Electrochemistry Concepts 1

1. Write a brief definition for each term below.

Reduction	
Oxidizing Agent	

2. Assign **oxidation states** (oxidation numbers) to the atoms in each of the species below.

   a. $S_2O_3^{2-}$ b) S_8 c) $AgNO_3$ d) $HAsO_4^{2-}$

3. Identify the substances being oxidized and reduced in the reaction below. Also identify the **oxidizing agent** and **reducing agent**.

$$8H^+ + Cr_2O_7^{2-} + 3\ NO_2^- \rightarrow 2\ Cr^{3+} + 3\ NO_3^- + 4\ H_2O$$

   Reduced: Reducing Agent:

   Oxidized: Oxidizing Agent:

4. A 25.0 mL sample of 0.025-M KNO_2 was titrated with solution of potassium dichromate, $K_2Cr_2O_7$. If it required 17.45 mL to reach the equivalence point, what was the concentration of the potassium dichromate solution? Refer to the balance equation in question 3 to help answer this.

5. Balance the equation below in ACIDIC solution.

$$H_2O_2 \quad + \quad Cr_2O_7^{2-} \quad \rightarrow \quad O_2 \quad + \quad Cr^{3+}$$

6. A 2.00-g sample of impure $K_2Cr_2O_7$ was titrated using hydrogen peroxide in acidic solution as described by the reaction in question 8. If the titration required 27.38 mL of 0.0500-M H_2O_2, what was the mass-percent of $K_2Cr_2O_7$ in the impure sample?

7. Balance the following redox reaction in BASIC solution.

$$IO_3^- \quad + \quad Re \quad \rightarrow \quad ReO_4^- \quad + \quad IO^-$$

Name: ______________________________ Slot: ____

Electrochemistry Concepts 2

1) Assign **oxidation states** to each atom.

   a) Ag_2SO_4 b) ClO_2 c) H_3PO_4 d) S_8

2) Identify the **oxidizing** and **reducing agents** in the reaction below.

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

3) Label each process as **oxidation** or **reduction**.

   a) $As_2O_3 \rightarrow H_3AsO_4$ b) $Cl_2 \rightarrow OCl^-$ c) $NO_2^- \rightarrow NH_3$

4) Balance the following reaction in ACIDIC solution:

$$CH_3OH + Cr_2O_7^{2-} \rightarrow CH_2O + Cr^{3+}$$

5) A 25.00-mL solution containing methanol (CH_3OH) required 15.73 mL of 0.126-M $K_2Cr_2O_7$ for titration. Calculate the concentration of the methanol solution.

Name: ______________________________ Slot: ____

6) A solution of triioide, $I_3^-(aq)$, can be **standardized** by titration with $As_4O_6(aq)$ in <u>acidic</u> solution. The titration of 0.1021-g of $As_4O_6(s)$ dissolved in 30.00 mL of water requires 36.55 mL of $I_3^-(aq)$. Calculate the concentration of the $I_3^-(aq)$ solution.

$$As_4O_6(s) + I_3^-(aq) \rightarrow As_4O_{10}(s) + I^-(aq)$$

7) Give a definition for each term below.

Oxidation	
Reducing Agent	

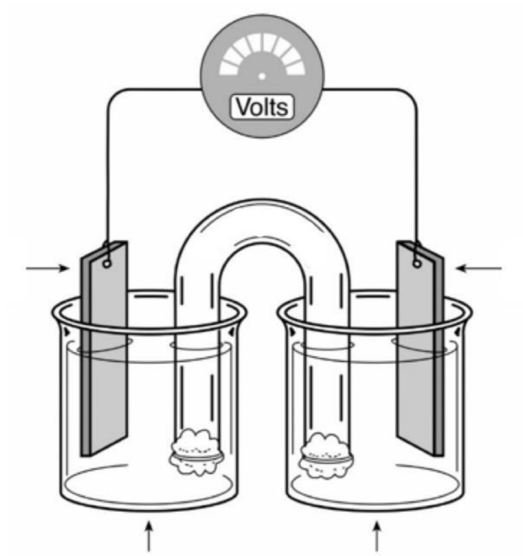
8) Circle the species in which <u>nitrogen</u> has the GREATEST oxidation state.

NO_2^- N_2O_4 NO_3^- NH_4^+ N_2

Electrochemistry Concepts 3

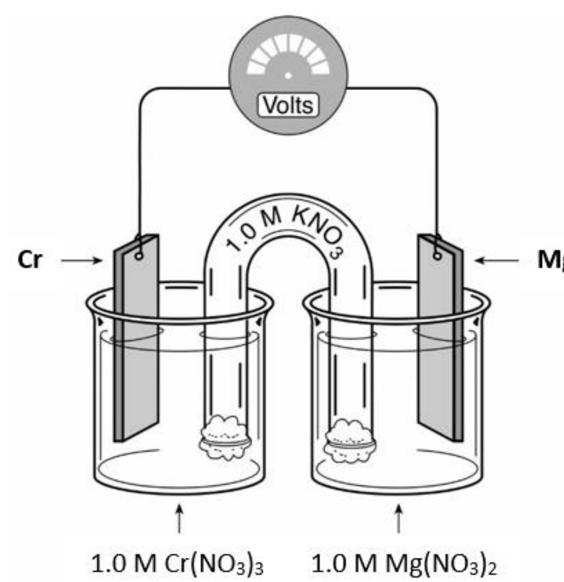
1. Sketch and **fully label** the **voltaic cell** described by this **line notation**. Include the half reaction at each electrode, the overall reaction and the cell potential calculation.

$$Al|Al^{3+}\ (1M)\ ||\ Sn^{2+}\ (1\ M)\ |Sn$$



Overall Reaction	
Cell Potential Calculation	

2. Describe the standard cell below using **line notation**. Write the overall reaction for the cell and calculate its **cell potential**.



3. Which of the following is TRUE regarding the cell in question 2?
   a. The magnesium electrode gets heavier over time.
   b. The $[Cr^{3+}]$ will increase over time.
   c. Cations will flow from the salt bridge to the chromium half-cell.
   d. Electrons are flowing in the external circuit from Cr to Mg.

4. Which of the following changes would affect the cell potential in a voltaic cell in question 2? Circle all that apply.
   a. Double the volume of the solution in the beaker for a half-cell.
   b. The atmospheric pressure changes in the room.
   c. The voltaic cell is placed in a refrigerator.
   d. Water is added to the chromium half cell.
   e. The magnesium electrode is replaced with one half its size.

5. Which of the following species is **most easily reduced**?
   a. Cu^{2+}
   b. Au^{3+}
   c. Zn
   d. Ag^{+}

6. Which of the following species is the **strongest reducing agent**?
   a. Cl^{-}
   b. Sn^{4+}
   c. Na
   d. Na^{+}

7. Which of the following would be capable of reducing Pb^{2+} to Pb?
   a. Br_2
   b. Cl^{-}
   c. Zn
   d. Mg^{2+}

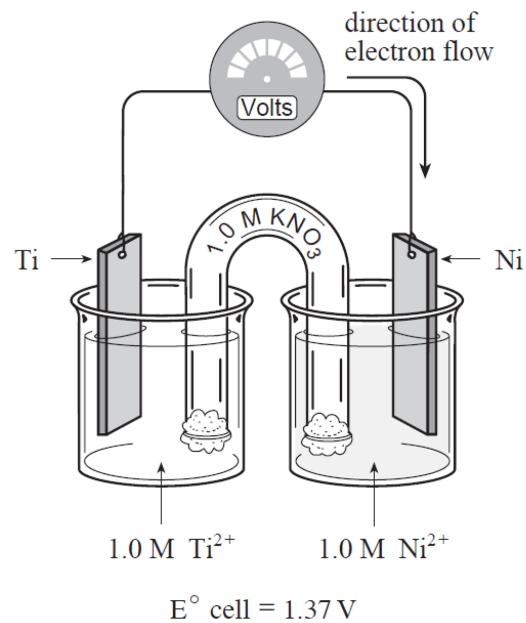
8. Which of the following could oxidize Fe^{2+} to Fe^{3+}?
   a. Cl_2
   b. Zn^{2+}
   c. Zn
   d. K^{+}

9. Which is the **strongest oxidizing agent**?
   a. Mg
   b. Cl_2
   c. Mg^{2+}
   d. Cl^{-}

10. In which cases below will a **spontaneous redox reaction** occur? Circle all that apply.
    a. A strip of nickel metal is put into a solution of zinc sulfate.
    b. A zinc metal wire is put into in a solution of silver nitrate.
    c. A copper metal coin is put into a beaker containing liquid bromine.
    d. A strip of gold metal is put into a beaker containing liquid bromine
    e. A silver ring is dropped into a copper(II) sulfate solution.
    f. A piece of lead is dropped into a beaker with acidified potassium dichromate solution.

Electrochemistry Concepts 4

1) Consider the **voltaic cell** shown here at 25°C.



   a) Which **electrode** is the **anode** and which is the **cathode**? Explain. Label each electrode with a + or – as it would be labeled <u>outside</u>.

   b) Write the **half-reaction** at each electrode.
      i) **Anode**:

      ii) **Cathode**:

   c) What happens to the **mass** of each electrode over time?
      i) Titanium:

      ii) Nickel:

   d) What happens to the **concentration** of metal cation in each cell over time?
      i) $[Ti^{2+}]$

      ii) $[Ni^{2+}]$

   e) Write the **overall reaction** in the cell.

   f) Use the information in the diagram as well as the table of standard reduction potentials in your data booklet to calculate the **standard reduction potential of Ti^{2+}**.

   g) What direction will the K^+ ions flow from the **salt bridge**? Explain.

   h) Describe the cell with **line notation**.

2) Draw and label a **standard voltaic cell** using Pb/Pb^{2+} and Mg/Mg^{2+} half-cells at 25°C.

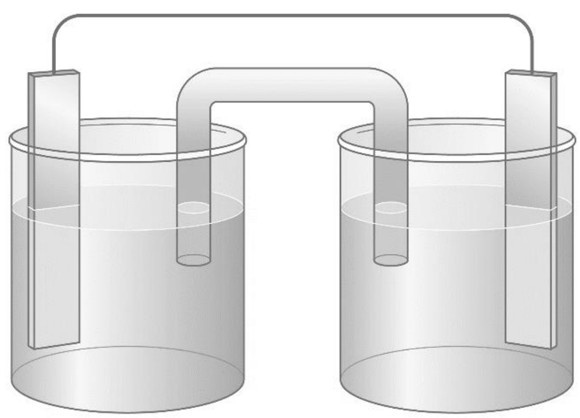
Be sure to include:

- ○ Solutions & metals
- ○ Anode & Cathode
- ○ Oxidation & Reduction
- ○ Electron direction
- ○ Salt for salt bridge
- ○ Movement of cations & anions from salt bridge
- ○ Internal charges on electrodes
- ○ External charge labels
- ○ Half reactions with $E°_{red}$ & $E°_{ox}$
- ○ Overall reaction
- ○ Cell potential

3) Describe the cell above using **line notation**.

4) Draw and completely label the following **electrochemical cell** at 25°C.

$Mn\ |\ Mn^{2+}(1M)\ ||\ Ag^{+}(1M)\ |\ Ag$

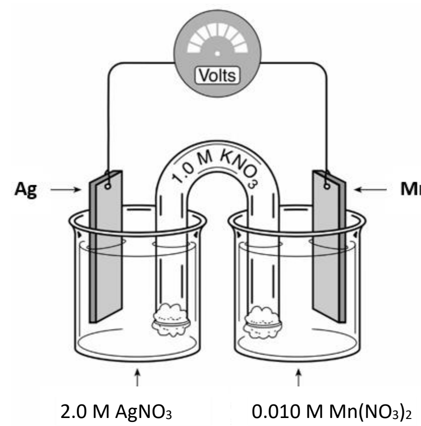


Electrochemistry Concepts 5

1. Describe **standard conditions** for electrochemical cells.

Temperature	Concentrations	Partial Pressures

2. Consider the electrochemical cell shown here, operating at 25°C.



   a. Explain why this cell is non-standard.

   b. Write the half-reaction for each electrode and then write the overall balanced equation for the cell.

   Anode:

   Cathode:

   Overall:

   c. Calculate the **standard cell potential**, **$E°_{cell}$**.

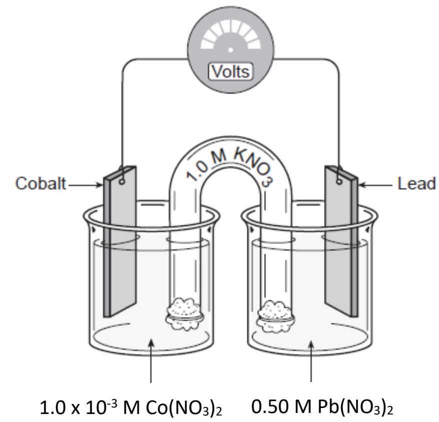
   d. Calculate the reaction quotient, Q, for this cell. What would Q be for a standard cell?

   e. Write the Nernst equation. Based on your value of Q, will this non-standard cell have a cell potential greater than, less than or equal to $E°_{cell}$? Explain briefly.

   f. Use the **Nernst equation** to calculate the cell potential, **E_{cell}**.

3. Consider the cell shown here, operating at 25°C.

   a. Calculate the reaction quotient, Q, for the cell shown here.



   b. Predict whether this cell would have a potential greater than, less than or equal to its standard cell potential.

   c. Calculate the cell potential, E_{cell}, for this cell.

4. Calculate the **equilibrium constant, Kc**, for each reaction redox below at 298K. Classify each reaction as *spontaneous* (product-favoured) or *non-spontaneous* (reactant-favoured).

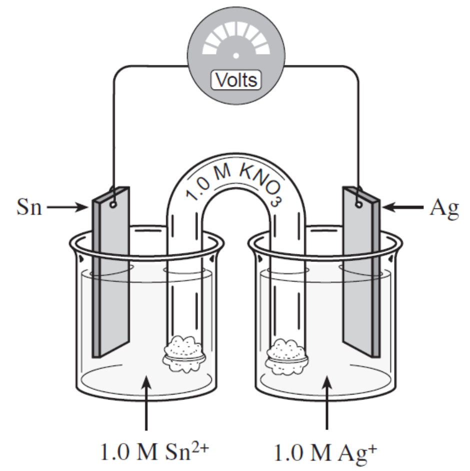
   a. $MnO_2 + 2\ Cr^{3+} + 4\ H^+ \rightarrow 2\ Cr^{3+} + Mn^{2+} + 2\ H_2O$

   b. $2\ Fe^{3+} + 2\ Br^- \rightarrow 2\ Fe^{2+} + Br_2$

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

Electrochemistry Concepts 6

1) Consider the **voltaic (galvanic) cell** shown here.



   a) Which electrode is the **anode**?

   b) Which electrode will <u>decrease</u> in mass?

   c) Into which half-cell will the K^+ ions flow from the **salt bridge**?

   d) Describe the cell with **line notation**.

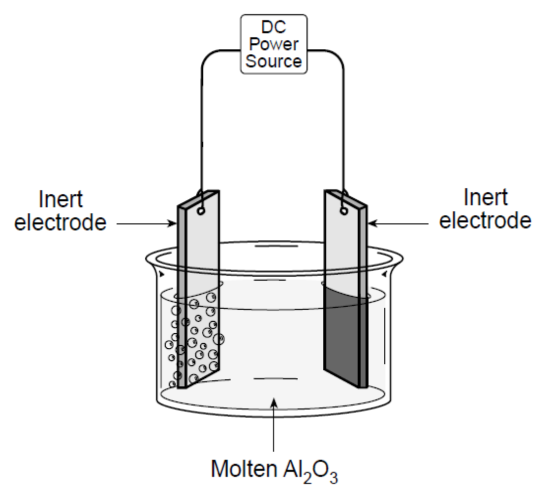
   e) What is the balanced equation for the overall reaction?

   f) Calculate the **cell potential**.

2) Circle the best answer in each case below.

a) The strongest oxidizing agent:	Zn^{2+}	I_2	Au^{3+}	F^-
b) The most easily reduced:	Ag^+	Br^-	Zn	Ca^{2+}
c) The strongest reducing agent:	Ni	Mg^{2+}	Fe^{2+}	Cl^-
d) Able to oxidize Pb	Br_2	Ni^{2+}	Al^{3+}	Ag
e) Able to reduce Sn^{4+}	Cl^-	Zn^{2+}	Mg	Au

3) Consider the **electrolytic cell** shown here.



   a) What is the product forming at the electrode on the left?

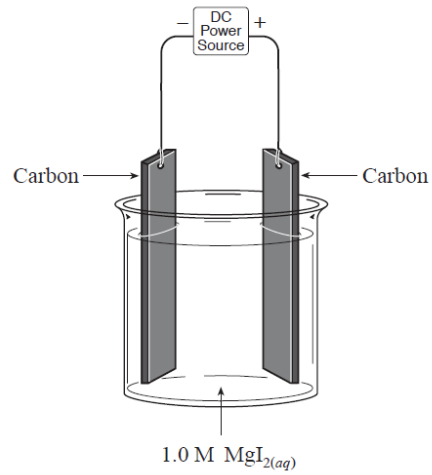
   b) Write a balanced half-reaction for the reaction occurring at the electrode on the right.

   c) Which electrode is the **anode**?

   d) Which electrode is connected to the power source's **negative** terminal?

4) What are the two possible half-reactions that could occur at the **ANODE** during the electrolysis of **aqueous** calcium bromide, $CaBr_2$ (aq)? What product will actually form?

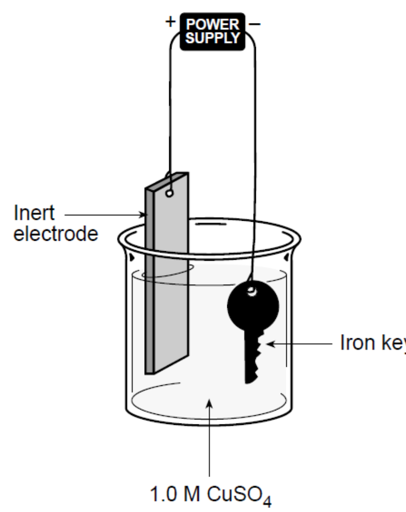
5) Consider the following electrolytic cell.



The cathode reaction is

A. $2I^- \rightarrow I_2 + 2e^-$
B. $Mg^{2+} + 2e^- \rightarrow Mg$
C. $H_2O \rightarrow \frac{1}{2}O_2 + 2H^+ + 2e^-$
D. $2H_2O + 2e^- \rightarrow H_2 + 2OH^-$

6) The electroplating shown here was allowed to run for 2.00 h. The power source provided a constant current of 10.0 A. What mass of copper will be plated on the key?



7) A student sets up an electroplating experiment by running a 20.0-A current through a concentrated solution of gold(III) nitrate, $Au(NO_3)_3$. If she wants to plate 1.00 g of gold metal, how long will her experiment need to run?

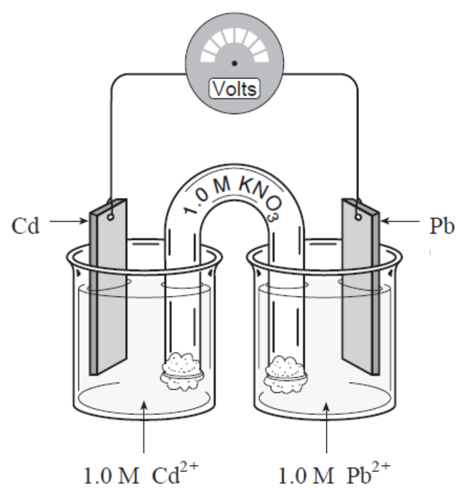
Electrochemistry Concepts 7

1) A new fertilizer, Kilzemall, contains iron(II) ammonium sulfate hexahydrate – $Fe(NH_4)_2(SO_4)_2{\cdot}6H_2O$ as a source of iron. A 6.500-g sample of Kilzemall is dissolved to make 250.0 mL of solution with dilute sulfuric acid. A 25.00-mL aliquot of this solution was titrated with 0.0100-M $KMnO_4$ and it required 23.48 mL to reach the equivalence point.

   a) Write a balanced chemical equation for the reaction between Fe^{2+} and MnO_4^- in ACIDIC solution.

   b) Based on your equation, calculate the mass of iron in the 25.00-mL aliquot.

   c) Calculate the % Fe by mass in the Kilzemall fertilizer.



2) Consider the **voltaic (galvanic) cell** shown here. As the cell operates, the lead electrode **<u>increases</u>** in mass. The cell's potential under standard conditions is 0.27 V.

   a) Write the half-reaction occurring at each electrode.

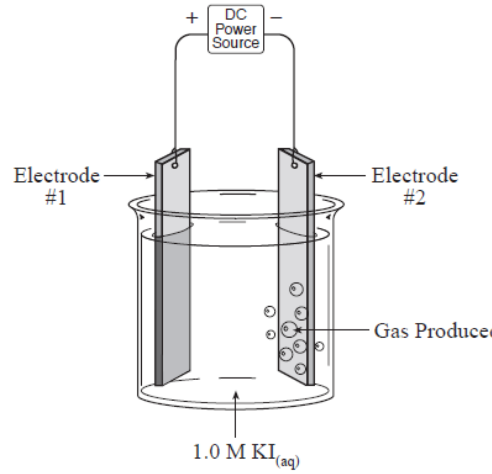
      Cathode:

      Anode:

   b) Write the <u>overall</u> balanced equation for the cell.

   c) Calculate the **standard reduction potential** for the cadmium(II) cation, Cd^{2+}.

3) Consider the **electrolytic cell** shown here. The DC power source produced a constant current of 2.50 A and operates for 30.0 min.



a) Write the balanced half reaction that is occurring at each electrode.

i) Electrode #1

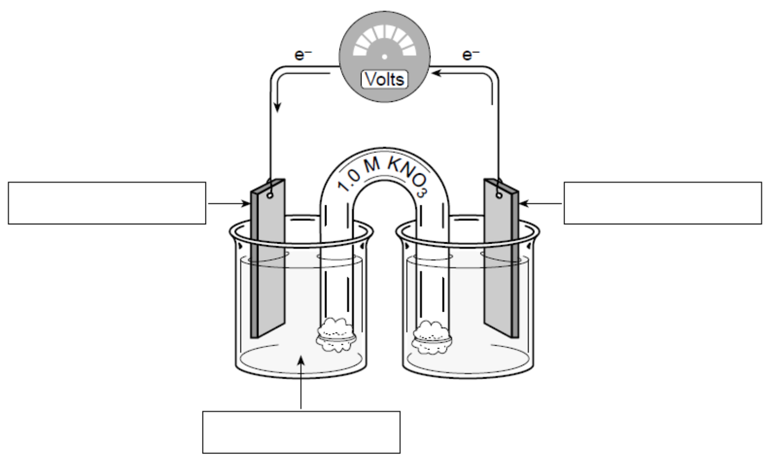
ii) Electrode #2

b) Calculate the mass of solid that forms at Electrode #1.

c) Calculate the number of moles of gas produced at Electrode #2.

4) From the list below, select materials capable of producing the greatest voltage and label the diagram.

- Silver, Aluminum and Nickel electrodes
- 1.0 M solutions of $AgNO_3$, $Al(NO_3)_3$ and $Ni(NO_3)_2$.



5) Use the laboratory observations below to complete the table of reduction half-reactions.

$Ce^{4+} + Pd \rightarrow Pd^{2+} + Ce^{3+}$

$In^{3+} + Cd \rightarrow$ no reaction

$Pd^{2+} + In^{2+} \rightarrow In^{3+} + Pd$

$Cd^{2+} + Pd \rightarrow$ no reaction

STRONGEST ↑ WEAKEST

Oxidizing Agents		Reducing Agents
	⇄	
	⇄	
	⇄	
	⇄	

WEAKEST ↓ STRONGEST

Corrosion and its Prevention

https://bit.ly/2RYUrVI and https://bit.ly/2RWzvhY

1) What is meant by the term corrosion? What term is used to refer specifically to corrosion of iron?

2) Consider iron, aluminum and zinc metal.
   a) Write their oxidation half-reactions and include the oxidation potential for each. List the three reactions in order from highest oxidation potential at the top to lowest at the bottom.

Oxidation Half-reactions for Fe, Al and Zn	$E°_{oxidation}$

   b) Zinc and aluminum are more easily oxidized than iron. Why do objects made from zinc and aluminum not experience the same extent of corrosion we see on iron/steel objects? Talk about the oxides of all three metals in your answer.

3) How are food cans made of steel prevented from corroding?

4) Write the half reaction at each of the following regions of iron's surface when corrosion is occurring. Include the potential for each reaction.
   a) Anode Region:
   b) Cathode Region when neutral or basic water is present:

5) What is the chemical formula for the final product in "rusting"?

6) Write the overall reaction with $E°_{cell}$ for the oxidation of iron when …
   a) Basic or neutral water is present

   b) Acidic water is present

7) Why does rusting occur faster when salts or other electrolytes are present?

8) What is meant by galvanic corrosion?

9) Explain how painting an iron or steel object … or greasing a bike chain … can prevent corrosion.

10) Briefly explain what is meant by “cathodic protection” by use of a “sacrificial anode”?

11) Why are metals like sodium or potassium not used for cathodic protection?

12) What is “galvanized steel”?

13) Stainless steel doesn’t rust easily. Explain why.

14) Briefly explain how an underground steel pipe can be cathodically protected using a rectifier.

Electrochemistry Multiple Choice Practice (No Electrolysis)

1. Identify the oxidizing agent in the following equation:

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

A. H^+

B. Pb

C. PbO_2

D. SO_4^{2-}

2. Which of the following is a redox equation?

A. $2H_2 + O_2 \rightarrow 2H_2O$

B. $Ag_2CrO_4 \rightarrow 2Ag^+ + CrO_4^{2-}$

C. $Ag(NH_3)_2^+ + 2H^+ + Cl^- \rightarrow AgCl + 2NH_4^+$

D. $Mn(OH)_2 + 2HC_2H_3O_2 \rightarrow Mn^{2+} + 2H_2O + 2C_2H_3O_2^-$

3. Which of the following contains molybdenum with its highest oxidation number?

A. $MoCl_5$

B. Mo_2S_3

C. MoO_4^{2-}

D. Mo_6Cl_{12}

4. Which of the following combinations will react spontaneously?

A. $I_2 + Cu^{2+}$

B. $Pb^{2+} + Ag$

C. $Zn^{2+} + Mg$

D. $Sn^{2+} + Ni^{2+}$

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

I.	$ClO^- \rightarrow ClO_3^-$
II.	$C_2H_5OH \rightarrow C_2H_4O_2$
III.	$NO_2 \rightarrow N_2O_4$

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

6. Consider the following equation:

$$16H^+ + 2MnO_4^- + 5C_2O_4^{2-} \rightarrow 10CO_2 + 2Mn^{2+} + 8H_2O$$

Identify the chemical species which is reduced.

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

7. Which of the following describes a strong oxidizing agent?

A. a substance which loses electrons readily
B. a substance which gains electrons readily
C. a substance which has a large increase in oxidation number
D. a substance which has a small increase in oxidation number

8. Consider the following spontaneous reactions:

$$3Cd^{2+} + 2Np \rightarrow 3Cd + 2Np^{3+}$$

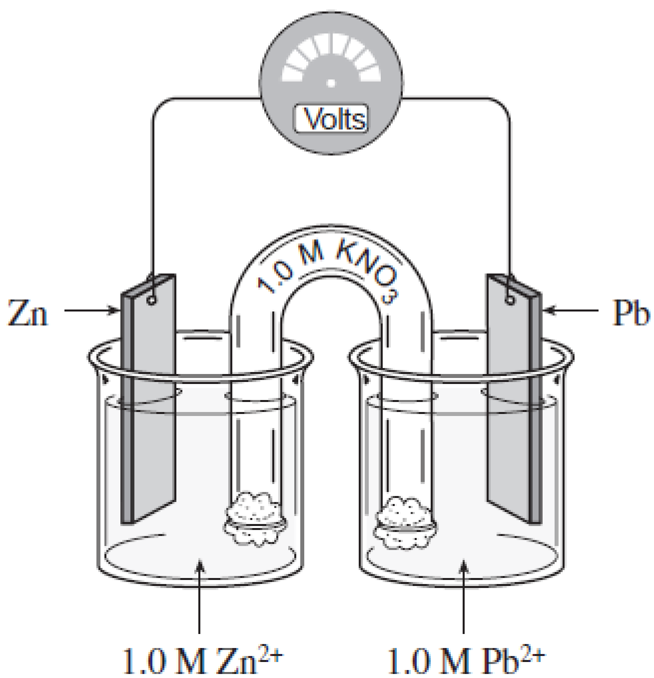
$$Cd + Pd^{2+} \rightarrow Pd + Cd^{2+}$$

$$Np^{3+} + Ce \rightarrow Np + Ce^{3+}$$

Which is the strongest oxidizing agent?

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

Questions 9 & 10 refer to this voltaic cell:



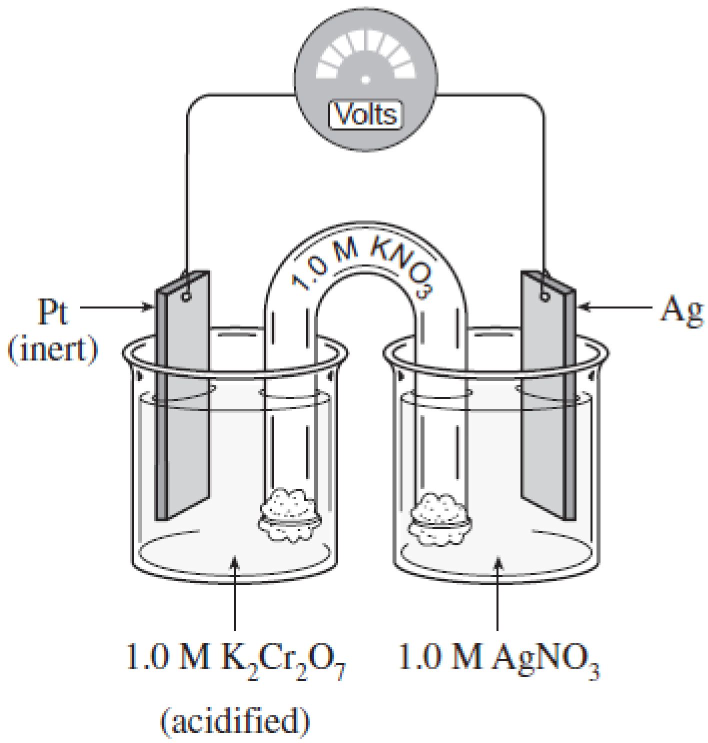
9. As this cell operates, the cations move towards the

A. Pb electrode and the electrode gains mass.
B. Pb electrode and the electrode loses mass.
C. Zn electrode and the electrode gains mass.
D. Zn electrode and the electrode loses mass.

10. As the cell operates, the electrons flow towards the

A. Zn electrode and the cell voltage increases over time.
B. Pb electrode and the cell voltage decreases over time.
C. Zn electrode and the cell voltage decreases over time.
D. Pb electrode and the cell voltage remains constant over time.

Questions 11 & 12 refer to this voltaic cell:



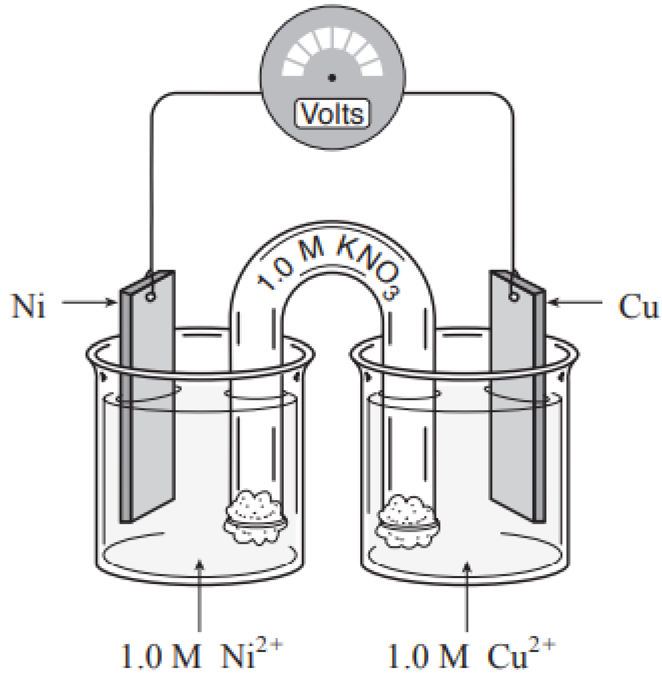
11. Which of the following represents the overall cell reaction?

A. $Cr_2O_7^{2-} + H^+ + Ag \rightarrow Ag^+ + Cr^{3+} + H_2O$
B. $Cr_2O_7^{2-} + 14H^+ + 9Ag \rightarrow 9Ag^+ + Cr^{3+} + 7H_2O$
C. $Cr_2O_7^{2-} + 14H^+ + 6Ag \rightarrow 6Ag^+ + 2Cr^{3+} + 7H_2O$
D. $Cr_2O_7^{2-} + 14H^+ + 6Ag^+ \rightarrow 6Ag + 2Cr^{3+} + 7H_2O$

12. What is the cell voltage at equilibrium?

A. −0.43 V
B. 0.00 V
C. +0.43 V
D. +2.03 V

Questions 13 & 14 refer to this voltaic cell



13. As the cell operates, observations include

	Mass of Nickel Electrode	Concentration of Copper Ions
A.	decreases	increases
B.	decreases	decreases
C.	increases	increases
D.	increases	decreases

14. What is the cell potential, E°, for this cell?

A. 0.08 V
B. 0.26 V
C. 0.60 V
D. 0.78 V

15. Which metal will react spontaneously with water?

A. Ca
B. Ni
C. Pb
D. Hg

16. Consider the following half-reaction:

$$Bi_2O_4 \rightarrow BiO^+ \qquad \text{(acidic)}$$

The balanced equation for this half-reaction is

A. $Bi_2O_4 + 6H^+ + 5e^- \rightarrow BiO^+ + 3H_2O$
B. $Bi_2O_4 + 8H^+ + 6e^- \rightarrow 2BiO^+ + 4H_2O$
C. $Bi_2O_4 + 4H^+ + 2e^- \rightarrow 2BiO^+ + 2H_2O$
D. $Bi_2O_4 + 4H^+ + 3e^- \rightarrow 2BiO^+ + 2H_2O$

17. In which of the following 1.0 M solutions will both ions react spontaneously with tin?

A. Ag^+ and Cu^{2+}
B. Ni^{2+} and Cu^{2+}
C. Zn^{2+} and Ni^{2+}
D. Mg^{2+} and Zn^{2+}

18. Consider the following redox reaction:

$$C_2H_5OH + 2Cr_2O_7^{2-} + 16H^+ \rightarrow 2CO_2 + 4Cr^{3+} + 11H_2O$$

Each carbon atom loses

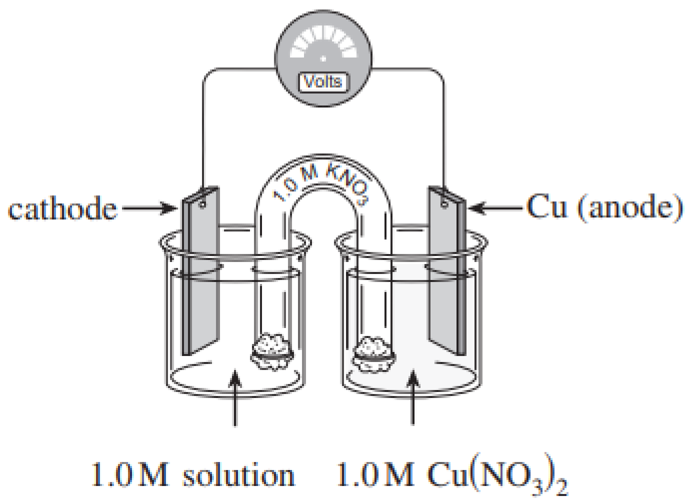
A. 2 electrons
B. 4 electrons
C. 6 electrons
D. 12 electrons

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

$$N_2O \rightarrow NH_3OH^+ \qquad \text{(acidic)}$$

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

Questions 20 & 21 refer to this voltaic cell:



20. Which material could be used as the cathode to produce an $E^\circ_{cell} = +0.46\text{ V}$?

A. Pb
B. Co
C. Ag
D. MnO_2

21. The concentration of Cu^{2+} in the copper half-cell will

A. increase as Cu loses electrons and is reduced.
B. increase as Cu loses electrons and is oxidized.
C. decrease as Cu gains electrons and is reduced.
D. decrease as Cu gains electrons and is oxidized.

22. A reducing agent

A. loses electrons and is reduced.
B. gains electrons and is reduced.
C. loses electrons and is oxidized.
D. gains electrons and is oxidized.

23. A piece of Au does not react spontaneously with 1.0 M HCl. Which of the following statements is true?

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

24. Consider the following:

$$2Cr^{2+} + Tl^{3+} \rightarrow 2Cr^{3+} + Tl^{+} \qquad E° = +1.19\ V$$

Identify the standard potential for the half-cell reaction:

$$Tl^{+} \rightarrow Tl^{3+} + 2e^{-}$$

A. −0.78 V
B. +1.60 V
C. +0.78 V
D. +1.19 V

25. Which of the following ions can be reduced by $Pb_{(s)}$ under standard conditions?

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