

Bubble in **A**, **B** or **C** as the test form code at the top right of your answer sheet. Bubble in your Section Number also.

VERSION X

- A reaction has $\Delta H = -75.0$ kJ and $\Delta S = -245$ J/K. At what temperatures will it be spontaneous?

 - Temperatures above 306K
 - The reaction is spontaneous at all temperatures.
 - The reaction is non-spontaneous at all temperatures.
 - Temperatures below 306K
 - Impossible to determine.
- Calculate ΔG at 298 K for the reaction $\text{XeF}_2(\text{g}) \rightarrow \text{Xe}(\text{g}) + \text{F}_2(\text{g})$ $\Delta G^\circ = -280$ kJ

if the gas pressures (p) are $p(\text{Xe}) = 0.50$ atm, $p(\text{F}_2) = 24$ atm, and $p(\text{XeF}_2) = 0.2$ atm.

 - 290 kJ
 - 280 kJ
 - 270 kJ
 - 260 kJ
 - none of these
- What is the equilibrium constant of the reaction below at 298 K?

$$\text{BaO}(\text{s}) + \text{CO}_2(\text{g}) \rightarrow \text{BaCO}_3(\text{s}) \quad \Delta G^\circ = -224 \text{ kJ}$$
 - 1.92×10^{39}
 - 5.20×10^{-40}
 - 1.09
 - 90.5
 - none of these
- The following reaction has $\Delta H^\circ = +25.68$ kJ and $\Delta S^\circ = -4.8$ J/K. What is the solubility product constant (K_{sp}) for $\text{PbCl}_2(\text{s})$ at 298 K?

$$\text{PbCl}_2(\text{s}) \rightarrow \text{Pb}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq})$$
 - 1.8×10^{-5}
 - 5.8×10^{-4}
 - 2.6×10^{-2}
 - 8.6×10^{-6}
 - impossible to determine
- At 350 K, the dissociation of HBrO in water has $K_a = 3.8 \times 10^{-9}$. Calculate ΔG° , and calculate ΔG when $[\text{H}_3\text{O}^+] = 4.0 \times 10^{-4}$ M, $[\text{BrO}^-] = 0.20$ M and $[\text{HBrO}] = 0.40$ M.

$$\text{HBrO}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{BrO}^-(\text{aq})$$
 - $\Delta G^\circ = -56$ kJ, $\Delta G = -81$ kJ
 - none of these
 - $\Delta G^\circ = 56$ kJ, $\Delta G = 32$ kJ
 - $\Delta G^\circ = 56$ kJ, $\Delta G = 81$ kJ
 - $\Delta G^\circ = -56$ kJ, $\Delta G = -32$ kJ
- Consider the following half-reactions:

$\text{Fe}^{2+}(\text{aq}) + 2 \text{e}^- \rightarrow \text{Fe}(\text{s})$	$E^\circ = -0.44$ V
$\text{I}_2(\text{s}) + 2 \text{e}^- \rightarrow 2 \text{I}^-(\text{aq})$	$E^\circ = 0.53$ V
$\text{O}_2(\text{g}) + 2 \text{H}^+(\text{aq}) + 2 \text{e}^- \rightarrow \text{H}_2\text{O}_2(\text{aq})$	$E^\circ = 0.68$ V

the order of oxidizing agent strength is:

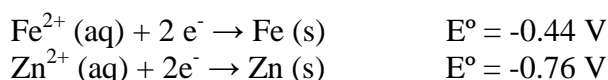
 - $\text{H}_2\text{O}_2(\text{aq}) > \text{I}^-(\text{aq}) > \text{Fe}(\text{s})$
 - $\text{O}_2(\text{g}) > \text{I}_2(\text{s}) > \text{Fe}^{2+}(\text{aq})$
 - $\text{Fe}(\text{s}) > \text{I}^-(\text{aq}) > \text{H}_2\text{O}_2(\text{aq})$
 - $\text{Fe}^{2+}(\text{aq}) > \text{I}_2(\text{s}) > \text{O}_2(\text{g})$
 - cannot be determined

7. Consider the following three half-reactions
- $$\begin{array}{l} \text{Sn}^{2+}(\text{aq}) + 2 \text{e}^{-} \rightarrow \text{Sn}(\text{s}) \quad E^{\circ} = -0.14 \text{ V} \\ \text{Mn}^{2+}(\text{aq}) + 2 \text{e}^{-} \rightarrow \text{Mn}(\text{s}) \quad E^{\circ} = -1.18 \text{ V} \\ \text{Ce}^{4+}(\text{aq}) + \text{e}^{-} \rightarrow \text{Ce}^{3+}(\text{aq}) \quad E^{\circ} = 1.61 \text{ V} \end{array}$$

the order of reducing agent strengths is:

- $\text{Mn}^{2+}(\text{aq}) > \text{Sn}^{2+}(\text{aq}) > \text{Ce}^{4+}(\text{aq})$
 - $\text{Ce}^{4+}(\text{aq}) > \text{Sn}^{2+}(\text{aq}) > \text{Mn}^{2+}(\text{aq})$
 - $\text{Mn}(\text{s}) > \text{Sn}(\text{s}) > \text{Ce}^{3+}(\text{aq})$
 - $\text{Ce}^{3+}(\text{aq}) > \text{Sn}(\text{s}) > \text{Mn}(\text{s})$
 - cannot be determined
8. In the redox reaction: $\text{N}_2(\text{g}) + 5 \text{H}^{+}(\text{aq}) + 2 \text{Mg}(\text{s}) \rightarrow \text{N}_2\text{H}_5^{+}(\text{aq}) + 2 \text{Mg}^{2+}(\text{aq})$
- the reducing agent is $\text{H}^{+}(\text{aq})$, and the oxidizing agent is $\text{Mg}(\text{s})$
 - the reducing agent is $\text{N}_2(\text{g})$, and the oxidizing agent is $\text{Mg}(\text{s})$
 - the oxidizing agent is $\text{H}^{+}(\text{aq})$, and the reducing agent is $\text{Mg}(\text{s})$
 - the oxidizing agent is $\text{N}_2(\text{g})$, and the reducing agent is $\text{Mg}(\text{s})$
 - the oxidizing agent is $\text{N}_2(\text{g})$, and the reducing agent is $\text{H}^{+}(\text{aq})$
9. In the redox reaction: $2\text{Fe}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 2\text{H}^{+}(\text{aq}) \rightarrow 2\text{Fe}^{3+}(\text{aq}) + \text{SO}_3^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
- The $\text{SO}_4^{2-}(\text{aq})$ is oxidized to $\text{SO}_3^{2-}(\text{aq})$ by reducing the $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$.
 - The $\text{H}^{+}(\text{aq})$ is the oxidizing agent which oxidizes the $\text{SO}_4^{2-}(\text{aq})$ to $\text{SO}_3^{2-}(\text{aq})$.
 - The $\text{SO}_4^{2-}(\text{aq})$ is the reducing agent and $\text{Fe}^{2+}(\text{aq})$ is the oxidizing agent.
 - The $\text{Fe}^{2+}(\text{aq})$ is the reducing agent, and the $\text{SO}_4^{2-}(\text{aq})$ is being oxidized.
 - The $\text{SO}_4^{2-}(\text{aq})$ is reduced to $\text{SO}_3^{2-}(\text{aq})$ by oxidizing the $\text{Fe}^{2+}(\text{aq})$ to $\text{Fe}^{3+}(\text{aq})$.
10. For the voltaic cell
- $$\text{Ni}(\text{s}) \mid \text{Ni}^{2+}(\text{aq}) \parallel \text{Cl}_2(\text{g}) \mid \text{Cl}^{-}(\text{aq}) \mid \text{graphite}$$
- the cell reaction is:
- $\text{Ni}(\text{s}) + \text{Cl}_2(\text{g}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq})$
 - $\text{Ni}(\text{s}) + 2 \text{Cl}^{-}(\text{aq}) \rightarrow \text{Ni}^{2+}(\text{aq}) + \text{Cl}_2(\text{g})$
 - $\text{Ni}^{2+}(\text{aq}) + 2 \text{Cl}^{-}(\text{aq}) \rightarrow \text{Ni}(\text{s}) + \text{Cl}_2(\text{g})$
 - $\text{Ni}^{2+}(\text{aq}) + \text{Cl}_2(\text{g}) \rightarrow \text{Ni}(\text{s}) + 2 \text{Cl}^{-}(\text{aq})$
 - impossible to determine without E°_{cell}
11. What is the cell notation for the voltaic cell in which the following reaction occurs?
- $$\text{Cr}(\text{s}) + \text{Sn}^{4+}(\text{aq}) \rightarrow \text{Cr}^{2+}(\text{aq}) + \text{Sn}^{2+}(\text{aq})$$
- $\text{Sn}(\text{s}) \mid \text{Sn}^{2+}(\text{aq}) \parallel \text{Cr}^{2+}(\text{aq}) \mid \text{Cr}(\text{s})$
 - $\text{Cr}(\text{s}) \mid \text{Cr}^{2+}(\text{aq}) \parallel \text{Sn}^{2+}(\text{aq}), \text{Sn}^{4+}(\text{aq}) \mid \text{graphite}$
 - $\text{graphite} \mid \text{Sn}^{4+}(\text{aq}), \text{Sn}^{2+}(\text{aq}) \parallel \text{Cr}^{2+}(\text{aq}) \mid \text{Cr}(\text{s})$
 - $\text{Cr}(\text{s}) \mid \text{Cr}^{2+}(\text{aq}) \parallel \text{Sn}^{4+}(\text{aq}), \text{Sn}^{2+}(\text{aq}) \mid \text{graphite}$
 - $\text{Cr}(\text{s}) \mid \text{Cr}^{2+}(\text{aq}) \parallel \text{Sn}^{4+}(\text{aq}) \mid \text{Sn}(\text{s})$
12. A voltaic cell prepared using bromine and nickel has the following cell notation:
- $$\text{Ni}(\text{s}) \mid \text{Ni}^{2+}(\text{aq}) \parallel \text{Br}_2(\text{l}) \mid \text{Br}^{-}(\text{aq}) \mid \text{graphite}$$
- Which of the following reactions occurs at the cathode?
- $2 \text{Br}^{-}(\text{aq}) \rightarrow \text{Br}_2(\text{l}) + 2 \text{e}^{-}$
 - $\text{Ni}(\text{s}) \rightarrow \text{Ni}^{2+}(\text{aq}) + 2 \text{e}^{-}$
 - none of these
 - $\text{Br}_2(\text{l}) + 2 \text{e}^{-} \rightarrow 2 \text{Br}^{-}(\text{aq})$
 - $\text{Ni}^{2+}(\text{aq}) + 2 \text{e}^{-} \rightarrow \text{Ni}(\text{s})$

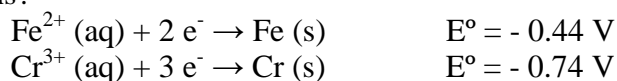
13. For the voltaic (galvanic) cell consisting of the Fe (s)/Fe²⁺ (aq, 1.0 M) and Zn (s) / Zn²⁺ (aq, 1.0 M) half-cells, given that



Which of the following statements is true?

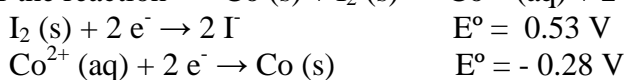
- The iron (Fe) electrode is the anode and electrons flow to it from the zinc (Zn) cathode through the external circuit.
 - The iron (Fe) electrode is the cathode and electrons flow to it from the zinc (Zn) electrode through the external circuit.
 - Since both E^o half-potentials are negative, no spontaneous reactions are possible and no current will flow.
 - The concentration of Fe²⁺ (aq) increases during operation of the cell.
 - Since [Fe²⁺] = [Zn²⁺], Q = 1, the cell is at equilibrium and no current will flow.
14. What is E_{cell}^o for the voltaic cell with cell reaction Fe (s) + Cl₂ (g) → Fe²⁺ (aq) + 2 Cl⁻ (aq)
- 0.00 V
 - 0.92 V
 - 2.13 V
 - 0.59 V
 - 1.80 V

15. What spontaneous reaction will occur in a voltaic cell constructed using the following two half-reactions?



- 2 Cr (s) + 3 Fe²⁺ (aq) → 2 Cr³⁺ (aq) + 3 Fe (s)
 - 2 Cr³⁺ (aq) + 3 Fe (s) → 2 Cr (s) + 3 Fe²⁺ (aq)
 - Cr³⁺ (aq) + 3 Fe²⁺ (aq) → Cr (s) + 3 Fe³⁺ (aq)
 - 2 Cr (s) + 2 Fe³⁺ (aq) → 2 Cr³⁺ (aq) + 2 Fe (s)
 - impossible to determine without ΔG^o
16. A metal M (s) spontaneously reacts with Ni²⁺ (aq) to produce nickel metal, Ni (s). Which of the three metals Cu, Fe and Pb can M be?
- $$\text{M} (\text{s}) + \text{Ni}^{2+} (\text{aq}) \rightarrow \text{M}^{2+} (\text{aq}) + \text{Ni} (\text{s})$$
- cannot be determined
 - all three of them
 - Cu and Pb only
 - Fe only
 - None of them

17. Calculate ΔG^o for the reaction Co (s) + I₂ (s) → Co²⁺ (aq) + 2 I⁻ (aq) at 298 K.



- 156 kJ
 - +28 kJ
 - +86 kJ
 - 78 kJ
 - 48 kJ
18. Which of the metals, cadmium (Cd), mercury (Hg), and strontium (Sr), can be used to produce H₂ gas from a solution of an acid such as HCl (aq)?
- all three of them
 - Hg only
 - Cd and Sr only
 - none of them
 - Cd only

19. In the electrolysis of a molten mixture of the salts CaCl_2 (l) and MnI_2 (l), the products at the cathode and anode are
- Ca metal and iodine (I_2)
 - Mn metal and chlorine (Cl_2)
 - Mn metal and iodine (I_2)
 - Ca metal and chlorine (Cl_2)
 - cannot be determined
20. For the voltaic (galvanic) cell $\text{Mg (s)} \mid \text{Mg}^{2+} \text{ (aq, 1.0 M)} \parallel \text{Cr}^{3+} \text{ (aq, 1.0 M)} \mid \text{Cr (s)}$ Which of the following statements is not true?
- The mass of the Cr (s) electrode will increase during operation of the cell.
 - $Q = [\text{Mg}^{2+}] / [\text{Cr}^{3+}]$
 - If the initial concentrations had been $[\text{Mg}^{2+}] = [\text{Cr}^{3+}] = 2.0 \text{ M}$, the cell potential would have been less than E°_{cell} .
 - The magnesium (Mg) electrode is the anode and electrons flow from it to the chromium (Cr) cathode through the external circuit.
 - $E^\circ_{\text{cell}} = 1.63 \text{ V}$
21. Consider the two half-reactions below:
- $$\text{Fe}^{3+} \text{ (aq)} + \text{e}^- \rightarrow \text{Fe}^{2+} \text{ (aq)} \quad E^\circ = 0.77 \text{ V}$$
- $$\text{Ce}^{4+} \text{ (aq)} + \text{e}^- \rightarrow \text{Ce}^{3+} \text{ (aq)} \quad E^\circ = 1.61 \text{ V}$$
- The spontaneous reaction obtained from these half-reactions is:
- cannot be determined without ΔG°
 - $\text{Fe}^{2+} \text{ (aq)} + \text{Ce}^{3+} \text{ (aq)} \rightarrow \text{Fe}^{3+} \text{ (aq)} + \text{Ce}^{4+} \text{ (aq)}$
 - $\text{Fe}^{3+} \text{ (aq)} + \text{Ce}^{3+} \text{ (aq)} \rightarrow \text{Fe}^{2+} \text{ (aq)} + \text{Ce}^{4+} \text{ (aq)}$
 - $\text{Fe}^{3+} \text{ (aq)} + \text{Ce}^{4+} \text{ (aq)} \rightarrow \text{Fe}^{2+} \text{ (aq)} + \text{Ce}^{3+} \text{ (aq)}$
 - $\text{Fe}^{2+} \text{ (aq)} + \text{Ce}^{4+} \text{ (aq)} \rightarrow \text{Fe}^{3+} \text{ (aq)} + \text{Ce}^{3+} \text{ (aq)}$
22. What is the cell potential for the following concentration cell at 298 K?
- $$\text{Zn (s)} \mid \text{Zn}^{2+} \text{ (aq, 0.22 M)} \parallel \text{Zn}^{2+} \text{ (aq, 2.2 M)} \mid \text{Zn (s)}$$
- 0.03 V
 - 0.06 V
 - 0.03 V
 - 0.06 V
 - 0.00
23. Consider the reaction (I) $\text{FeO (s)} \rightarrow \text{Fe (s)} + \frac{1}{2} \text{O}_2 \text{ (g)}$ $\Delta G^\circ = +250 \text{ kJ}$ Which of the following reactions (II) and (III) could be coupled with reaction (I) to provide a spontaneous reaction for converting FeO (s) to Fe (s) ?
- $$\text{(II)} \quad \text{C (s)} + \text{O}_2 \text{ (g)} \rightarrow \text{CO}_2 \text{ (g)} \quad \Delta G^\circ = -394 \text{ kJ}$$
- $$\text{(III)} \quad 2 \text{H}_2 \text{ (g)} + \text{O}_2 \text{ (g)} \rightarrow 2 \text{H}_2\text{O (g)} \quad \Delta G^\circ = -229 \text{ kJ}$$
- neither (II) nor (III)
 - reaction (II)
 - reaction (III)
 - either (II) or (III)
 - cannot be determined without ΔH° values.
24. Which of the following is true for a voltaic (galvanic) cell?
- the E°_{cell} value is negative if $Q > 1$
 - the cell reaction can never reach equilibrium
 - the cell potential E_{cell} does not depend on the concentrations of species in solution
 - the cell potential (E_{cell}) decreases during operation of the cell
 - E_{cell} cannot be greater than E°_{cell}
25. What mass of nickel (Ni) will be deposited on the cathode by electrolysis of a Ni^{2+} (aq) solution for 60.0 minutes using a constant current of 5.00 amperes? Atomic weight of Ni is 58.7 g/mol.
- 2.74 g
 - 10.94 g
 - 5.47 g
 - 0.912 g
 - none of these

26. Consider the following voltaic cell at 298 K.

$$\text{Cr (s)} \mid \text{Cr}^{3+} (\text{aq}, 1.0 \text{ M}) \parallel \text{Ni}^{2+} (\text{aq}, 1.0 \text{ M}) \mid \text{Ni (s)}$$
 What is E°_{cell} ?
- 0.49 V
 - 0.49 V
 - 0.99 V
 - 0.99 V
 - cannot be determined
27. Consider the voltaic cell below, containing equal volumes of the two solutions:

$$\text{Co (s)} \mid \text{Co}^{2+} (\text{aq}, 1.0 \text{ M}) \parallel \text{Ni}^{2+} (\text{aq}, 1.0 \text{ M}) \mid \text{Ni (s)}$$
 After current has flowed for 30 minutes, the $[\text{Ni}^{2+}]$ is 0.60 M. Which of the following statements is true?
- the $[\text{Co}^{2+}]$ after 30 minutes is 0.60 M
 - the $[\text{Co}^{2+}]$ after 30 minutes is 1.0 M
 - none of these
 - the $[\text{Co}^{2+}]$ after 30 minutes is 0.02 M
 - the $[\text{Co}^{2+}]$ after 30 minutes is 1.40 M
28. Consider the following voltaic cell at 298 K.

$$\text{Co (s)} \mid \text{Co}^{2+} (\text{aq}, 0.8 \text{ M}) \parallel \text{Ni}^{2+} (\text{aq}, 8.0 \text{ M}) \mid \text{Ni (s)}$$
 What is E_{cell} if $E^\circ_{\text{cell}} = 0.03 \text{ V}$?
- 0.00 V
 - 0.03 V
 - 0.09 V
 - 0.06 V
 - cannot be determined
29. Consider the equilibrium below at 298 K. At equilibrium, $[\text{X}] = [\text{Y}] = 2.0 \text{ M}$.

$$3 \text{ X (aq)} \rightleftharpoons 2 \text{ Y (aq)} \quad \Delta S^\circ = -225 \text{ J/K}$$
 What is ΔG° for this reaction?
- $\Delta G^\circ = +1.7 \text{ kJ}$
 - $\Delta G^\circ = 0.0 \text{ kJ}$
 - $\Delta G^\circ = -1.7 \text{ kJ}$
 - $\Delta G^\circ = -3.4 \text{ kJ}$
 - cannot be determined without ΔH°
30. A voltaic cell is driven by the following spontaneous reaction involving the metal M (s).

$$\text{M (s)} + 2 \text{ Cu}^{2+} (\text{aq}) \rightarrow \text{M}^{2+} (\text{aq}) + 2 \text{ Cu}^+ (\text{aq})$$
 If the $E^\circ_{\text{cell}} = 0.28 \text{ V}$, the metal M is
- Cd
 - Zn
 - Pb
 - Fe
 - Al
-