Chemistry 12

Electrochemistry Practice Test

Name: Key
Date:
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I. Multiple Choice

A 1. When an electrode loses mass, it also:

A. loses electrons
B. acts as an oxidizing agent
C. becomes reduced
D. decreases in oxidation number

D 2. At standard conditions, Fe^{+2} reacts spontaneously with

A. I_2
B. Br^-
C. Co
D. Ag^+

3. Explain your answer to the question above:

- Fe^{2+} is an OA & RA
- Spontaneous = OA is above RA
- Ag^+ is the only RA above Fe^{2+} (acting as OA)

C 4. Which of the following half-reactions is balanced?

A. $SO_4^{-2} + H_2O \rightarrow SO_3^{-2} + 2H^+ + 2e^-$
B. $SO_4^{-2} + H_2O + 2e^- \rightarrow SO_3^{-2} + 2H^+$
C. $SO_4^{-2} + 2H^+ + 2e^- \rightarrow SO_3^{-2} + H_2O$
D. $SO_4^{-2} + 2H^+ \rightarrow SO_3^{-2} + H_2O + 2e^-$

C 5. During a redox reaction, the oxidizing agent:

A. reduces other species
B. increases in oxidation number
C. gains electrons
D. becomes oxidized

C 6. For a given redox rxn, the oxidation # of tin changed from +2 to +4. As a result, tin:

A. lost 2 electrons and was reduced
B. gained 2 electrons and was reduced
C. lost 2 electrons and was oxidized
D. gained 2 electrons and was oxidized

D 7. In which of the following compounds does carbon have an oxidation number of -2?

A. CO
B. CH_2O
C. CO_2
D. CH_3OH

B 8. Which of the following equations represents a redox reaction?

A. $ZnCl_2 \rightarrow Zn^{2+} + 2Cl^-$
B. $Zn + Br_2 \rightarrow ZnBr_2$
C. $H_2CO_3 \rightarrow H_2O + CO_2$
D. $2NaI + Pb(NO_3)_2 \rightarrow PbI_2 + NaNO_3$

$\overset{0}{Zn} + \overset{0}{Br_2} \rightarrow \overset{+2}{Zn}\overset{-1}{Br_2}$

A 9. Consider the following reaction:

$$SO_4^{-2} + \overset{-1}{8I^-} + 8H^+ \rightarrow S^{-2} + \overset{0}{4I_2} + 4H_2O$$

The reducing agent is

A. I^-
B. S in SO_4^{2-}
C. H+
D. O in SO_4^{2-}

C 10. When $\overset{+6}{Mn}O_4^{-2}$ undergoes oxidation, it may form:

A. MnO Mn = +2
B. MnO_3 Mn = +6
C. MnO_4^- Mn = +7
D. Mn_2O_3 Mn = +3

11. Explain your answer to the question above:

In MnO_4^{-2}, Mn = +6
↳ Oxidation means loss of e^- and ox # ↑

B 12. Consider the following reaction:

$$\overset{0}{3I_2} + 3H_2O \rightarrow 6H^+ + \overset{-1}{5I^-} + \overset{+5}{IO_3^-}$$

In this reaction, the I_2 atoms undergo:

A. oxidation only
B. both oxidation and reduction
C. reduction only
D. neither oxidation nor reduction

B 13. In an electrochemical cell, electrons flow from the

A. anode to the cathode through the salt bridge
B. anode to cathode through the external circuit
C. cathode to the anode through the salt bridge
D. cathode to anode through the external circuit

14. Explain your answer to the question above:

Anode = oxidation (loss of e^-)
↳ e^- travel through wire to
Cathode = reduction (gain of e^-)

B 15. When molten aluminum oxide is electrolyzed, the cathode reaction is:

A. $Al \rightarrow Al^{3+} + 3e^-$
B. $Al^{3+} + 3e^- \rightarrow Al$
C. $O_2 + 4e \rightarrow 2\,O^{2-}$
D. $2\,O^{2-} \rightarrow O_2 + 4e^-$

Al^{3+} OA reduction cathode

O^{2-} RA oxidation anode

B 16. Gold is found in nature in its pure form because:

A. Au^{3+} is a strong reducing agent
B. Au^{3+} a strong oxidizing agent
C. Au is a strong reducing agent
D. Au is a strong oxidizing agent

17. Explain your answer to the question above:

Au^{3+} is a strong OA which means it has a strong tendency to gain e^- and become pure $Au_{(s)}$

_____ 18. The electrolysis of $Na_2SO_{4\,(aq)}$ would produce this gas at the anode.

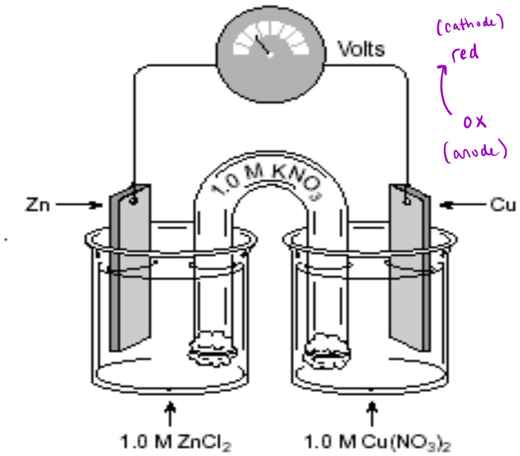
A. Oxygen
B. Hydrogen
C. Water vapour
D. Sulfur dioxide

$\cancel{Na^+}$, SO_4^{2-} , H_2O stronger RA than

Anode / RA

$2H_2O \rightarrow O_{2(g)} + 4H^+ + 4e^-$

Use the following cell diagram for questions 19 and 20.



(cathode) red

ox (anode)

$Cu^{2+} + 2e^- \rightarrow Cu$ $E° = +0.34$

$Zn \rightarrow Zn^{2+} + 2e^-$ $E° = +0.76V$

$E°_{total} = +1.10V$

C 19. In the above electrochemical cell,

A. the mass of the anode increases and the mass of the cathode increases.
B. the mass of the anode decreases and the mass of the cathode decreases.
C. the mass of the anode decreases and the mass of the cathode increases.
D. the mass of the anode increases and the mass of the cathode decreases.

A 20. In the above electrochemical cell,

A. The anode is Zn and the cathode is Cu
B. The anode is Cu and the cathode is Zn
C. The anode is Zn^{2+} and the cathode is Cu^{2+}
D. The anode is Cu^{2+} and the cathode is Zn^{2+}

D 21. In the operating electrochemical cell above, the voltage produced is:

A. -1.10 V
B. -0. 42 V
C. 0.00 V
D. +1.10 V

22. Explain your answer to the question above:

$Cu^{2+} + 2e^- \rightarrow Cu$ $E° = +0.34$

$Zn \rightarrow Zn^{2+} + 2e^-$ $E° = +0.76V$

$E°_{total} = +1.10V$

II. Problems

1) For each of the following compounds, identify the oxidation number of the atom(s) indicated.

a) Mg_2TiO_4 Mg= +2 Ti= +4 O= -2

b) $K_2Cr_2O_7$ K= +1 Cr= +6 O= -2

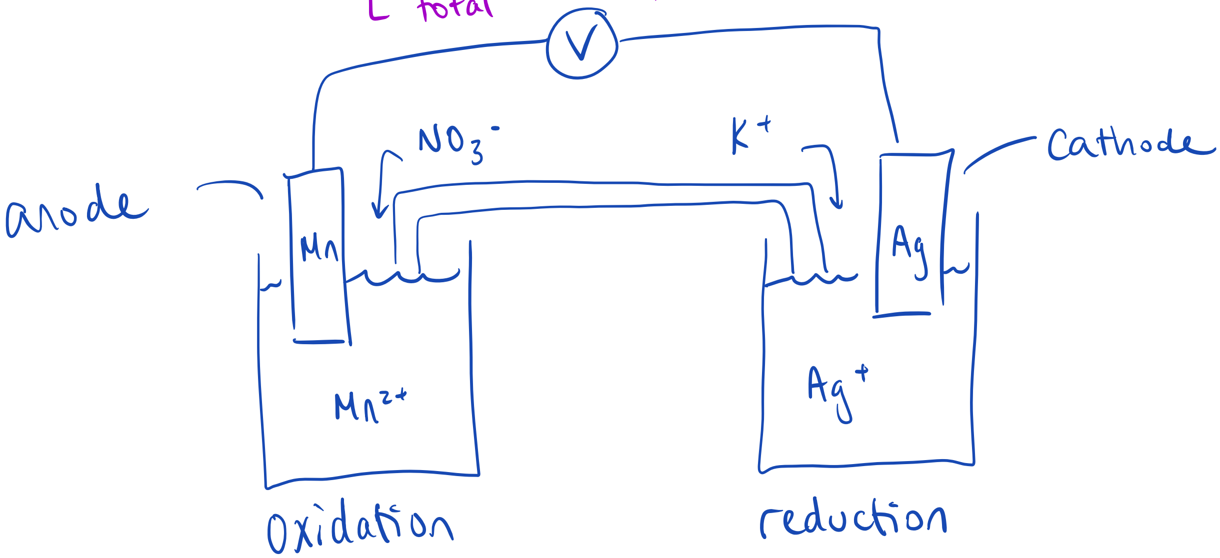
$2K^{+} + Cr_2O_7^{2-}$

c) MnO_4^- Mn= +7 O= -2

RA OA OA RA

2) A Mn/Mn^{2+} and Ag^+/Ag electrochemical cell is set up at standard conditions. Draw the electrochemical cell for this particular reaction. Label all parts of the cell, including the voltage produced.

$E^\circ_{total} = +1.19V + 0.80V = +1.99V$



$Mn \rightarrow Mn^{2+} + 2e^-$ $Ag^+ + e^- \rightarrow Ag$

3) Balance the following redox reactions.

a) Mn^{+2} + BiO_3^- → MnO_4^- + Bi^{+3} (acidic)

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

$$(2e^- + 6H^+ + BiO_3^- \rightarrow Bi^{3+} + 3H_2O) \times 5$$

$$\cancel{8H_2O} + 2Mn^{2+} \rightarrow 2MnO_4^- + \cancel{16H^+} + \cancel{10e^-}$$

$$\cancel{10e^-} + \underset{14}{\cancel{30}}H^+ + 5BiO_3^- \rightarrow 5Bi^{3+} + \underset{7}{\cancel{15}}H_2O$$

$$2Mn^{2+} + 14H^+ + 5BiO_3^- \rightarrow 2MnO_4^- + 5Bi^{3+} + 7H_2O$$

a) Sb_2S_3 + NO_3^- → NO_2 + SO_4^{2-} + Sb_2O_5 (basic)

$$\cancel{17H_2O} + Sb_2S_3 \rightarrow \underset{=}{3}SO_4^{2-} + Sb_2O_5 + \cancel{34H^+} + \cancel{28e^-}$$

-6 $\cancel{0}$ $+34$

$$\cancel{(e^- + 2H^+ + NO_3^- \rightarrow NO_2 + H_2O) \times 28}$$

$$\cancel{28e^-} + \underset{22}{\cancel{56}}H^+ + 28NO_3^- \rightarrow 28NO_2 + \underset{11}{\cancel{28}}H_2O$$

$$Sb_2S_3 + \cancel{22H^+} + 28NO_3^- \rightarrow 3SO_4^{2-} + Sb_2O_5 + 28NO_2 + \cancel{11H_2O}$$

$$\cancel{+22OH^-} \qquad + 22OH^-$$

$$\underset{11}{\cancel{22}}H_2O$$

$$Sb_2S_3 + 11H_2O + 28NO_3^- \rightarrow 3SO_4^{2-} + Sb_2O_5 + 28NO_2 + 22OH^-$$

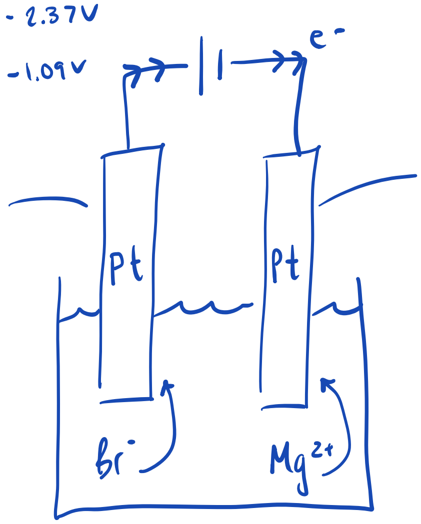
4) For each of the following, draw the electrolytic cell, including the half-reactions occurring within it:

a) Platinum electrodes in molten $MgBr_2$

$E^\circ_{total} = -3.46V$

$Mg^{2+} \rightarrow$ OA $-2.37V$

$Br^- \rightarrow$ RA $-1.09V$



anode
oxidation

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

Cathode
reduction

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

b) Copper electrodes in NaCl $_{(aq)}$

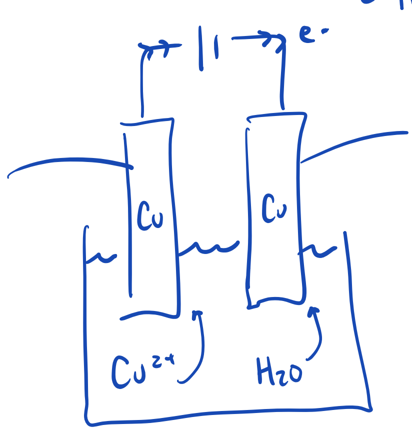
$Cu \rightarrow$ RA $-0.34V$

$Na^+ \rightarrow$ ~~OA~~

$Cl^- \rightarrow$ ~~RA~~

$H_2O \rightarrow$ ~~RA~~ or OA $-0.41V$

$E^\circ_{total} = -0.75V$



anode
oxidation

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

cathode
reduction

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

5. SRP Table Ranking

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

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

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

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

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