(D)** 59.2 mV

***Knowledge Required:*** (1) Definition of a concentration cell. (2) Ability to interpret electrochemical cell notation.

***Thinking it Through:*** You are being asked about a concentration cell. You recall that in a concentration cell the source of the electrical potential, and the current flow, is the differences in the concentrations of the species. The given cell notation can be converted to half-reactions:

$$Fe(s)\,|\,Fe^{3+}(aq),\,0.100\text{ M}\,||\,Fe^{3+}(aq),\,1.00\text{ M}\,|\,Fe(s)$$

oxidation ↙ ↘ reduction

$Fe(s) \rightarrow Fe^{3+}(aq, 0.100\text{ M}) + 3e^-$ $Fe^{3+}(aq, 1.00\text{ M}) + 3e^- \rightarrow Fe(s)$

After adding the half-reactions, you get: $Fe(s) + Fe^{3+}(aq, 1.00\text{ M}) \rightarrow Fe^{3+}(aq, 0.100\text{ M}) + Fe(s)$

The Nernst equation takes into account the dependence of cell potential on concentrations:

$$E_{cell} = E^{\circ}_{cell} - \left(\frac{0.0592}{n}\right)\log Q$$ remembering that $Q$ is the ratio of the concentrations of products to reactants.

You also recall that the $E^{\circ}_{cell}$ for a concentration cell is zero, because the two half-reactions are the same:

$$E_{cell} = 0 - \left(\frac{0.0592}{3}\right)\log\left(\frac{0.100\text{ M}}{1.00\text{ M}}\right) = 0.0197\text{ V}$$

$$0.0197\text{ V}\left(\frac{1000\text{ mV}}{1\text{ V}}\right) = 19.7\text{ mV}$$

which is choice **(B)**.

Choice **(A)** is not correct because it used 6 for the number of electrons in the Nernst equation. Choice **(C)** is not correct because it used 2 for the number of electrons in the Nernst equation. Choice **(D)** is not correct because it did not use the number of electrons in the Nernst equation.

***Practice Questions Related to This:*** **PQ-29, PQ-30**

## Practice Questions (PQ)

Conceptual **PQ-1.** What is the coefficient of the hydroxide ion for the half reaction when the hypobromite ion reacts to form the bromide ion in basic solution? $BrO^-(aq) \rightarrow Br^-(aq)$ *unbalanced*

**(A)** 1 **(B)** 2 **(C)** 3 **(D)** 4

Conceptual **PQ-2.** What is the balanced half reaction of the nitrate ion reacting to form nitrogen monoxide in acidic solution?

**(A)** $NO_3^-(aq) + H_2O(l) \rightarrow NO(g) + 3e^- + 2H^+(aq)$

**(B)** $NO_3^-(aq) + 3e^- + 4H^+(aq) \rightarrow NO(g) + 2H_2O(l)$

**(C)** $NO_3^-(aq) + 3e^- + 2H_2O(l) \rightarrow NO(g) + 4OH^-(aq)$

**(D)** $NO_3^-(aq) + 4e^- + 5H^+(aq) \rightarrow NO(g) + 2H_2O(l)$

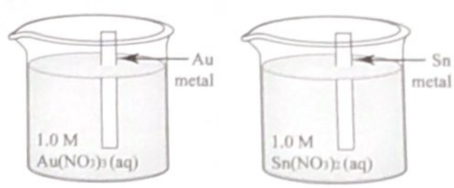
Conceptual **PQ-3.** What is the coefficient for water when the reaction is balanced in acidic solution?

$Cr_2O_7^{2-}(aq) + Mn^{2+}(aq) \rightarrow Cr^{3+}(aq) + MnO_2(s)$ *unbalanced*

**(A)** 1 **(B)** 2 **(C)** 4 **(D)** 5

**PQ-4.** Which reaction occurs at the cathode of a galvanic cell constructed from these two half-cells?

The standard reduction potential of gold(III) is +1.50 V.
The standard reduction potential of tin(II) is –0.14 V.




**(A)** $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

**(B)** $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

**(C)** $Au(s) \rightarrow Au^{3+}(aq) + 3e^-$

**(D)** $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

Conceptual **PQ-5.** Which reaction will occur if each substance is in its standard state?
Assume potentials are given in water at 25 °C.

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$ | –0.28 V |
| $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$ | 0.15 V |
| $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ | 1.06 V |

**(A)** $Ni^{2+}$ will oxidize $Sn^{2+}$ to give $Sn^{4+}$

**(B)** $Sn^{4+}$ will oxidize $Br^-$ to give $Br_2$

**(C)** $Br_2$ will oxidize Ni(s) to give $Ni^{2+}$

**(D)** $Ni^{2+}$ will oxidize $Br_2$ to give $Br^-$

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$ | –0.04 V |
| $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$ | 1.36 V |

**PQ-6.** Which represents a spontaneous reaction and what is the correct $E°_{cell}$?

| Reaction | $E°_{cell}$ |
|---|---|
| **(A)** $Fe(s) + Cl_2(g) \rightarrow Fe^{3+}(aq) + 2Cl^-(aq)$ | 1.40 V |
| **(B)** $Fe^{3+}(aq) + Cl_2(g) \rightarrow Fe(s) + 2Cl^-(aq)$ | 1.40 V |
| **(C)** $2Fe(s) + 3Cl_2(g) \rightarrow 2Fe^{3+}(aq) + 6Cl^-(aq)$ | 1.40 V |
| **(D)** $2Fe(s) + 3Cl_2(g) \rightarrow 2Fe^{3+}(aq) + 6Cl^-(aq)$ | 4.16 V |

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Cu^+(aq) + e^- \rightarrow Cu(s)$ | 0.52 V |
| $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$ | 0.15 V |
| $Cr^{3+}(aq) + e^- \rightarrow Cr^{2+}(aq)$ | –0.41 V |

**PQ-7.** Which combination of reactants will produce the greatest voltage based on these standard electrode potentials?

**(A)** $Cu^+$ and $Sn^{2+}$ **(B)** $Cu^+$ and $Cr^{2+}$ **(C)** Cu and $Sn^{2+}$ **(D)** $Sn^{4+}$ and $Cr^{2+}$

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Cr^{3+}(aq) + 3e^- \rightarrow Cr(s)$ | –0.74 V |
| $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$ | –0.28 V |

**PQ-8.** What is the standard electrode potential for a voltaic cell constructed in the appropriate way from these two half-cells?

**(A)** –1.02 V **(B)** 0.46 V **(C)** 0.64 V **(D)** 1.02 V

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Pd^{2+}(aq) + 2e^- \rightarrow Pd(s)$ | ? |
| $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ | 0.337 V |

**PQ-9.** What is the value of the missing standard reduction potential?
$Cu(s) + Pd^{2+}(aq) \rightarrow Cu^{2+}(aq) + Pd(s)$ $E° = 0.650$ V

**(A)** 0.987 V **(B)** –0.987 V **(C)** 0.313 V **(D)** –0.313 V

**PQ-10.** The standard cell potential, $E°$, for this reaction is 0.79 V.

$$6I^-(aq) + Cr_2O_7^{2-}(aq) + 14H^+(aq) \rightarrow 3I_2(aq) + 2Cr^{3+}(aq) + 7H_2O(l)$$

What is the standard potential for $I_2(aq)$ being reduced to $I^-(aq)$ given that the standard reduction potential for $Cr_2O_7^{2-}(aq)$ changing to $Cr^{3+}(aq)$ is 1.33 V?

**(A)** 0.54 V **(B)** –0.54 V **(C)** 0.18 V **(D)** –0.18 V

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$ | 1.61 V |
| $Br_2(l) + 2e^- \rightarrow 2Br^-(aq)$ | 1.06 V |

Conceptual **PQ-11.** What is the standard cell potential, $E°$, for this reaction?
$Br_2(l) + 2Ce^{3+}(aq) \rightarrow 2Br^-(aq) + 2Ce^{4+}(aq)$

**(A)** –2.67 V **(B)** –2.16 V **(C)** –0.55 V **(D)** 2.67 V

Conceptual **PQ-12.** Which statement is true for an electrochemical cell built from an oxidation-reduction reaction if $K$ for the reaction is greater than 1?

**(A)** $\Delta G°$ is negative, $E°$ is negative
**(B)** $\Delta G°$ is negative, $E°$ is positive
**(C)** $\Delta G°$ is positive, $E°$ is negative
**(D)** $\Delta G°$ is positive, $E°$ is positive

| Reduction Potentials | $E°$ |
|---|---|
| $Cu^{2+}$\|Cu | 0.34 V |
| $Ag^+$\|Ag | 0.80 V |

**PQ-13.** What is the equilibrium constant at 298 K for the spontaneous reaction that occurs when a $Cu^{2+}$|Cu half-cell is connected to a $Ag^+$|Ag half-cell?

**(A)** $2.8\times10^{-16}$ **(B)** $6.0\times10^{7}$ **(C)** $3.6\times10^{15}$ **(D)** $3.7\times10^{38}$

**PQ-14.** Consider the reaction at 298 K.

$$Sn^{2+}(aq) + 2Fe^{3+}(aq) \rightarrow Sn^{4+}(aq) + 2Fe^{2+}(aq) \qquad E° = 0.617\ V$$

What is the value of $E$ when $[Sn^{2+}]$ and $[Fe^{3+}]$ are equal to 0.50 M and $[Sn^{4+}]$ and $[Fe^{2+}]$ are equal to 0.10 M?

**(A)** 0.069 V **(B)** 0.679 V **(C)** 0.658 V **(D)** 0.576 V

*Conceptual* **PQ-15.** Consider this reaction.

$$2Al(s) + 6HCl(aq) \rightleftharpoons 2AlCl_3(aq) + 3H_2(g) \quad E° = 1.66\ V$$

Which statement is true if the hydrogen ion concentration is initially at 1.0 M and the initial pressure of hydrogen gas is 1.0 atm?

**(A)** Addition of a base should result in a value of $E$ which is less than 1.66 V.

**(B)** $n = 3$, because three moles of hydrogen is being produced.

**(C)** $E°$ is independent of the pH of the solution.

**(D)** $Q = \dfrac{[H_2]^3[AlCl_3]^2}{[Al]^2[HCl]^6}$

*Conceptual* **PQ-16.** According to the Nernst equation, when $E_{cell} = 0$ then

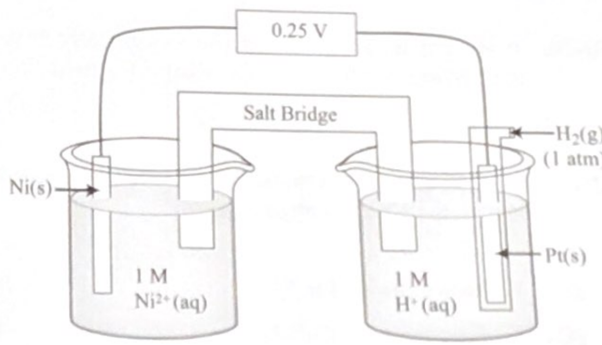
**(A)** $Q = K$ **(B)** $K = 1$ **(C)** $\Delta G° = 0$ **(D)** $E°_{cell} = 0$

**PQ-17.** What is $E_{cell}$ at 25 °C if $[Ag^+] = 0.200$ M and $[Cu^{2+}] = 0.200$ M?

$$2Ag^+(aq) + Cu(s) \rightarrow Cu^{2+}(aq) + 2Ag(s) \qquad E°_{cell} = 0.460\ V$$

**(A)** 0.419 V **(B)** 0.439 V **(C)** 0.460 V **(D)** 0.481 V

*Conceptual* **PQ-18.** A voltmeter connected to the cell under standard state conditions reads +0.25 V, and the standard free energy ($\Delta G°$) is –44 kJ·mol$^{-1}$.




If the concentration of $Ni^{2+}$ is increased to 2.0 M, the cell potential will _____ and $\Delta G$ will be ______ than $\Delta G°$.

$$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s) \qquad E° = -0.25\ V$$

**(A)** decrease, less negative **(B)** decrease, more negative

**(C)** increase, less negative **(D)** increase, more negative

**PQ-19.** What is the potential of a galvanic cell made from $Mn^{2+}|Mn$ and $Ag^+|Ag$ half-cells at 25 °C if $[Ag^+] = 0.10$ M and $[Mn^{2+}] = 0.10$ M?

| Standard Reduction Potentials | $E°$ |
|---|---|
| $Mn^{2+}(aq) + 2e^- \rightarrow Mn(s)$ | –1.18 V |
| $Ag^+(aq) + e^- \rightarrow Ag(s)$ | 0.80 V |

**(A)** 1.92 V **(B)** 1.95 V **(C)** 1.98 V **(D)** 2.01 V

**PQ-20.** What is the cell potential of the voltaic zinc-copper cell at 25 °C?

$$Zn(s)|Zn^{2+}(1.00 \times 10^{-5}\ M)||Cu^{2+}(1.00 \times 10^{-2})|Cu(s)$$

The standard cell potential of the cell is 1.10 V.

**(A)** 0.92 V **(B)** 1.28 V **(C)** 1.01 V **(D)** 1.19 V

Conceptual **PQ-21.** During the electrolysis of an aqueous solution of $CuSO_4$ using inert electrodes

**(A)** the anode loses mass and the cathode gains mass.

**(B)** the mass of the anode decreases but the mass of the cathode remains constant.

**(C)** the mass of the anode remains the same but the cathode gains mass.

**(D)** the anode and the cathode neither gain nor lose mass.

Conceptual **PQ-22.** Which products are formed during the electrolysis of a concentrated aqueous solution of sodium chloride?

**I** $Cl_2(g)$ **II** $NaOH(aq)$ **III** $H_2(g)$

**(A)** **I** only **(B)** **I** and **II** only **(C)** **I** and **III** only **(D)** **I, II,** and **III**

Conceptual **PQ-23.** The half-reaction occurring at the anode during the electrolysis of molten sodium bromide, NaBr, is ___.

**(A)** $Na^+ + e^- \rightarrow Na$ **(B)** $Na \rightarrow Na^+ + e^-$ **(C)** $Br_2 + 2e^- \rightarrow 2Br^-$ **(D)** $2Br^- \rightarrow Br_2 + 2e^-$

**PQ-24.** What mass of platinum could be plated on an electrode from the electrolysis of a $Pt(NO_3)_2$ solution with a current of 0.500 A for 55.0 s? The molar mass of platinum is 195.1 $g \cdot mol^{-1}$.

**(A)** 27.8 mg **(B)** 45.5 mg **(C)** 53.6 mg **(D)** 91.0 mg

**PQ-25.** A current of 5.00 A is passed through an aqueous solution of chromium(III) nitrate for 30.0 min. How many grams of chromium metal will be deposited at the cathode? The molar mass of chromium is 52.0 $g \cdot mol^{-1}$.

**(A)** 0.027 g **(B)** 1.62 g **(C)** 4.85 g **(D)** 6.33 g

**PQ-26.** A vanadium electrode is oxidized electrically. If the mass of the electrode decreases by 114 mg during the passage of 650 coulombs, what is the oxidation state of the vanadium product?

**(A)** +1 **(B)** +2 **(C)** +3 **(D)** +4

Conceptual **PQ-27.** Given 1 amp of current for 1 hour, which solution would deposit the smallest mass of metal?

| Molar mass / g·mol–1 | |
|---|---|
| Fe 55.9 | Ni 58.7 |
| Cu 63.5 | Ag 107.9 |

**(A)** Fe from aqueous $FeCl_2$ **(B)** Ni from aqueous $NiCl_2$

**(C)** Cu from aqueous $CuSO_4$ **(D)** Ag from aqueous $AgNO_3$

**PQ-28.** How many minutes will be required to deposit 1.00 g of chromium metal from an aqueous $CrO_4^{2-}$ solution using a current of 6.00 A?

**(A)** 186 min **(B)** 30.9 min **(C)** 15.4 min **(D)** 5.15 min

**PQ-29.** What is the $[Cu^{2+}]$ in the oxidation half-cell if the observed cell potential is 0.068 V?

$Cu(s) \mid Cu^{2+}(?\ M) \parallel Cu^{2+}(2.0\ M) \mid Cu(s)$

**(A)** 0.043 M **(B)** 0.010 M **(C)** 0.14 M **(D)** 2.0 M

Conceptual **PQ-30.** Which statement best describes a concentration cell when the solutions in each compartment are 0.50 M?

**(A)** The cell potential is zero. **(B)** The cell is an electrolytic cell.

**(C)** Anions will flow away from the anode. **(D)** Electrons will flow away from the cathode.

## Answers to Study Questions

1. C
2. C
3. B
4. B
5. B
6. A
7. B
8. B

## Answers to Practice Questions

1. B
2. B
3. A
4. A
5. C
6. C
7. B
8. B
9. A
10. A
11. C
12. B
13. C
14. B
15. A
16. A
17. B
18. A
19. B
20. D
21. C
22. D
23. D
24. A
25. B
26. C
27. D
28. B
29. B
30. A

# Chapter 16 – Nuclear Chemistry

*Chapter Summary:*

This chapter will focus on expressing and predicting nuclear reactions.

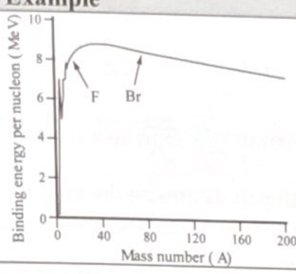
Specific topics covered in this chapter are:

- Radioactivity
- Radioactive decay
- Balancing nuclear reactions

Previous material that is relevant to your understanding of questions in this chapter include:

- Atomic Structure (***Chapter 1***)
- Electronic Structure (***Chapter 2***)
- Stoichiometry (***Chapter 4***)

*Common representations used in questions related to this material:*

| Name | Example | Used in questions related to |
|---|---|---|
| Graphs |  | nuclear binding energies |

*Where to find this in your textbook:*

The material in this chapter typically aligns to "Nuclear Chemistry" in your textbook. The name of your chapter(s) may vary.

*Practice exam:*

There are practice exam questions aligned to the material in this chapter. Because there are a limited number of questions on the practice exam, a review of the breadth of the material in this chapter is advised in preparation for your exam.

*How this fits into the big picture:*

The material in this chapter aligns to the Big Idea of Atomic Structure (1) as listed on page 12 of this study guide.

# Study Questions (SQ)

**SQ-1.** What is the missing particle in the nuclear equation?

$$^{230}_{90}\text{Th} \rightarrow {}^{226}_{88}\text{Ra} + \_\_\_\_$$

**(A)** $^{0}_{-1}\beta$ **(B)** $^{4}_{2}\text{He}$ **(C)** $^{1}_{1}\text{H}$ **(D)** $^{1}_{0}n$

***Knowledge Required:*** (1) How to balance a nuclear equation.

***Thinking it Through:*** You are asked in the question to determine the missing particle from a balanced nuclear equation. You can assume that all mass units are accounted for in the missing particle. Therefore, you note that the starting material (Th) undergoes the loss of 4 mass units in becoming the product shown (Ra) as *A* (remember, *A* is mass number or the number of protons plus the number of neutrons) goes from 230 ($^{230}Th$) to 226 ($^{226}Ra$). The missing particle will have a mass number of 4 ($^{4}X$).

In addition, you note that the starting material (Th) is 2 atomic numbers (or two protons) more than the product (Ra), as $Z$ goes from 90 ($_{90}Th$) to 88 ($_{88}Ra$). The missing particle will have an atomic number of 2 ($_{2}X$).

Four mass units with two of those as protons is a He particle. Therefore, you can conclude that this is an example of alpha decay and that $^{4}_{2}He$ is the missing particle, choice **(B)**.

Choice **(A)** is not correct because it is the particle lost for a beta decay equation.
Choice **(C)** is not correct because it is a proton particle.
Choice **(D)** is not correct because it is a neutron particle.

***Practice Questions Related to This:*** **PQ-1, PQ-2, PQ-3,** and **PQ-4**

**SQ-2.** A $^{235}_{92}U$ atom undergoes nuclear fission by being bombarded with a neutron to produce a total of four neutrons, a $^{140}_{55}Cs$ atom, and _____.

**(A)** $^{235}_{92}U$  **(B)** $^{235}_{93}Np$

**(C)** $^{92}_{37}Rb$  **(D)** $^{96}_{37}Rb$

***Knowledge Required:*** (1) Definition of nuclear fission. (2). How to write and balance a nuclear equation.

***Thinking it Through:*** You are asked in the question to determine the missing particle for a given nuclear fission reaction. The reaction is expressed in words; it is helpful in this situation to express the reaction as an equation:

$$^{235}_{92}U + ^{1}_{0}n \rightarrow ^{140}_{55}Cs + 4^{1}_{0}n + \_\_\_\_$$

Next, you balance the protons; there are 92 on the reactant side ($_{92}U$ and $_{0}n$) and 55 on the product side ($_{55}Cs$ and $4 \times {}_{0}n$) or $92-55 = 37$ protons missing for the unknown atom.

You then balance the mass numbers; there are $235 + 1$ on the reactant side ($^{235}U$ and $^{1}n$) and $140 + 4\times1$ on the product side ($^{140}Cs$ and $4 \times {}^{1}n$) or $236 - 144 = 92$ mass numbers missing for the unknown atom. Thus, $^{92}_{37}X$ is missing in order to balance the reaction; this is the atom $^{92}_{37}Rb$, Choice **(C)**.

Choice **(A)** is not correct because it assumes no change in the starting material.
Choice **(B)** is not correct because it is the result of positron emission, i.e. a neutron becomes a proton.
Choice **(D)** is not correct because it does not account for the four neutrons that are produced in the nuclear fission reaction.

***Practice Questions Related to This:*** **PQ-5**

**SQ-3.** An atom of the element of atomic number 53 and mass number 131 undergoes beta decay. The residual atom after this change has an atomic number of _____ and a mass number of ____.

**(A)** 54, 131  **(B)** 53, 131  **(C)** 53, 131  **(D)** 51, 127

***Knowledge Required:*** (1) Radiation emitted when alpha decay, beta decay, gamma emission, and positron emission occur. (2) How to write and balance a nuclear reaction.

***Thinking it Through:*** You are asked in the question to determine the atomic number and mass number for an atom after a particular radiation emission has occurred. In this instance, the atom undergoes beta decay (a $^{0}_{-1}\beta$ particle); this emission results in a net gain of one atomic number (or transformation such that the nucleus now has one additional proton). During beta decay, there is no overall change in mass number. You identify the element in the question as iodine, I ($Z = 53$).

$$^{131}_{53}I \rightarrow ^{0}_{-1}\beta + ^{131}_{54}Xe$$

Therefore, if a particle that is $^{131}_{53}I$ undergoes beta decay, the resulting atom is $^{131}_{54}Xe$ ($Z = 54$, $A = 131$), choice **(A)**.

Choice **(B)** is not correct because it would be the residual atom after a positron emission.
Choice **(C)** is not correct because it would be the residual atom after a gamma emission.
Choice **(D)** is not correct because it would be the residual atom after alpha decay.

*Practice Questions Related to This:* **PQ-6, PQ-7, PQ-8, PQ-9, PQ-10, PQ-11,** and **PQ-12**

## Practice Questions (PQ)

**PQ-1.** Cobalt-60 undergoes beta decay. What is the balanced nuclear reaction for this process?

**(A)** $^{60}_{27}Co \longrightarrow {}^{0}_{1}\beta + {}^{60}_{26}Fe$
**(B)** $^{60}_{27}Co \longrightarrow {}^{0}_{-1}\beta + {}^{60}_{28}Ni$
**(C)** $^{60}_{27}Co \longrightarrow {}^{4}_{2}\alpha + {}^{56}_{25}Mn$
**(D)** $^{60}_{27}Co + {}^{0}_{-1}e \longrightarrow {}^{60}_{26}Fe$

**PQ-2.** What is the missing particle in the balanced nuclear equation?

$$^{23}_{12}Mg \rightarrow {}^{23}_{11}Na + \_\_\_\_ + {}^{0}_{0}\nu$$

**(A)** $^{0}_{-1}\beta$ **(B)** $^{0}_{1}\beta$ **(C)** $^{1}_{1}H$ **(D)** $^{1}_{0}n$

**PQ-3.** What is the missing product of this reaction?

$$^{32}_{15}P \rightarrow {}^{32}_{16}S + \_\_\_\_$$

**(A)** $^{0}_{-1}\beta$ **(B)** $^{0}_{+1}\beta$ **(C)** $^{0}_{0}\gamma$ **(D)** $^{4}_{2}\alpha$

**PQ-4.** What is the missing particle in the balanced nuclear equation?

$$^{27}_{13}Al + {}^{4}_{2}He \rightarrow {}^{30}_{15}P + \_\_\_\_$$

**(A)** $^{0}_{-1}e$ **(B)** $^{0}_{1}e$ **(C)** $^{1}_{1}H$ **(D)** $^{1}_{0}n$

**PQ-5.** When $^{242}_{96}Cm$ captures an alpha particle, it produces a neutron and which isotope?

**(A)** $^{242}_{96}Cm$ **(B)** $^{245}_{98}Cf$ **(C)** $^{238}_{94}Pu$ **(D)** $^{239}_{94}Pu$

Conceptual **PQ-6.** Which particle, if lost from the nucleus, will *not* result in a change in the atomic mass number?

**(A)** proton **(B)** alpha particle **(C)** beta particle **(D)** neutron

**PQ-7.** An atom of the element of atomic number 84 and mass number 199 emits an alpha particle. The residual atom after this change has an atomic number of _____ and a mass number of _____.

**(A)** 82, 195 **(B)** 84, 203 **(C)** 85, 195 **(D)** 86, 199

**PQ-8.** What is the product when $^{218}_{84}Po$ undergoes positron emission?

**(A)** $^{214}_{82}Pb$ **(B)** $^{222}_{86}Rn$ **(C)** $^{214}_{81}Tl$ **(D)** $^{218}_{83}Bi$

**PQ-9.** Which equation qualifies as electron capture?

**(A)** $^{6}_{3}Li + {}^{1}_{0}n \rightarrow {}^{3}_{1}H + {}^{4}_{2}\alpha$
**(B)** $^{26}_{13}Mg + {}^{1}_{1}p \rightarrow {}^{23}_{11}Na + {}^{4}_{2}\alpha$
**(C)** $^{37}_{18}Ar + {}^{0}_{-1}e \rightarrow {}^{37}_{17}Cl + h\nu$
**(D)** $^{38}_{19}K \rightarrow {}^{38}_{20}Ca + {}^{0}_{-1}\beta$

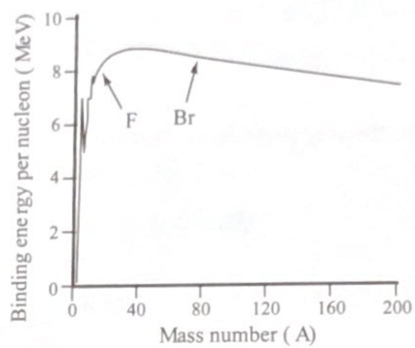
Conceptual **PQ-10.** Which nuclear process does not necessarily result in nuclear transmutation (the transformation of one element into another)?

**(A)** alpha ($\alpha$) emission
**(B)** beta ($\beta$) emission
**(C)** gamma ($\gamma$) emission
**(D)** positron emission

**PQ-11.** Medical imaging is used when a patient ingests a radioactive emitter, called a tracer, so a doctor can view an internal system. Which tracer emission will penetrate through the body?

**(A)** alpha particle
**(B)** beta particle
**(C)** gamma particle
**(D)** helium nucleus

Conceptual **PQ-12.** Based on their binding energies, what nuclear transformations are predicted for F and Br?




**(A)** Both could undergo nuclear fission.
**(B)** Both could undergo nuclear fusion.
**(C)** F could undergo nuclear fission and Br could undergo nuclear fusion.
**(D)** Br could undergo nuclear fission and F could undergo nuclear fusion.

## Answers to Study Questions

1. B
2. C
3. A

## Answers to Practice Questions

1. B
2. B
3. A
4. D
5. B
6. C
7. A
8. D
9. C
10. C
11. C
12. D