KEY

RedOx Homework Booklet Name________________________________Period______

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1. When SO_4^{-2} reacts to form $S_2O_6^{-2}$, the sulphur atoms
   A. gain electrons and are ~~oxidized~~
   B. ~~lose~~ electrons and are oxidized
   C. gain electrons and are reduced
   D. ~~lose~~ electrons and are reduced

SO_4^{-2} $\quad$ $S_2O_6^{-2}$

$x + 4(-2) = -2$ $\quad$ $2x + 6(-2) = -2$

$x = +6 \longrightarrow x = +5$

2. Consider the following unbalanced redox reaction:

$$Sn^{+2} + MnO_4^{-} \rightarrow Sn^{+4} + Mn^{+2}$$

Which of the following describes the change in manganese in the MnO_4^{-}?
   A. gains electrons and is ~~oxidized~~
   B. gains electrons and is reduced
   C. ~~loses~~ electrons and is reduced
   D. ~~loses~~ electrons and is ~~oxidized~~

MnO_4^{-} $\quad$ Mn^{+2}

$x + 4(-2) = -1$

$x = +7 \longrightarrow Mn^{+2}$

3. Consider the following equation for the breathalyzer reaction:

$$CH_3CH_2OH + K_2Cr_2O_7 + H_2SO_4 \rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$$

The change in oxidation number for each carbon is equivilent to
   A. 2 electrons gained
   B. 2 electrons lost
   C. 1 electron lost
   D. 1 electron gained

$CH_3CH_2OH \longrightarrow CH_3COOH$

$2x + 6(+1) + 1(-2) = 0$ $\quad$ $2x + 4(+1) + 2(-2) = 0$

$x = -2$ $\quad$ $x = 0$

4. Which of the following gives the correct oxidation numbers for the nitrogen atoms in all three chemical species?

	N_2	Li_3N	NO_2^-
A.	~~-3~~	-3	-3
B.	0	-3	+3
C.	~~-3~~	-3	+3
D.	0	~~+3~~	-3

NO_2^-

$x + 2(-2) = -1$

$x = +3$

5. The oxidation number of manganese changes as MnO_4^- is converted to MnO_2. How many electrons are gained or lost by the manganese during the change?
   A. one electron gained
   B. 3 electrons gained
   C. 3 electrons lost
   D. one electron lost

$MnO_4^- \longrightarrow MnO_2$

$x + 4(-2) = -1$ $\quad$ $x + 2(-2) = 0$

$x = +7$ $\quad$ $x = +4$

6. A reducing agent gets oxidized
   A. gains electrons and is ~~reduced~~
   B. gains electrons and is oxidized
   C. loses electrons and is oxidized
   D. loses electrons and is ~~reduced~~

7. Which of the following describes a strong oxidizing agent? gets reduced
   A. a substance which loses electrons easily
   B. a substance which has a large increase in oxidation numbers
   C. a substance which has a small increase in oxidation numbers
   D. a substance which gains electrons easily

8. An equation for the rusting of iron is shown below:

$$4Fe + 3O_2 \rightarrow 2\,Fe_2O_3$$

   Which of the following is **false**?
   A. O_2 is the oxidizing agent. true
   B. Metallic iron is reduced to Fe^{+3}. oxidized
   C. Metallic iron is the reducing agent.
   D. This is a redox reaction.

9. Two separate reactions involved in the refining of copper ore are:

Reaction I	$2\overset{+1}{Cu}_2S + 3O_2 \rightarrow 2\overset{+1}{Cu}_2O + 2SO_2$
Reaction II	$\overset{+1}{Cu}_2S + 2\overset{+1}{Cu}_2O \rightarrow 6\overset{0}{Cu} + SO_2$

   What happens to the copper ions during this process?
   A. Copper ions are oxidized in Reaction II.
   B. Copper ions are oxidized in Reaction I.
   C. Copper ions are reduced in Reaction II.
   D. Copper ions are reduced in Reaction I.

10. Which of the following could be produced by the reduction of NO_2?
   A. NO
   B. N_2O_5
   C. HNO_3
   D. N_2O_4

11. Consider the following redox reaction:

$$3NO + SO_4^{-2} + Fe^{+3} + 2H_2O \rightarrow FeS + 3NO_3^- + 4H^+$$

Which of the following is being oxidized?

A. NO B. ~~H_2O~~ C. SO_4^{-2} reduced D. Fe^{+3} reduced

12. Consider the following redox reaction:

$$\overset{+2}{3NO} + SO_4^{-2} + Fe^{+3} + 2H_2O \rightarrow FeS + \overset{+5}{3NO_3^-} + 4H^+$$

The oxidation number for nitrogen in this reaction has

A. increased by 3
B. decreased by 2
C. increased by 2
D. decreased by 1

13. What is the oxidation number for chromium in $Cr_2O_7^{-2}$?

A. +7
B. +6
C. +14
D. +12

$2x + 7(-2) = -2$

$2x = +12$

$x = +6$

14. Consider the following unbalanced half-reaction:

$$\overset{+4}{H_2SeO_3} \rightarrow \overset{0}{Se}$$

The oxidation number of Se

A. increases as it undergoes reduction
B. decreases as it undergoes oxidation
C. decreases as it undergoes reduction
D. increases as it undergoes oxidation

15. What is the oxidation number for sulfur in $S_2O_8^{-2}$?

A. +9
B. -2
C. +7
D. +8

$2x + 8(-2) = -2$

$2x = +14$

$x = +7$

16. Consider the following equation:

$$16H^+ + 2\overset{+7}{Mn}O_4^- + 5\overset{+3}{C_2}O_4^{-2} \rightarrow 10\ CO_2 + 2\overset{+2}{Mn^{+2}} + H_2O$$

Identify the chemical species which is reduced.

A. Mn^{+2} B. H^+ C. $C_2O_4^{-2}$ D. MnO_4^-

Consider the following :

$$\overset{+5}{H_3AsO_4} + 4Zn + 8H^+ \rightarrow \overset{+3}{As_2O_3} + 4Zn^{+2} + 4H_2O$$

Which of the following statements is **true**?

A. The oxidation of hydrogen decreases by 1. ×

B. The H_3AsO_4 is the reducing agent. ×

C. The oxidation number of arsenic decreases by 2. ✓

D. The oxidation number of arsenic increases by 2.

18. What is the oxidation number of iron in Fe_3O_4?

A. +2

B. +3

C. +8/3

D. +4/3

19. What is the oxidation number of iron in FeO_2?

A. +3

B. +4

C. +2

D. +1/2

20. In which of the following chemical changes will there be an oxidation number change of +3?

A. $ClO^- \rightarrow ClO_2^-$ $+1 \rightarrow +3$

B. $Cr^{+3} \rightarrow Cr_2O_7^{-2}$ $+3 \rightarrow +6$

C. $Mn^{+2} \rightarrow MnO_4^-$ $+2 \rightarrow +7$

D. $Cr^{+3} \rightarrow Cr^{+2}$ $+3 \rightarrow +2$

21. Which of the following represents an oxidation?

A. $SO_2 + H_2O \rightarrow 2H^+ + SO_3^{-2}$ $+4 \rightarrow +4$

B. $Na^+ + Cl^- \rightarrow NaCl$ no change

C. $2SO_4^{-2} \rightarrow S_2O_8^{-2}$ $+6 \rightarrow +14$

D. $2H^+ + S \rightarrow H_2S$ $0 \rightarrow -2$

22. Which equation represents a redox reaction?

A. $NH_3 + HCl \rightarrow NH_4Cl$

B. $C + O_2 \rightarrow CO_2$

C. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

D. $2CrO_4^{-2} + 2H^+ \rightarrow Cr_2O_7^{-2} + H_2O$

gets reduced

23. Identify the oxidizing agent in the following equation:

$$Pb + PbO_2 + 4H^+ + 2\,SO_4^{-2} \rightarrow 2PbSO_4 + 2H_2O$$

Car battery

A. Pb   B. H^+ C. SO_4^{-2} D. PbO_2

24. Which of the following is a redox equation?
A. $2H_2 + O_2 \rightarrow 2H_2O$
B. $Ag(NH_3)^{2+} + 2H^+ + Cl^- \rightarrow AgCl + 2NH_4^+$
C. $Mn(OH)_2 + 2HC_2H_3O_2 \rightarrow Mn^{+2} + 2H_2O + 2C_2H_3O_2^-$
D. $Ag_2CrO_4 \rightarrow 2\,Ag^+ + CrO_4^{-2}$

25. Which of the following contains molybedenum with its highest oxidation number?
A. Mo_2S_3 +3
B. $MoCl_5$ +5
C. Mo_6Cl_{12} +2
D. MoO_4^{-2} +7

26. Which of the following skeletal half-reactions are **not** oxidations?

I. $ClO^- \rightarrow ClO_3^-$ ✓
II. $C_2H_5OH \rightarrow C_2H_4O_2$ $-2 \rightarrow 0$ ✓
III. $NO_2 \rightarrow N_2O_4$ × stays the same

A. II   B. III   C. I and II   D. I

27. Which of the following is **not** a redox equation?
A. $Cl_2 + 2NaI \rightarrow 2NaCl + I_2$ ✓
B. $Sn^{+4} + 2\,Fe^{+2} \rightarrow 2\,Fe^{+3} + Sn^{+2}$ ✓
C. $Zn + S \rightarrow ZnS$ ✓
D. $K_3PO_4 + 3\,AgNO_3 \rightarrow Ag_3PO_4 + 3KNO_3$ double replacement

28. Consider the following redox equation:

$$MnO_4^- + 5Fe^{+2} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{+3} + 4H_2O$$

Which of the following statements is **false**?
A. The equation is balanced. ✓
B. Manganese is reduced. ✓
C. Hydrogen is reduced. ×
D. Iron is oxidized. ✓

29. Consider the following unbalanced redox equation:

$$Pt + NO_3^- + Cl^- \rightarrow PtCl_6^{-2} + NO + H_2O$$

Which chemical species is reduced?

A. NO_3^- B. Pt C. Cl^- D. H_2O

30. What is the oxidation number of carbon in $Mn(CH_3COO)_3$?

A. +1
B. -1
C. +1/2
D. 0

CH_3COO^- $2x + 3(+1) + 2(-2) = -1$
$2x = 0$
$x = 0$

gets oxidized

31. Identify the reducing agent in the following equation:

$$Zn + 2MnO_2 + 2H_2O \rightarrow Zn(OH)_2 + 2MnO(OH)$$

A. H_2O B. MnO_2 C. Zn D. $Zn(OH)_2$

32. Consider the following equation:

$$Co + SO_4^{-2} + 4H^+ \rightarrow Co^{+2} + H_2SO_3 + H_2O$$

Which statement is correct?

A. The hydrogen is ~~oxidized~~ and the sulphur is reduced
B. The cobalt is oxidized and the sulphur is reduced
C. The sulphur is oxidized and the cobalt is ~~reduced~~
D. The hydrogen is reduced and the cobalt is oxidized

33. In an experiment Ag^+ reacts spontaneously with Ru but not with Pd. The relative strength of the metals from strongest to weakest reducing agent is

A. Ru > Ag > Pd
B. Pd > Ag > Ru
C. Ru > Pd > Ag
D. Ag > Ru > Pd

34. A solution containing Pd^{+2} reacts spontaneously with Ga to produce Pd and Ga^{+3}. However, a solution containing Pd^{+2} does not react with Pt. The metals, in order of increasing strength as reducing agents, are:

A. Pt < Ga < Pd
B. Ga < Pt < Pd
C. Ga < Pd < Pt
D. Pt < Pd < Ga

35. Consider the following spontaneous reactions:

$$3Cd^{+2} + 2Np \rightarrow 3\,Cd + 2Np^{+3}$$
$$Cd + Pd^{+2} \rightarrow Pd + Cd^{+2}$$
$$Np^{+3} + Ce \rightarrow Np + Ce^{+3}$$

$Cd^{+2} > Np^{+3}$

$Pd^{+2} > Cd^{+2}$

$Np^{+3} > Ce^{+2}$

Which is the strongest oxidizing agent?

A. Pd^{+2}
B. Np^{+3}
C. Ce^{+3}
D. Cd^{+2}

36. The metals Hg, Cd, and Pd reacts as follows:

$$Hg^{+2} + Pd \rightarrow Pd^{+2} + Hg$$
$$Pd^{+2} + Cd \rightarrow Cd^{+2} + Pd$$
$$Hg^{+2} + Cd \rightarrow Cd^{+2} + Hg$$

Pd > Hg — weakest

Cd > Pd

Cd > Hg

Put the metals in order from weakest reducing agents to strongest:

A. Hg < Pd < Cd
B. ~~Cd~~ < Pd < ~~Hg~~
C. Hg < Cd < ~~Pd~~
D. ~~Pd~~ < Hg < Cd

37. Consider the following experimental results:

$$Mo^{+3} + Mn \rightarrow Mn^{+2} + Mo^{+2}$$
$$Mo^{+3} + Pd \rightarrow \text{no reaction}$$
$$Ti + Mo^{+3} \rightarrow \text{no reaction}$$
$$Ti + Pd^{+2} \rightarrow Pd + Ti^{+2}$$

$Mo^{+3} > Mn^{+2}$

$Pd^{+2} > Mo^{+3}$

Use these results to complete the table of reduction half reactions below:

STRONGEST ↑ (Oxidizing Agents) — WEAKEST ↓ (Reducing Agents)

Oxidizing Agents		Reducing Agents
$Pd^{+2} + 2e^-$	⇄	Pd
$Ti^{+2} + 2e^-$	⇄	Ti
$Mo^{+3} + e^-$	⇄	Mo^{+2}
$Mn^{+2} + 2e^-$	⇄	Mn

WEAKEST (Oxidizing Agents) — STRONGEST (Reducing Agents)

38. Consider the following spontaneous reaction:

$$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2\,(aq)} + H_{2\,(g)}$$

Which of the following statements is correct?

A. Mg is a weaker reducing agent than H^+
B. Mg is a stronger reducing agent than H^+
C. Mg is a stronger reducing agent than H_2
D. Mg is a weaker reducing agent than H_2

39. A piece of Ag does not react spontaneously with 1.0 M Ti^{+2} because

A. Ag^+ is a weaker oxiding agent than Ti^{+2}
B. Ag^+ is a stronger reducing agent than Ti^{+2}
C. Ag^+ is a stronger oxidizing agent than Ti^{+2}
D. Ag^+ is a weaker reducing agent than Ti^{+2}

$Ag^+ + Ti^{+2} \leftarrow Ti + Ag^+$

40. Which of the following can act as an oxidizing agent, but **not** as a reducing agent?

A. Cl^-
B. Na^+
C. Cu^+ both
D. Cr right only

∴ is on the left but not right

41. Which of the following can act as an reducing agent, but **not** as a oxidizing agent?

A. Sn
B. Ca^{+2}
C. Br_2
D. Fe^{+2}

∴ can get oxidized but not reduced

42. Solid copper forms spontaneously in the following reaction:

$$V + Cu^{+2} \rightarrow Cu + V^{+2}$$

Based on these observations, V^{+2} is a

A. stronger reducing agent than Cu^{+2}
B. weaker oxidizing agent than Cu^{+2}
C. stronger oxidizing agent than Cu^{+2}
D. weaker reducing agent than Cu^{+2}

43. Which of the following is the strongest reducing agent?

A. H_2Te
B. H_2Se
C. H_2O
D. H_2S

look on page 8 bottom right

44. Which of the following is the strongest oxidizing agent?

A. H_2O
B. Sn^{+2}
C. I_2
D. Fe^{+2}

∴ highest up on left

45. Which of the following ions can be reduced by Pb (s) under standard conditions?

A. Cu^{+} ✓
B. Sn^{+2} ✗
C. Ca^{+2} ✗
D. Cr^{+3} ✗

∴ must be above
$Pb \rightarrow Pb^{+2} + 2e^-$

46. Which of the following is more difficult to reduce than the $H^+_{(aq)}$ ion?

A. Zn^{+2}
B. Ag^{+}
C. Cu^{+2}
D. I_2

∴ lower than $2H^+ + 2e^- \rightarrow H_{2(s)}$

47. A piece of Cu reacts spontaneously with 1.0 M Pd^{+2} because

A. Cu is a stronger reducing agent than Pd and $E^o > 0$
B. Cu is a weaker reducing agent than Pd and $E^o < 0$
C. Cu is a stronger reducing agent than Pd and $E^o < 0$
D. Cu is a weaker reducing agent than Pd and $E^o > 0$

48. Which two species will **not** react spontaneously at standard conditions?

A. Cu with Ag^{+} ✓
B. Mg with Cr^{+3} ✓
C. Ag with Zn^{+2} ✗
D. Co with Cl_2 ✓

49. Which of the following will not react spontaneously with H_2O at standard conditions?

A. Sn will not
B. Na will
C. Ca will
D. F_2 will

50. Which of the following will react spontaneously with Ag_2S at standard conditions?

A. Au
B. Co
C. Pb
D. Al

∴ must be below
$Ag_2S + 2e^- \rightarrow 2Ag + S^{-2}$

51. Which of the following represents a spontaneous redox reaction?

A. $AuCl_4^- + 3Ag \rightarrow Au + 4Cl^- + 3Ag^+$

B. $2Br^- + Mg^{+2} \rightarrow MgBr_2$

C. $Cu + Sn^{+4} \rightarrow Cu^{+2} + Sn^{+2}$

D. $Ag_2S + 2Fe^{+2} \rightarrow 2Fe^{+3} + 2Ag + S^{-2}$

52. Which of the following will react spontaneously with Br_2 but not with I_2?

A. Cr^{+2}

B. Fe^{+2}

C. Mn^{+2}

D. F^-

oxidation below $Br_2 + 2e^- \rightarrow 2Br^-$

above $I_2 + 2e^- \rightarrow 2I^-$

53. Consider the following:

$$Sn^{+4} + 2Cl^- \rightarrow Sn^{+2} + Cl_2$$

$Sn^{+4} + 2e^- \rightarrow Sn^{+2}$ $E^\circ = +0.15$

$2Cl^- \rightarrow Cl_2 + 2e^-$ $E^\circ = -1.36V$

What is true for this reaction?

A. $E^\circ_{cell} = +1.21$ V and it is spontaneous

B. $E^\circ_{cell} = -1.51$ V and it is not spontaneous

C. $E^\circ_{cell} = -1.21$ V and it is not spontaneous

D. $E^\circ_{cell} = +1.51$ V and it is spontaneous

54. Which of the following could react spontaneously with Ag metal?

oxidation $\therefore$ reduction above

$Ag \rightarrow Ag^+ + e^-$

A. Fe^{+2}

B. acidified NO_3^-

C. acidified SO_4^{-2}

D. Cl^-

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$

55. Which of the following could react spontaneously with Fe^{+2}?

A. MnO_4^- in base reduction ✗

B. Cl^- oxidation ✗

C. acidified SO_4^{-2} oxidation ✗

D. Ag^+ reduction ✓

reduction above $Fe^{+2} \rightarrow Fe^{+3} + e^-$

oxidation below $Fe^{+2} + \cancel{2e^-} \rightarrow Fe$

56. Which of the following metals can oxidized by 1.0 M Fe^{+2}?

A. Co

B. Ag

C. Cr

D. Sn

$\therefore$ must be below $Fe^{+2} + 2e^- \rightarrow Fe$

57. Consider the following half-reactions under standard conditions:

I. $ClO_2 + e^- \rightarrow ClO_2^-$

II. $PbSO_4 + 2e^- \rightarrow Pb + SO_4^{-2}$

III. $Fe^{+3} + 3e^- \rightarrow Fe$

$\therefore$ I > III $\therefore$ III > II

In an experiment when ClO_2 and Fe were combined, they reacted. In a second experiment when $PbSO_4$ and Fe were combined, there was no observable change. Which of the following shows the reduction half-reactions I, II and III in order of decreasing E°?

A. I, III, II   B. II, III, I   C. III, II, I   D. I, II, III

58. Which of the following combinations will react spontaneously?

A. Zn^{+2} + Mg ✓
B. Pb^{+2} + Ag ✗
C. Sn^{+2} + Ni^{+2} ✗
D. I_2 + Cu^{+2} ✗

59. A sample of copper is placed in $HNO_{3(aq)}$ and another sample of copper is placed in $HCl_{(aq)}$.

a. In which acid does a spontaneous redox reaction occur with the copper? HNO_3

b. Calculate the E° for the reaction that occurs. +0.62 V

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ $E^\circ = +0.96\,V$

$Cu \rightarrow Cu^{+2} + 2e^-$ $E^\circ = -0.34\,V$

$E^\circ = +0.62\,V$

60. An excess of copper solid is dropped into a solution which contains $AgNO_3$, $Fe(NO_3)_3$ and $Zn(NO_3)_2$. Write the equations for any reduction half-reactions that occur over time under standard conditions.

Nothing for NO_3^-: but reductions must be above $Cu \rightarrow Cu^{+2} + 2e^-$

$\therefore$ $Ag^+ + e^- \rightarrow Ag$ ✓

$Fe^{+3} + e^- \rightarrow Fe^{+2}$ ✓

$Zn^{+2} + 2e^- \rightarrow Zn$ ✗

61. A reaction occurs when copper metal is dropped into a solution of silver nitrate. Write the balanced formula equation and the balanced net ionic equation for this reaction.

Formula equation: $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

Net ionic equation: $Cu + 2Ag^+ \rightarrow Cu^{+2} + 2Ag$

62. Which of the following is the balanced half-reaction for

$N_2O \rightarrow NH_3OH^+$ (acidic) $4e^- + 6H^+ + H_2O + N_2O \rightarrow 2NH_3OH^+$

A. $N_2O + 2H^+ + H_2O + e^- \rightarrow NH_3OH^+$
B. $N_2O + 4H^+ + 3e^- \rightarrow NH_3OH^+$
C. $N_2O + 6H^+ + H_2O + 4e^- \rightarrow 2NH_3OH^+$
D. $N_2O + 6H^+ + H_2O \rightarrow 2NH_3OH^+ + 4e^-$

63. When the skeletal equation $S_2O_3^{-2} \rightarrow HSO_3^-$ is balanced in acidic solution H^+ and e^- will appear. Which of the following best describes the H^+ and e^- for the balanced half-reaction?

$3H_2O + S_2O_3^{-2} \rightarrow 2HSO_3^- + 4H^+ + 4e^-$

$2(+2) \rightarrow 2(+4) + 4e^-$

A. $4H^+$ on the right and $4e^-$ on the right
B. $4H^+$ on the left and $3e^-$ on the left
C. $1H^+$ on the left and $1e^-$ on the left
D. $1H^+$ on the left and $1e^-$ on the right

64. When the skeletal equation $Br_2 \rightarrow BrO_3^-$ is balanced in acidic solution, H_2O, H^+ and e^- will appear. Which of the following are the correct balancing coefficients?

$6H_2O + Br_2 \rightarrow 2BrO_3^- + 12H^+ + 10e^-$

	H_2O	H^+	e^-
A.	6	12	10
B.	3	3	2
C.	3	6	4
D.	6	12	5

65. Consider the following skeletal equation for a redox reaction in acidic solution:

$Br^- + BrO_3^- \rightarrow Br_2$ $2BrO_3^- + 12H^+ + 10e^- \rightarrow 6H_2O + Br_2$

What is the equation for the balanced reduction half-reaction

A. $2Br^- + 2e- \rightarrow Br_2$
B. $5e- + 6H+ + BrO_3^- \rightarrow Br_2 + 3H_2O$
C. $10e- + 12H^+ + 2BrO_3^- \rightarrow Br_2 + 6H_2O$
D. $2Br^- \rightarrow Br_2 + 2e-$

66. Consider the following balanced redox equation in acidic solution:

$$4OH^- + \cancel{5}\overset{H_2O}{H_2O} + 2CoCl_2 + OCl^- \rightarrow 2\,Co(OH)_3 + 5Cl^- + 4H^+ + 4OH^- / \cancel{4H_2O}$$

Which of the following describes the amounts and locations of OH^- and H_2O if the equation is balanced in basic solution?

A. $1H_2O$ on the left and 4 OH^- on the left
B. $5H_2O$ on the left and 4 OH^- on the left
C. $1H_2O$ on the left and 4 OH^- on the ~~right~~
D. $1H_2O$ on the left and no OH^-

67. Balance the following redox reaction:

$$\overset{+2}{NO} + SO_4^{-2} + \overset{+3}{Fe^{+3}} \rightarrow \overset{+3}{Fe}S + \overset{+5}{NO_3^-} \quad \text{(acidic)}$$

$$(2H_2O + \overset{+6}{NO} \rightarrow NO_3^{-2} + 4H^+ + 3e^-)3$$

$$9e^- + 8H^+ + SO_4^{-2} + Fe^{+3} \rightarrow FeS + 4H_2O$$

$$2H_2O + 3NO + SO_4^{-2} + Fe^{+3} \rightarrow FeS + 3NO_3^- + 4H^+$$

68. Balance the following redox equation:

$$SO_3^{-2} + MnO_4^- \rightarrow SO_4^{-2} + MnO_2 \quad \text{(basic)}$$

$$3(SO_3^{-2} + H_2O \rightarrow SO_4^{-2} + 2H^+ + 2e^-)$$

$$2(MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O)$$

$$3SO_3^{-2} + \cancel{3H_2O} + 2MnO_4^- + \overset{2H^+}{\cancel{8H^+}} + \cancel{6e^-} \rightarrow 3SO_4^{-2} + 2MnO_2 + \overset{H_2O}{\cancel{4H_2O}} + \cancel{6H^+} + \cancel{6e^-}$$

$$3SO_3^{-2} + 2MnO_4^- + \cancel{2H^+} \rightarrow 3SO_4^{-2} + 2MnO_2 + \cancel{H_2O}$$

$$\cancel{2}H_2O \rightarrow \cancel{2H^+} + 2OH^-$$

$$\boxed{3SO_3^{-2} + 2MnO_4^- + H_2O \rightarrow 3SO_4^{-2} + 2MnO_2 + 2OH^-}$$

69. Balance the following redox reaction: (acidic)

$$Ag_2O + Si \rightarrow Ag + SiO_3^{-2} \text{ (acidic)}$$

$\cancel{3H_2O} + Si \rightarrow SiO_3^{-2} + \cancel{6H^+} + \cancel{4e^-}$

$\cancel{4H^+} + 4e^- + 2Ag_2O \rightarrow 4Ag + 2\cancel{H_2O}$

$H_2O + Si + 2Ag_2O \rightarrow 4Ag + SiO_3^{-2} + 2H^+$

70. Balance the following redox reaction: (acidic)

$$Cr_2O_7^{-2} + I_2 \rightarrow Cr^{+3} + IO_3^- \text{ (acidic)}$$

$\cancel{18H_2O} + 3I_2 \rightarrow 6IO_3^- + \cancel{36H^+} + \cancel{30e^-}$

$\cancel{30e^-} + 70H^+ + 5Cr_2O_7^{-2} \rightarrow 10Cr^{+3} + \cancel{35H_2O}$

$34H^+ + 5Cr_2O_7^{-2} + 3I_2 \rightarrow 10Cr^{+3} + 6IO_3^- + 17H_2O$

71. Balance the following redox equation:

$$ClO_3^- + S_2O_3^{-2} \rightarrow S_4O_6^{-2} + Cl^- \quad \text{(acidic)}$$

$$\cancel{6e^-} + 6H^+ + ClO_3^- \rightarrow Cl^- + 3H_2O$$

$$6S_2O_3^{-2} \rightarrow 3S_4O_6^{-2} + \cancel{6e^-}$$

$$6H^+ + ClO_3^- + 6S_2O_3^{-2} \rightarrow 3S_4O_6^{-2} + Cl^- + 3H_2O$$

72. Balance the following skeletal redox equation in acidic solution:

$$MnO_4^- + As_2O_3 \rightarrow Mn^{+2} + H_3AsO_4 \quad \text{(acidic)}$$

$$\cancel{20e^-} + \cancel{32}H^+ + 4MnO_4^- \rightarrow 4Mn^{+2} + \cancel{16}H_2O$$

$$\cancel{25}H_2O + 5As_2O_3 \rightarrow 10H_3AsO_4 + \cancel{20}H^+ + \cancel{20e^-}$$

$$9H_2O + 12H^+ + 5As_2O_3 + 4MnO_4^- \rightarrow 4Mn^{+2} + 10H_3AsO_4$$

73. Balance the following skeletal redox equation in basic solution:

$$NO_3^- + As_2S_3 \rightarrow N_2O + HAsO_4^{-2} + SO_4^{-2} \text{ (basic)}$$

$\cancel{56}e^- + 7\cancel{0}H^+ + 14\,NO_3^- \rightarrow 7\,N_2O + 3\cancel{5}H_2O$

$4\cancel{0}H_2O + 2\,As_2S_3 \rightarrow 4\,HAsO_4^{-2} + 6\,SO_4^{-2} + \cancel{76}H^+ + \cancel{56}e^-$

$\cancel{5}H_2O + 14\,NO_3^- + 2\,As_2S_3 \rightarrow 4\,HAsO_4^{-2} + 6\,SO_4^{-2} + \cancel{6H^+} + 7N_2O$

$\cancel{6H^+} + 6\,OH^- \rightarrow \cancel{6H_2O}$

$6\,OH^- + 14\,NO_3^- + 2\,As_2S_3 \rightarrow 4\,HAsO_4^{-2} + 6\,SO_4^{-2} + H_2O + 7N_2O$

74. Balance the following redox reaction:

$$Sb + NO_3^- \rightarrow Sb_2O_5 + NO \text{ (acidic)}$$

$\cancel{15H_2O} + 6\,Sb \rightarrow 3\,Sb_2O_5 + \cancel{30H^+} + \cancel{30e^-}$

$\cancel{30e^-} + \cancel{40H^+} + 10\,NO_3^- \rightarrow 10\,NO + \cancel{20H_2O}$

$10H^+ + 6\,Sb + 10\,NO_3^- \rightarrow 10\,NO + 3\,Sb_2O_5 + 5H_2O$

75. Balance the following redox reaction:

$$As_4 + CrO_4^{-2} \rightarrow Cr^{+3} + H_2AsO_4^- \quad \text{(basic)}$$

$$\cancel{48H_2O} + 3As_4 \rightarrow 12H_2AsO_4^- + \cancel{72H^+} + \cancel{60e^-}$$

$$\cancel{60e^-} + \cancel{160H^+} + 20CrO_4^{-2} \rightarrow 20Cr^{+3} + \cancel{80H_2O}$$

$$\cancel{88H^+} + 3As_4 + 20CrO_4^{-2} \rightarrow 20Cr^{+3} + 12H_2AsO_4^- + \cancel{32H_2O}$$

$$\cancel{88H_2O} \rightarrow \cancel{88H^+} + 88OH^-$$

$$56H_2O + 3As_4 + 20CrO_4^{-2} \rightarrow 20Cr^{+3} + 12H_2AsO_4^- + 88OH^-$$

76. Balance the following redox reaction:

$$H_3As + H_2AsO_4^- \rightarrow As_2O_3 \quad \text{(basic)}$$

$$\cancel{3H_2O} + 2H_3As \rightarrow As_2O_3 + \cancel{12H^+} + \cancel{12e^-}$$

$$\cancel{12e^-} + \cancel{18H^+} + 6H_2AsO_4^- \rightarrow 3As_2O_3 + \cancel{15H_2O}$$

$$\cancel{6H^+} + 2H_3As + 6H_2AsO_4^- \rightarrow 4As_2O_3 + \overset{6H_2O}{\cancel{12H_2O}}$$

$$\cancel{6H_2O} \rightarrow \cancel{6H^+} + 6OH^-$$

$$2H_3As + 6H_2AsO_4^- \rightarrow 4As_2O_3 + 6H_2O + 6OH^-$$

77. Balance the following equation.

$Cr_2O_7^{-2} + C_2O_4^{-2} \rightarrow Cr^{+3} + CO_2$ (acidic)

$\cancel{6e^-} + 14H^+ + Cr_2O_7^{-2} \rightarrow 2Cr^{+3} + 7H_2O$

$3C_2O_4 \rightarrow 6CO_2 + \cancel{6e^-}$

$14H^+ + Cr_2O_7^{-2} + 3C_2O_4^{-2} \rightarrow 6CO_2 + 2Cr^{+3} + 7H_2O$

78. Balance the following equation.

$Cr_2O_7^{-2} + C_2H_5OH \rightarrow Cr^{+3} + CH_3COOH$ (acidic)

$\cancel{12e^-} + \overset{16H^+}{\cancel{28}H^+} + 2Cr_2O_7^{-2} \rightarrow 4Cr^{+3} + \overset{11H_2O}{\cancel{14H_2O}}$

$\cancel{3H_2O} + 3C_2H_5OH \rightarrow 3CH_3COOH + \cancel{12H^+} + \cancel{12e^-}$

$16H^+ + 2Cr_2O_7^{-2} + 3C_2H_5OH \rightarrow 3CH_3COOH + 4Cr^{+3} + 11H_2O$

79. Balance the following redox reaction:

$$H_2S + CrO_4^{-2} \rightarrow S_8 + Cr^{+3} \quad \text{(acidic)}$$

$3(\ 8H_2S \longrightarrow S_8 + 16H^+ + 16e^-)$

$16(\ 3e^- + 8H^+ + CrO_4^{-2} \longrightarrow Cr^{+3} + 4H_2O)$

$24H_2S + 16CrO_4^{-2} + \cancel{48e^-} + \overset{80H^+}{\cancel{128H^+}} \longrightarrow 3S_8 + 16Cr^{+3} + 64H_2O + \cancel{48H^+} + \cancel{48e^-}$

$$\boxed{24H_2S + 16CrO_4^{-2} + 80H^+ \longrightarrow 3S_8 + 16Cr^{+3} + 64H_2O}$$

80. Balance the following redox reaction in basic solution:

$$ClO_3^- + N_2H_4 \rightarrow NO_3^- + Cl^- \quad \text{(basic)}$$

$\cancel{42e^-} + \cancel{42H^+} + 7ClO_3^- \longrightarrow 7Cl^- + \overset{3H_2O}{\cancel{21H_2O}}$

$\cancel{18H_2O} + 3N_2H_4 \longrightarrow 6NO_3^- + \overset{6H^+}{\cancel{48H^+}} + \cancel{42e^-}$

$7ClO_3^- + 3N_2H_4 \longrightarrow 7Cl^- + 6NO_3^- + \cancel{3H_2O} + \cancel{6H^+}$

$\cancel{6H^+} + 6OH^- \longleftrightarrow \overset{9H_2O}{\cancel{6H_2O}}$

$7ClO_3^- + 3N_2H_4 + 6OH^- \longrightarrow 7Cl^- + 6NO_3^- + 9H_2O$

81. Consider the following skeletal redox equation for a reaction in basic solution:

$$Zn + NO_3^- \rightarrow ZnO_2^{-2} + NH_3 \quad \text{(basic)}$$

Which of the following best represents the reduction half-reaction occurring in this solution?

A. $6H_2O + NO_3^- + 8e^- \rightarrow NH_3 + 9OH^-$
B. $3H_2O + NO_3^- + 5e^- \rightarrow NH_3 + 6OH^-$
C. $4OH^- + Zn + 2e^- \rightarrow ZnO_2^{-2} + 2H_2O$
D. $9H^+ + NO_3^- + 8e^- \rightarrow NH_3 + 3H_2O$

$8e^- + 9H^+ + NO_3^- \rightarrow NH_3 + \cancel{3H_2O}$
$9OH^- + \cancel{9H_2O}$ $6H_2O$ $+ 9OH^-$

$8e^- + 6H_2O + NO_3^- \rightarrow NH_3 + 9OH^-$

82. When a piece of Ag is placed in 1.0 M $NiCl_2$

A. the $[Ag^+]$ decreases
B. no changes occur
C. the $[Ni^{+2}]$ decreases
D. the $[Cl^-]$ increases

$Ag \rightarrow Ag^+ + e^-$ ∴ need reduction above

$Ni^{+2} + 2e^- \rightarrow Ni$ Not above

No reaction

83. When a peice of Cu is placed in 1.0 M $AgNO_3$

A. the $[Cu^{+2}]$ increases
B. the $[NO_3^-]$ decreases
C. no change occurs
D. the $[Ag^+]$ increases

$Cu \rightarrow Cu^{+2} + 2e^-$ ∴ need reduction above

$Ag^+ + e^- \rightarrow Ag$ ✓ ∴ $[Cu^{+2}]\uparrow$ and $[Ag^+]\downarrow$

84. When a piece of Cu is placed in 1.0 M HNO_3

A. the $[NO_3^-]$ decreases
B. the $[Cu^{+2}]$ decreases
C. no change occurs
D. the $[H^+]$ increases

$Cu \rightarrow Cu^{+2} + 2e^-$ ↑

$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$ ✓
↓ ↓

85. Consider the following equation for the breathalyzer reaction:

$$CH_3CH_2OH + K_2Cr_2O_7 + H_2SO_4 \rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$$

The balanced redox reaction would be:

A. $3CH_3CH_2OH + K_2Cr_2O_7 + 4H_2SO_4 \rightarrow 3CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$
B. $CH_3CH_2OH + K_2Cr_2O_7 + 4H_2SO_4 \rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + 8H_2O$
C. $3CH_3CH_2OH + 2K_2Cr_2O_7 + 8H_2SO_4 \rightarrow 3CH_3COOH + 2Cr_2(SO_4)_3 + 2K_2SO_4 + 11H_2O$
D. $CH_3CH_2OH + K_2Cr_2O_7 + H_2SO_4 \rightarrow CH_3COOH + Cr_2(SO_4)_3 + K_2SO_4 + H_2O$

See question 78

86. Nitric oxide (NO) can be prepared by the oxidation of Cu with NO_3^- in acidic solution. Which of the following equations correctly describes this process?

A. $Cu + NO_3^- + 3H^+ \rightarrow Cu^{+2} + NO + H_2O$ Not balanced.

B. $3Cu + 2NO_3^- + 8H^+ \rightarrow 3Cu^{+2} + 2NO + 4H_2O$

C. $Cu + NO_3^- + 4H^+ \rightarrow Cu^{+2} + NO_2 + 2H_2O$

D. $Cu + 2NO_3^- + 4H^+ \rightarrow Cu^{+2} + 2NO_2 + 2H_2O$

$NO_3^- \rightarrow NO$

$(4H^+ + 3e^- + NO_3^- \rightarrow NO + 2H_2O)\,2$

$(Cu \rightarrow Cu^{+2} + 2e^-)\,3$

87. Which of the following could be used to determine the $[Fe^{+2}]$ by a redox reaction?

A. Cl^-

B. MnO_4^- (acidified)

C. Cu^{+2}

D. I_2

$MnO_4^- + 8H^+ + 5e^- \leftrightarrow Mn^{+2} + 4H_2O$

$Fe^{+2} \leftrightarrow Fe^{+3} + e^-$

88. A 2.000 g strip of cobalt metal is suspended in 100.0 mL of 0.20 M $AgNO_3$ and a reaction occurs. When the reaction is complete, there is an excess of cobalt. The excess cobalt is removed from the solution, washed and dried and its mass is found to be 1.411 g.

a. Using the table of Standard Reduction Potentials of Half-cells, write the balanced net ionic equation for the redox reaction.

$$2Ag^+ + Co \rightarrow Co^{+2} + 2Ag$$

b. Using the experimental data, calculate the moles of Co and Ag+ reacting, and show how these values support the balanced equation.

$$0.1000\,L \times \frac{0.20\text{ mol }Ag^+}{1\,L} = 0.0200\text{ mol }Ag^+$$

$$(\underbrace{2.000 - 1.411}_{0.589})\text{ g of Co} \times \frac{1\text{ mol Co}}{58.9\text{ g}} = 0.0100\text{ mol Co}$$

$\therefore$ 2 : 1 Mol. ratio

89. Which of the following could be used to determine the acidified [BrO_3^-] by a redox reaction?

A. MnO_4^- (acidified) reduction
B. NO_3- (acidified) reduction
C. Cu^{+2} reduction
D. I^- oxidation.

$2\ BrO_3^- + 12H^+ + 10e^- \rightarrow Br_2 + 6H_2O$

$2I^- \rightarrow I_2 + 2e^-$

90. Which of the following could be titrated using acidified MnO_4^- ions?

A. IO_3^- reduction ×
B. SO_4^{-2} oxidation above ×
C. Na^+ reduction ×
D. H_2O_2 $\rightarrow O_2 + 2H^+ + 2e^-$ ✓

$MnO_4^- + 8H^+ + 5e^- \leftrightarrow Mn^{+2} + 4H_2O$

91. Acidified potassium permanganate ($KMnO_4$) solution is often used in redox titrations. Permanganate reacts with Sn^{+2} as follows:

$$2MnO_4^- + 5Sn^{+2} + 16H^+ \rightarrow 2Mn^{+2} + 5Sn^{+4} + 8H_2O$$

A 10.0 mL solution containing Sn^{+2} is titrated with 19.3 mL of 0.10 M $KMnO_4$. What is the [Sn^{+2}]?

$$0.0193L \times \frac{0.10\,mol\ \cancel{MnO_4^-}}{1L} \times \frac{5\ mol\ Sn^{+2}}{2\ mol\ \cancel{MnO_4}} \times \frac{1}{0.010L} =$$

$0.4825\ M$ or $\boxed{0.48M}$

92. The titation of a 25.0 mL $SnCl_2$ sample, in acidic solution, requires 14.4 mL of 0.030 M $K_2Cr_2O_7$. The balanced equation for the reaction is shown below:

$$Cr_2O_7^{-2} + 14H^+ + 3Sn^{+2} \rightarrow 3Sn^{+4} + 2Cr^{+3} + 7H_2O$$

What is the number of moles of $SnCl_2$ in the original sample?

A. 4.3 x 10^{-4} moles
B. 1.3 x 10^{-3} moles
C. 5.2 x 10^{-2} moles
D. 1.4 x 10^{-4} moles

$$0.0144\cancel{L} \times \frac{0.030\ mol\ Cr_2O_7^{-2}}{\cancel{L}} \times \frac{3\ mol\ Sn^{+2}}{1\ mol\ Cr_2O_7^{-2}} = 1.296 \times 10^{-3}\ mol\ Sn^{+2}$$

93. A 10.0 mL water sample was analyzed for $[Fe^{+2}]$ using a redox titration with acidified $KMnO_4$. The equation for the reaction is:

$$MnO_4^- + 5Fe^{+2} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{+3} + 4H_2O$$

A 10.0 mL sample was titrated with 12.5 mL of 0.10 M $KMnO_4$ solution. What is the $[Fe^{+2}]$ in the water sample?

A. 0.13 M
B. 0.28 M
C. 0.63 M
D. 0.025 M

$0.0125\,\cancel{L}\ MnO_4^- \times \frac{0.10\ mol\ MnO_4}{1\ \cancel{L}} \times \frac{5\ mol\ Fe^{+2}}{1\ mol\ MnO_4^-} \times \frac{1}{0.010\ L} = 0.625\ M$

94. Consider the following redox reaction in acidic solution:

$$2\ \overset{+7}{KMnO_4} + 5\overset{+2}{H_2O_2} + 3H_2SO_4 \rightarrow K_2SO_4 + 2\overset{+2}{MnSO_4} + 8H_2O + 5O_2$$

a. Balance the above redox reaction.

b. The above reaction was used for a redox titration. At the equivilence point 5.684×10^{-4} mol $KMnO_4$ was required to titrate 5.00 mL of H_2O_2 solution. Calculate $[H_2O_2]$

$5.684 \times 10^{-4}\ mol\ MnO_4^- \times \frac{5\ mol\ H_2O_2}{2\ mol\ MnO_4^-} \times \frac{1}{0.05000\ L} = 0.2842\ M$

$\boxed{0.284\ M}$

95. A titration is performed to determine the $[Fe^{+2}]$ in 25.00 mL of an $FeSO_4$ solution. It requires 22.52 mL of 0.015 M $KMnO_4$ to reach the equivilence point in which Mn^{+2} and Fe^{+3} are produced.

a. balance the redox reaction:

$$MnO_4^- + Fe^{+2} \rightarrow Mn^{+2} + Fe^{+3} \text{ (acidic)}$$

$$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$$

$$5(Fe^{+2} \rightarrow Fe^{+3} + e^-)$$

$$MnO_4^- + 5Fe^{+2} + 8H^+ \rightarrow Mn^{+2} + 5Fe^{+3} + 4H_2O$$

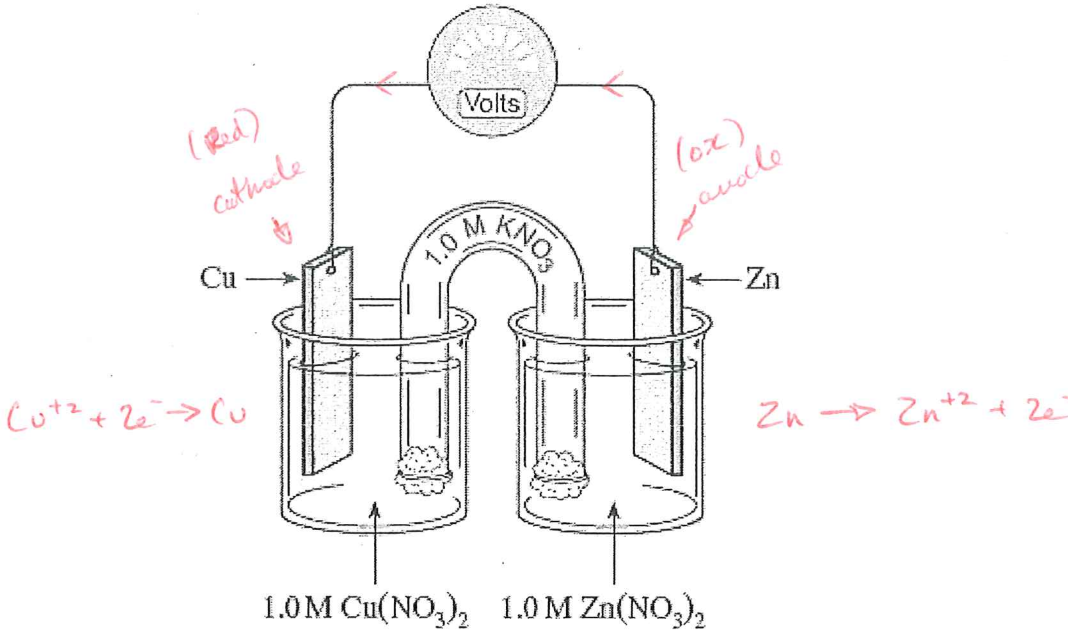
b. Calculate the $[Fe^{+2}]$

$$0.0225\,\cancel{L} \times \frac{0.015\,\cancel{mol\ MnO_4^-}}{\cancel{L}} \times \frac{5\,mol\ Fe^{+2}}{\cancel{1\,mol\ MnO_4^-}} \times \frac{1}{0.02500\,L} =$$

$$= 0.6756\,M\ Fe^{+2}$$

$$\boxed{= 0.068\,M\ Fe^{+2}}$$

96. Consider the following cell for the next three questions:



The [Cu^{+2}] in the copper half-cell will

A. decrease as Cu^{+2} gains electrons and is reduced.
B. increase as Cu loses electrons and is oxidized.
C. decrease as Cu^{+2} gains electrons and is oxidized.
D. increase as Cu loses electrons and is reduced.

97. The [Zn^{+2}] in the zinc half-cell will

A. decrease as Zn^{+2} gains electrons and is oxidized.
B. increase as Zn loses electrons and is oxidized.
C. increase as Zn loses electrons and is reduced.
D. decrease as Zn^{+2} gains electrons and is reduced.

98. What will happen to the electrodes as the cell operates?

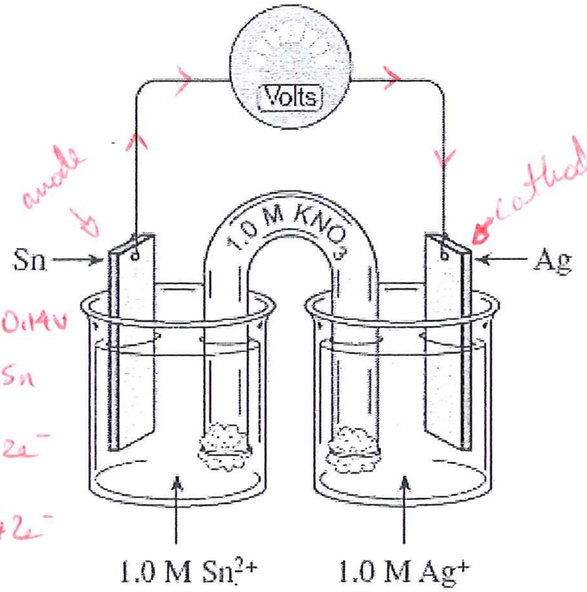
	Copper electrode	Zinc electrode
A.	decreases	increases
B.	increases	increases
C.	increases	decreases
D.	decreases	decreases

99. Consider the following cell for the next three questions:

flip for oxidation

anode

cathode



$E^\circ = -0.14V$

✗ $Sn^{+2} + \cancel{2e^-} \rightarrow Sn$

✓ $Sn \rightarrow Sn^{+2} + 2e^-$

✗ $Sn^{+2} \rightarrow \cancel{Sn^{+4}} + 2e^-$

$Ag^+ + e^- \rightarrow Ag$ $\quad E^\circ = +0.80\ V$

$+0.80V$
$+0.14V$
$E^\circ\ +0.94V$ Spontaneous

What is the overall reaction:

A. $2Ag + Sn \rightarrow Sn^{+2} + 2Ag^+$
B. $2Ag^+ + Sn \rightarrow Sn^{+2} + 2Ag$
C. $2Ag^+ + Sn^{+2} \rightarrow Sn^{+2} + 2Ag^+$
D. $2Ag + Sn^{+2} \rightarrow Sn + 2Ag^+$

100. Calculate the cell's voltage

A. $E^o = +0.94$ V B. $E^o = -0.76$ V C. $E^o = +0.76$ V D. $E^o = +1.74$ V

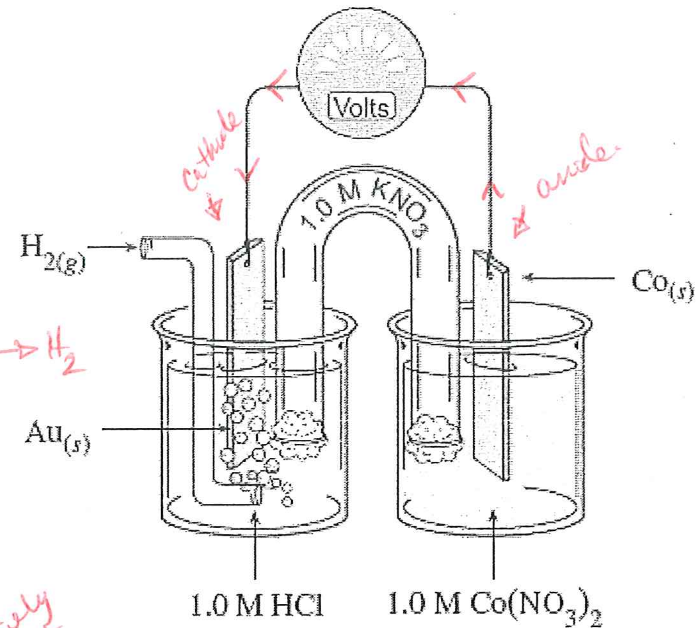
~~101. Calculate the cell's voltage~~

~~A. $E^o = +0.76$ V B. $E^o = -0.76$ V C. $E^o = +1.74$ V D. $E^o = +0.94$ V~~

102. Which way will the electrons flow, and why will the electrons go in that direction

A. Electrons will travel from the Sn to the Ag electrodes because Ag is the anode.
B. Electrons will travel from the Ag to the Sn electrodes because Sn is the anode.
C. Electrons will travel from the Ag to the Sn electrodes because Ag is the anode.
D. Electrons will travel from the Sn to the Ag electrodes because Sn is the anode.

103. Consider the diagram below for the next three questions:



$2H^+ + 2e^- \rightarrow H_2$

$Co \rightarrow Co^{+2} + 2e^-$

Gold relatively inert

Identify the overall cell reaction

A. $Au^{+3} + Co^{+2} \rightarrow Au + Co$

B. $Co^{+2} + H_2 \rightarrow 2H^+ + Co$

C. $2Au^{+3} + 3Co \rightarrow 2Au + 3Co^{+2}$

D. $2H^+ + Co \rightarrow Co^{+2} + H_2$

104. Which direction will the anions and cations travel?

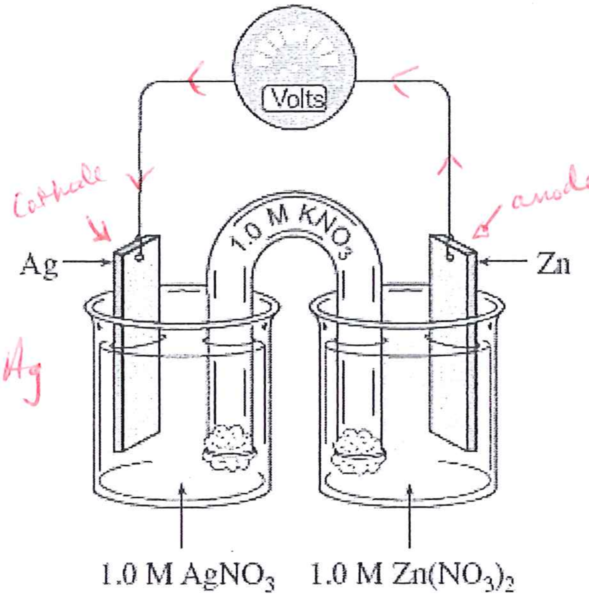
	Anions (anode)	Cations (cathode)
A.	to the Co cell	to the Co cell
B.	to the Au cell	to the Au cell
C.	to the Co cell	to the Au cell
D.	to the Au cell	to the Co cell

105. What will happen to the mass of the Co in the Co cell and what will happen to the [Cl-] in the Au cell?

	Mass of Co	[Cl-]
A.	decreases	decreases
B.	increases	increases
C.	increases	decreases
D.	decreases	increases

Cl⁻ migrate to salt bridge.

106. Consider the following cell for the next three questions



$Ag^+ + e^- \to Ag$

$Zn \to Zn^{+2} + 2e^-$

What is the equation for the half-reaction that occurs at the cathode?

A. $Zn \to Zn^{+2} + 2e^-$

B. $Ag^+ + e^- \to Ag$

C. $Zn^{+2} + 2e^- \to Zn$

D. $Ag \to Ag^+ + e^-$

107. What will happen to the $[Ag^+]$ and the $[Zn^{+2}]$ as the cell operates?

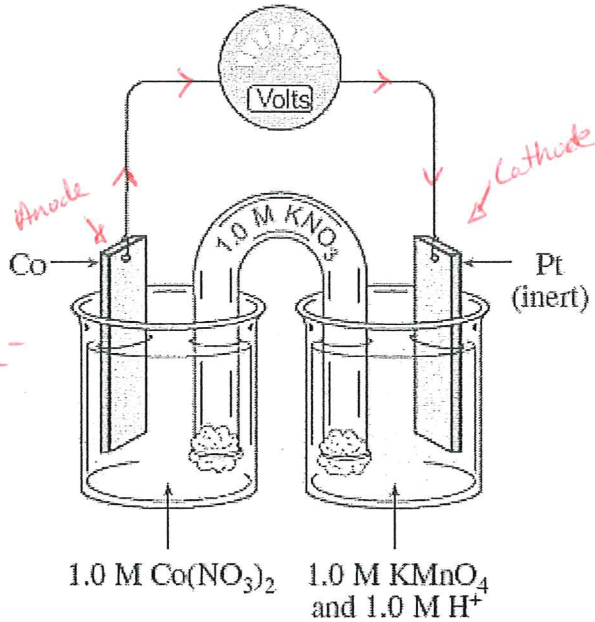
	$[Ag^+]$	$[Zn^{+2}]$
A.	increases	decreases
B.	decreases	increases
C.	increases	increases
D.	decreases	decreases

108. At $\boxed{\text{equilibrium}}$, what will be the voltage?

A. E^o = -1.56 V B. E^o = +0.04 V C. E^o = 0.0 V D. E^o = +1.56 V

*

109. Consider the following cellfor the next three questions



$Co \rightarrow Co^{+2} + 2e^-$

$K^+ + e^- \rightarrow K$

OR

$2H^+ + 2e^- \rightarrow H_2$

OR

✓ $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$

Identify the anode reaction for the cell shown in the diagram.

A. $Co^{+2} + 2e^- \rightarrow Co$

B. $Co \rightarrow Co^{+2} + 2e^-$

C. $MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{+2} + 4H_2O$

D. $H_2 \rightarrow 2H^+ + 2e^-$

110. What is the function of the salt bridge in this electrochemical cell?

A. It allows the anode to become positively charged.

B. It maintains electrical neutrality in each half cell.

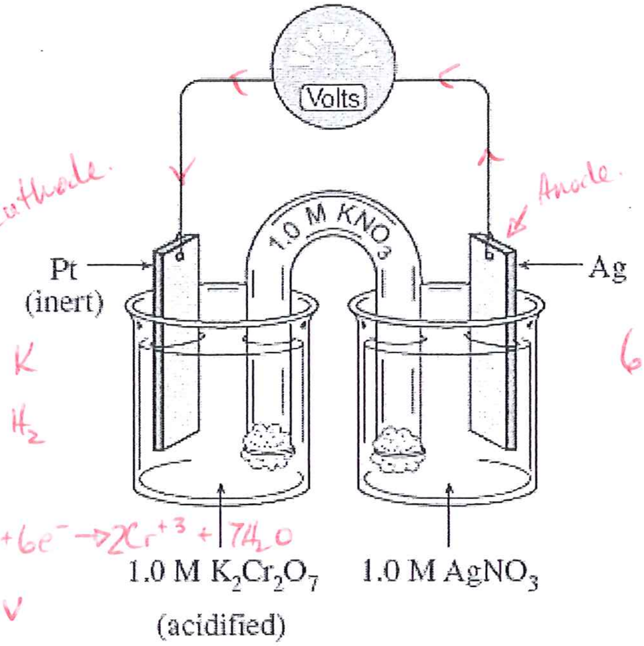
C. It allows the cathode to become negatively charged.

D. It provides a path for electrons

111. What will happen to the mass of the electrodes in this cell?

	Anode electrode	Cathode electrode (inert)
A.	stays the same	stays the same
B.	decreases	stays the same
C.	decreases	increases
D.	stays the same	increases

112. Consider the following electrochemical cell for the next three questions



Cathode.

Anode.

$K^+ + e^- \rightarrow K$

or

$2H^+ + 2e^- \rightarrow H_2$

or

✓ $Cr_2O_7^{-2} + 14H^+ + 6e^- \rightarrow 2Cr^{+3} + 7H_2O$

$E^\circ = +1.23\ V$

$6(Ag^+ + e^- \rightarrow Ag)$ $E^\circ = +0.80$

(reverse) oxidation. $\boxed{E^\circ = -0.80}$

Which of the following represents the overall cell reaction?

A. $Cr_2O_7^{-2} + 14H^+ + 9Ag \rightarrow 9Ag^+ + Cr^{+3} + 7H_2O$

B. $Cr_2O_7^{-2} + H^+ + Ag \rightarrow Ag^+ + Cr^{+3} + H_2O$

C. $Cr_2O_7^{-2} + 14H^+ + 6Ag^+ \rightarrow 6Ag + 2Cr^{+3} + 7H_2O$

D. $Cr_2O_7^{-2} + 14H^+ + 6Ag \rightarrow 6Ag^+ + 2Cr^{+3} + 7H_2O$

113. What would be the voltage for the above cell?

A. $E^o = 0.0\ V$ B. $E^o = +2.03\ V$ C. $E^o = +0.43\ V$ D. $E^o = -1.17\ V$

114. What will happen to the $[Cr_2O_7^{-2}]$ and $[Ag^+]$ as the cell operates?

	$[Cr_2O_7^{-2}]$	$[Ag^+]$
A.	decreases	decreases
B.	increases	increases
C.	increases	decreases
D.	decreases	increases

115. What happens to the pH at each electrode?

	pH at anode	pH at cathode
A.	stays the same	increases
B.	stays the same	decreases
C.	increases	increases
D.	increases	decreases

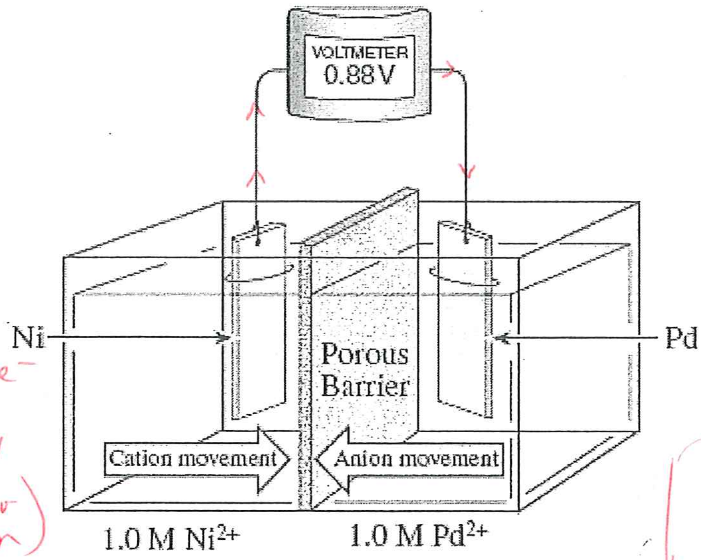
116. Consider the following diagram:

Anode

$Ni \rightarrow Ni^{+2} + 2e^-$

$E^\circ = +0.26\ V$ (Reverse for oxidation)

$E^\circ = +0.26\ V$



Cathode

$Pd^{+2} + 2e^- \rightarrow Pd$

Anions go to anode
Cations go to cathode

What is the half-cell reaction at the anode? (oxidation)

A. $Pd \rightarrow Pd^{+2} + 2e^-$
B. $Ni^{+2} + 2e^- \rightarrow Ni$
C. $Pd^{+2} + 2e^- \rightarrow Pd$
D. $Ni \rightarrow Ni^{+2} + 2e^-$

$+0.88 = x + (+0.26)$

117. Calculate the E° value for the following half-reaction: $Pd^{+2} + 2e^- \rightarrow Pd$

A. E° = -0.62 V B. E° = +1.14 V C. E° = -1.14 V D. E° = + 0.62 V

118. What will happen to the $[Ni^{+2}]$ and in what direction will the electrons travel?

	$[Ni^{+2}]$	direction
A.	increases	to the cathode
B.	~~decreases~~	~~to the anode~~
C.	~~decreases~~	to the cathode
D.	increases	~~to the anode~~

119. What is the half-cell reaction at the cathode?

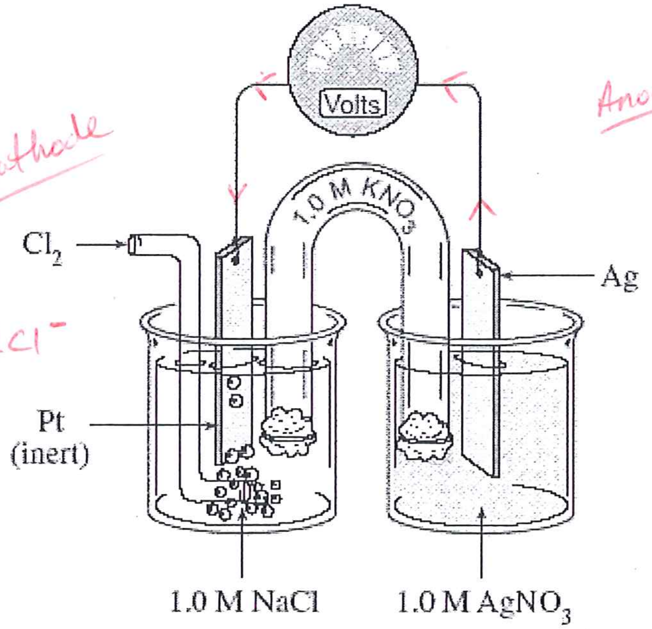
A. $Pd^{+2} + 2e^- \rightarrow Pd$
B. $Ni^{+2} + 2e^- \rightarrow Ni$
C. $Pd \rightarrow Pd^{+2} + 2e^-$
D. $Ni \rightarrow Ni^{+2} + 2e^-$

120. Consider the following cell:

Cathode

Anode

$Cl_2 + 2e^- \rightarrow 2Cl^-$



$Ag \rightarrow Ag^+ + e^-$

Which of the following represents the anode half-cell reaction?

A. $Ag \rightarrow Ag^+ + e^-$
B. $Cl_2 + 2e^- \rightarrow 2Cl^-$
C. $2Cl^- \rightarrow Cl_2 + 2e^-$
D. $Ag^+ + e^- \rightarrow Ag$

121. Which of the following represents the cathode half-cell reaction?

A. $2Cl^- \rightarrow Cl_2 + 2e^-$
B. $Cl_2 + 2e^- \rightarrow 2Cl^-$
C. $Ag^+ + e^- \rightarrow Ag$
D. $Ag \rightarrow Ag^+ + e^-$