Recitation Worksheet 13: Electrochemistry

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3. Answers must be written in the corresponding answer box, or no credit will be awarded.
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1. Which assignment of oxidation number is **INCORRECT** for the underlined element? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

$\boxed{BE}$

~~A.~~ $K_2\underline{Cr}_2O_7$; +6
B. $\underline{N}H_3$; +3
~~C.~~ $H_2\underline{P}O_2^-$; +1
D. $\underline{Se}O_3^{2-}$; +4
E. $Cu(\underline{N}O_3)_2$; +2

A) K = 1+, Cr = ?, and O = 2- ($K_2Cr_2O_7$)
$2(1+) + 2Cr + 7(2-) = 0$
$\therefore 2Cr + (12-) = 0$
$2Cr = 12$ $\therefore Cr = 6+$

B) N = ?, H = 1+ (NH_3)
$N + (3 \times 1+) = 0$
$N = 3-$

C) H = 1+, P = ?, and O = 2- ($H_2PO_2^-$)
$(2 \times 1+) + P + (2 \times 2-) = 1-$
$(2) + P + (-4)$
$P + (-2) = -1$
$P = +1$

D) Se = ?, O = 2- (SeO_3^{2-})
$Se + (3 \times -2) = -2$
$Se + (-6) = -2$
$Se = +4$

E) Cu = 2+, N = ?, O = 2- ($Cu(NO_3)_2$)
$(2+) + 2N + 6(2-) = 0$
$2N + (10-) = 0$
$2N = 10$
$\therefore N = +5$

2. What is the coefficient of **Mn^{2+}** if the reaction below occurs in **acidic solution**?

$$S_2O_3^{2-} + MnO_4^- \longrightarrow SO_4^{2-} + Mn^{2+}$$

$\boxed{C}$

A. 5
B. 7
C. 8
D. 10
E. 14

① Reduction: $(5e^- + 8H^+ + MnO_4^- \longrightarrow Mn^{2+} + 4H_2O) \times 8$

Oxidation: $(5H_2O + S_2O_3^{2-} \longrightarrow 2SO_4^{2-} + 10H^+ + 8e^-) \times 5$

② Reduction: $\cancel{40e^-} + \overset{14H^+}{\cancel{64}H^+} + 8MnO_4^- \longrightarrow 8Mn^{2+} + \overset{7H_2O}{\cancel{32H_2O}}$

Oxidation: $\cancel{25H_2O} + 5S_2O_3^{2-} \longrightarrow 10SO_4^{2-} + \cancel{50H^+} + \cancel{40e^-}$

$$14H^+ + 8MnO_4^- + 5S_2O_3^{2-} \longrightarrow 8Mn^{2+} + 10SO_4^{2-} + 7H_2O$$

3. What is the **reducing agent** in the following reaction between hexacyanoferrate (II) complex and hydrogen peroxide in acidic solution?

Oxidizing agent

$$2\,[\overset{2+}{Fe}(\overset{1-}{CN})_6]^{4-} + \overset{+1}{H_2}\overset{-1}{O_2} + 2\,\overset{1+}{H^+} \longrightarrow 2\,[\overset{3+}{Fe}(\overset{1-}{CN})_6]^{3-} + 2\,\overset{1+}{H_2}\overset{2-}{O}$$

Reducing agent

$\boxed{A}$

A. $[Fe(CN)_6]^{4-}$
B. H_2O_2
C. H^+
D. $[Fe(CN)_6]^{3-}$
E. H_2O

Strategy: track the oxidation state of each element & the reducing agent will contain the element that its oxidation state increases

4. Which statement is true about redox reactions?

$\boxed{D}$

~~A.~~ A half-reaction can occur by itself.
~~B.~~ A redox reaction in base can include excess H^+ after it has been balanced.
~~C.~~ Two oxidations can occur instead of one oxidation and one reduction.
D. At least 2 atoms must have their oxidation states change during a redox reaction.
~~E.~~ None of these statements are true.

(A) the oxidation & reduction half-reactions occur by themselves
(B) the excess H+ will be balanced by adding OH- on both sides of the equation
(C) should be one oxidation & one reduction

Reduction

Example: $2Al^0 + 3Cu^{2+} \longrightarrow 2Al^{3+} + 3Cu^0$

Oxidation

In this example the oxidation states of both Al & Cu change

5. Which of the following equations are an oxidation-reduction reaction? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

Types of Redox Reactions
① Single displacement
② Combustion

$\boxed{BCD}$

~~A.~~ $H_2CO_3(aq) \longrightarrow H_2O(l) + CO_2(g)$
B. $2\ Li(s) + 2\ H_2O(l) \longrightarrow 2\ LiOH(aq) + H_2(g)$
C. $C_3H_8(g) + 5\ O_2(g) \longrightarrow 3\ CO_2(g) + 5\ H_2O(g)$
D. $4\ Ag(s) + PtCl_4(aq) \longrightarrow 4\ AgCl(s) + Pt(s)$
~~E.~~ $2\ HClO_4(aq) + Ca(OH)_2(aq) \longrightarrow 2\ H_2O(l) + Ca(ClO_4)_2(aq)$
~~F.~~ $3\ Cu(NO_3)_2(aq) + 2\ Na_3PO_4(aq) \longrightarrow 6\ NaNO_3 + Cu_3(PO_4)_2$

6. For the reaction:

$$3\ Pt(s) + 12\ Cl^-(aq) + 2\ NO_3^-(aq) + 8\ H^+(aq) \longrightarrow 3\ [PtCl_4]^{2-}(aq) + 2\ NO(g) + 4\ H_2O(l) \quad E°_{cell} = 0.201\ V$$

What is *E*° for the reduction half-cell reaction of $[PtCl_4]^{2-}$ to Pt in acidic solution? Please refer to the table below for additional information. Keep your answers to two significant figures.

$\boxed{0.76}$ V

Standard Reduction Potentials at 25°C (298 K) for Many Common Half-reactions

Half-reaction	ℰ° (V)	Half-reaction	ℰ° (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$IO_4^- + 2H^+ + 2e^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	−0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	−0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	−0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	−0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	−0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	−0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	−0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	−0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	−1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	−2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	−2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	−2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	−2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	−2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	−2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	−3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

oxid: $3\ Pt(s) + 12\ Cl^-(aq) \longrightarrow 3\ [PtCl_4]^{2-}(aq)$ $E°_{cell} = ?$
Red: $2NO_3^-(aq) + 8H^+(aq) \longrightarrow 2NO + 4H_2O(l)$ $E°_{cell} = +0.96V$

$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode})$
$0.201\ V = 0.96V - E_{red}(\text{anode})$
$\therefore\ E_{anode} = 0.759 \sim 0.76V$

7. Which of the reactions do you predict to be **non-spontaneous** in the forward direction? Assume all the reactants and products in their standard states. Please refer to the table in question 6 for additional information.

$\boxed{E}$

A. $Cu^{2+}(aq) + 2I^{-}(aq) \longrightarrow Cu(s) + I_2(s)$
B. $4\ NO_3^{-}(aq) + 4\ H^{+}(aq) \longrightarrow 3\ O_2(g) + 4\ NO(g) + 2\ H_2O(l)$
C. $2\ Br^{-}(aq) + I_2(s) \longrightarrow Br_2(aq) + 2\ I^{-}(aq)$
D. $Au(s) + NO_3^{-}(aq) \longrightarrow Au^{3+}(aq) + NO(g)$ (in acidic solution)
E. All of the above reactions are non-spontaneous in the forward direction.

Strategy: Determine the E° cell for each reaction. If the E° cell for the reaction is positive ∴ the reaction is spontaneous

(A) $Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$ $E^\circ(red) = 0.34V$ Reduction (cathode)

$2I^-(aq) \longrightarrow I_2(s) + 2e^-$ $E^\circ(red) = 0.54V$ Oxidation (Anode)

non-spontaneous

$E^\circ cell = E^\circ cathode - E^\circ Anode = 0.34 - 0.54 = -0.20V$

(B) $(NO_3^- + 4H^+ + 3e^- \longrightarrow NO + 2H_2O) \times 4$ Reduction $E^\circ red = 0.96$

$(2H_2O \longrightarrow O_2 + 4H^+ + 4e^-) \times 3$ oxidation $E^\circ red = 1.23$

$4\ NO_3^- + \cancel{16H}^+ + \cancel{12e^-} \longrightarrow 4\ NO + 8\cancel{H_2O}$

4 H^+ $2H_2O$

$6\cancel{H_2O} \longrightarrow 3O_2 + \cancel{12H}^+ + \cancel{12e^-}$

$E^\circ cell = 0.96 - 1.23 = -0.27V$ non-spontaneous

(C) $2Br^- \longrightarrow Br_2 + 2e^-$ oxidation $E^\circ cell = 1.09V$

$2e^- + I_2 \longrightarrow 2I^-$ Reduction $E^\circ cell = 0.54V$

$E^\circ cell = 0.54 - 1.09 = -0.55V$ non-spontaneous

(D)

$Au \rightarrow Au^{3+} + 3e^-$ oxidation ($E^\circ cell = 1.50$)

$3e^- + 4H^+ + NO_3^- \longrightarrow NO + 2H_2O$ Reduction (E° cell = 0.96)

$E^\circ cell = 0.960 - 1.50 = -0.54V$

non-spontaneous

8. Calculate the $E°_{cell}$ for the following reactions. Please refer to the table in question 6 for additional information. Keep your answers to two decimal places.

(Anode: Oxidation)
(Cathode = reduction)

A. Anode: $Cu(s)|Cu^{2+}(aq)||Cu^{+}(aq)|Cu(s)$ Cathode; || = salt bridge

$\boxed{0.18}$ V

$E°_{cell} = E°_{cathode} - E°_{anode} = 0.52 - 0.34 = 0.18\,V$

$Cu(s) \rightarrow Cu^{2+} + 2e^-$ $E°_{red} = 0.34\,V$ (Oxidation)

$2e^- + 2Cu^{+}(aq) \rightarrow 2Cu(s)$ $E°_{red} = 0.52\,V$ (Reduction)

B. Anode: $Pt(s)|Ce^{4+}(aq), Ce^{3+}(aq)||I^{-}(aq)|I_2(s)|C(s)$ Cathode

$\boxed{-1.16}$ V

$2Ce^{3+}(aq) \rightarrow 2Ce^{4+}(aq) + 2e^-$ $E° = 1.70$ (Oxidation)

$2e^- + I_2(s) \rightarrow 2I^-(aq)$ $E° = 0.54$ (Reduction)

$E°_{cell} = 0.54 - 1.70 = -1.16\,V$

C. Anode: $Al(s)|Al^{3+}(aq)||Sn^{2+}(aq)|Sn(s)$ Cathode

$\boxed{1.52}$ V

$2Al(s) \rightarrow 2Al^{3+}(aq) + 6e^-$ $E° = -1.66\,V$ (Oxidation)

$6e^- + 3Sn^{2+}(aq) \rightarrow 3Sn(s)$ $E° = -0.14\,V$ (Reduction)

$E°_{cell} = -0.14 - (-1.66) = 1.52\,V$

9. Which statements below is **FALSE** regarding standard cell potentials? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

$\boxed{BCE}$

~~A.~~ $E°_{cell}$ is positive for spontaneous reactions. True ($E°_{cell}$ is negative for non-spontaneous reactions).

B. Electrons will flow from the positive electrode to the negative electrode in a galvanic cell. False (from negative to positive)

C. $E°_{cell}$ is the difference in voltage between the anode and the cathode, $E°_{cell} = E°_{cell}$(anode) - $E°_{cell}$ (cathode). False. $E°_{cell} = E°_{cathode} - E°_{anode}$

~~D.~~ The electrode potential of the standard hydrogen electrode is exactly zero. true

E. The electrode in any half-cell with a greater tendency to undergo reduction is negatively charged relative to the standard hydrogen electrode and therefore has a negative $E°$. False → positively charged

10. What is the redox reaction represented by the following cell notation?

Anode: $Mn(s) \mid Mn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$ Cathode; || = salt bridge

$\boxed{B}$

A. $Cu(s) + Mn^{2+}(aq) \rightarrow Mn(s) + Cu^{2+}(aq)$
B. $Mn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Mn^{2+}(aq)$
C. $2\,Mn(s) + Cu^{2+}(aq) \rightarrow Cu(s) + 2\,Mn^{2+}(aq)$
D. $2\,Cu(s) + Mn^{2+}(aq) \rightarrow Mn(s) + 2\,Cu^{2+}(aq)$
E. $3\,Mn(s) + 2\,Cu^{2+}(aq) \rightarrow 2\,Cu(s) + 3\,Mn^{2+}(aq)$

Anode = Oxidation
Cathode = Reduction

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

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

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

11. Which of the choices represents the correct cell notation for $Sn(s) + 2\ H^+(aq) \longrightarrow Sn^{2+}(aq) + H_2(g)$?

$\boxed{D}$

~~A.~~ $H^+(aq)\ |\ H_2(g)\ |\ Pt\ \|\ Sn(s)\ |\ Sn^{2+}(aq)$
~~B.~~ $H_2(g)\ |\ H^+(aq)\ |\ Pt\ \|\ Sn^{2+}(aq)\ |\ Sn(s)$
~~C.~~ $Sn^{2+}(aq)\ |\ Sn(s)\ \|\ H_2(g)\ |\ H^+(aq)\ |\ Pt$
D. $Sn(s)|\ Sn^{2+}\ (aq)\ \|\ H^+(aq)\ |\ H_2(g)\ |\ Pt$
~~E.~~ $Sn(s)\ |\ H_2(g)\ \|\ Sn^{2+}(aq)\ |\ H^+(aq)\ |\ Pt$

$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$ Oxidation = Anode → left side of cell notation

$2e^- + 2H^+(aq) \longrightarrow H_2(g)$ Reduction = Cathode → left side of cell notation

A solid surface must be present to accept electrons ∴ an inert electrode (Pt) is used

12. A galvanic cell consists of one half-cell that contains Ag(*s*) and Ag^+(*aq*), and one half-cell that contains Pb(*s*) and Pb^{2+}(*aq*). What species are produced at the electrodes under standard conditions?

$\boxed{B}$

$Ag^+(aq) + e^- \rightarrow Ag(s)\quad E° = +0.80\ V$
$Pb^{2+}(aq) + 2\ e^- \rightarrow Pb(s)\quad E° = -0.13\ V$

Ag+ has a higher reduction potential than Pb^{2+} ∴ will be the oxidizing agent

∴ Ag+ acts as the cathode (reduction) half reaction
↓
you can also think about Ag+ has the oxidizing agent which gets reduced at the cathode

A. Ag(aq) is formed at the cathode and, Pb(s) is formed at the anode.
B. Ag(s) is formed at the cathode, and Pb^{2+}(aq) is formed at the anode.
C. Pb(s) is formed at the cathode, and Ag^+(aq) is formed at the anode.
D. Pb^{2+}(aq) is formed at the cathode, and Cu(s) is formed at the anode.

oxidizing agent

$2Ag^+(aq) + 2e^- \longrightarrow 2Ag(s)$
$Pb(s) \longrightarrow Pb^{2+} + 2e^-$

** Remember oxidizing agents get reduced & reducing agents are oxidized in a redox reaction **

13. Given the following laboratory observation, which of the following statements is **FALSE**?

$\boxed{C}$

OA

$\overset{0}{Zn}(s) + \overset{2+}{Cu}Br_2(aq) \longrightarrow \overset{0}{Cu}(s) + \overset{2+}{Zn}Br_2(aq)$
$Cu(s) + ZnBr_2(aq) \longrightarrow$ No Reaction

Cu^{2+} = oxidizing agent ∴ has a higher reduction potential ∴ undergoes reduction

A. Zn is a stronger reducing agent than Cu. true
B. Cu^{2+} is a stronger oxidizing agent than Zn^{2+}. true
C. Cu is a stronger reducing agent than Zn. False
D. The fact that Cu doesn't react with $ZnBr_2$ proves that copper attracts electrons more than does Zn. true
E. None of the above.

$Cu^{2+} + 2e^- \longrightarrow Cu(s)\quad E°_{red} = 0.34V$ oxidizing agent gets reduced
$Zn^{2+} + 2e^- \longrightarrow Zn(s)\quad E°_{red} = -0.76V$ reducing agent gets oxidized

Rewriting the equation

$Cu^{2+}(aq) + \cancel{2e^-} \longrightarrow Cu(s)$ cathode
$Zn(s) \longrightarrow Zn^{2+} + \cancel{2e^-}$ Anode

$Cu^{2+}(aq) + Zn(s) \longrightarrow Cu(s) + Zn^{2+}(aq)$ $E°cell = 0.34 - (-0.76) = 1.10V$

14. Calculate the standard cell potential for the galvanic cell reaction given below and determine whether this reaction is spontaneous or non-spontaneous under standard conditions.

A

$Cu^{2+}(aq) + 2\,Fe^{2+}(aq) \rightarrow 2\,Fe^{3+}(aq) + Cu(s)$

Split the reaction into two half cell reactions:

$Fe^{3+}(aq) + e^- \rightarrow Fe^{2+}(aq)$ $E° = +0.77$ V
$Cu^{2+}(aq) + 2\,e^- \rightarrow Cu(s)$ $E° = +0.34$ V

$Cu^{2+}_{(aq)} + 2e^- \rightarrow Cu(s)$ Reduction (Cathode)
$2Fe^{2+}(aq) \rightarrow Fe^{3+}(aq)$ Oxidation (Anode)

A. $E°$ = - 0.43 V, nonspontaneous.
B. $E°$ = - 0.43 V, spontaneous.
C. $E°$ = + 0.43 V, nonspontaneous.
D. $E°$ = + 0.43 V, spontaneous.

$E°_{cell} = E°_{(red)}$ Cathode $- E°_{(red)}$ Anode
$= 0.34 - 0.77$
$= -0.43V \rightarrow$ negative $E°$ cell value indicates that the galvanic cell reaction is non-spontaneous

15. A galvanic cell consists of a silver electrode in 1.0 mol/L solution of silver nitrate, a copper electrode in 1.0 mol/L solution of copper(II) nitrate, and a salt bridge. The spontaneous cell reaction is:

CD

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

When the two electrodes are connected by a wire, which of the following **does not** take place? Select all that apply. Insert letters without spaces in the answer box, example **ABCD.**

a. Electrons flow in the wire from the copper electrode to the silver electrode. True (in a galvanic cell, electrons flow from the anode to cathode)
b. The silver electrode increases in mass as the cell operates. True
c. There is a net movement of silver ions through the salt bridge from the silver half-cell to the copper half-cell. False (counter ions from the salt bridge to anode & cathode solutions to avoid charge build up)
d. There is a net movement of copper ions through the salt bridge from the copper half-cell to the silver half-cell. False
e. The copper electrode decreases in mass as the cell operates. True

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ Anode
$2Ag^+ + 2e^- \rightarrow 2\,Ag(s)$ Cathode

16. Determine which of the following pairs of reactants will result in a spontaneous reaction at 25°C. Please refer to the table in question 6 for additional information.

A

A. $Sn^{2+}(aq) + Zn(s)$

$2e^- + Sn^{2+} \rightarrow Sn(s)$ Reduction = cathode $E° = -0.14$
$Zn(s) \rightarrow Zn^{2+}_{(aq)} + 2e^-$ Oxidation : anode $E° = -0.76$
$E°_{cell} = -0.14 - (-0.76) = 0.62\,V$ spontaneous

B. $Cr^{3+}(aq) + Sn(s)$

$2\,Cr^{3+}_{(aq)} + 6e^- \rightarrow 2Cr(s)$ Red, $E° = -0.73V$
$3\,Sn(s) \rightarrow 3Sn^{2+}(aq) + 6e^-$ Oxid, $E° = -0.14V$
$E°_{cell} = -0.73 - (-0.14) = -0.59V$ non-spontaneous

C. $Zn(s) + K^+(aq)$

$Zn(s) \rightarrow Zn^{2+}_{(aq)} + 2e^-$ Oxidation $E°_{red} = -0.76V$
$2K^+(aq) + 2e^- \rightarrow 2K(s)$ Reduction $E°_{red} = -2.92V$
$E°_{cell} = -2.92 - (-0.76) = -2.16V$
non-spontaneous

D. $Fe(s) + Ba^{2+}(aq)$

$Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$ oxid $E° = -0.44V$
$Ba^{2+}(aq) + 2e^- \rightarrow Ba(s)$ Red $E° = -2.90V$

E. None of the above pairs will react.

$E°_{cell} = -2.90 - (-0.44) = -2.46$ non-spontaneous