Electrochemistry

UNIT 5: ELECTROCHEMISTRY OUTCOMES

All important vocabulary is in Italics and bold.

- Outline the development of our understanding of ***oxidation*** and ***reduction*** reactions
- Develop an activity series
- Determine the ***oxidation numbers*** for atoms in compounds and ions
- Identify reactions as redox or non-redox
  Include: ***oxidizing agent, reducing agent***
- Balance oxidation-reduction reactions using redox methods
  Include: ***oxidation number method****, and* ***half-reaction method***
- Explain the operation of a ***voltaic*** (***galvanic***) cell at the visual, particulate and symbolic level
  Include: writing ***half-cell*** *reactions and overall reaction*
- Define standard ***electrode potential***
- Calculate ***standard cell potentials*** given standard electrode potentials
- Predict the spontaneity of reactions using standard electrode potentials
- Compare and contrast voltaic (galvanic) and ***electrolytic cells***
- Explain the operation of an electrolytic cell at the visual, particulate and symbolic levels
  Include: a ***molten*** *and* ***aqueous*** *electrolytic cells*
- Using ***Faraday's law***, solve problems related to electrolytic cells

Additional KEY Terms

line notation	anode	cathode
salt bridge	molten	

COMPLETE THE FOLLOWING QUESTIONS ON OX.NUMBER AND REDOX REACTIONS:

In each of the following chemicals, determine the oxidation states of each element:

1) sodium nitrate ______________________________

2) ammonia ______________________________

3) zinc oxide ______________________________

4) water ______________________________

5) calcium hydride ______________________________

6) carbon dioxide ______________________________

7) nitrogen ______________________________

8) sodium sulfate ______________________________

9) aluminum hydroxide ______________________________

10) magnesium phosphate ______________________________

In each of the following reactions, determine what was oxidized and what was reduced.

11) $Ca + H_2O \rightarrow CaO + H_2$

Element oxidized: ______________________________

Element reduced: ______________________________

12) $2\ H_2 + O_2 \rightarrow 2\ H_2O$

Element oxidized: ______________________________

Element reduced: ______________________________

Indicate the oxidation states of each of the elements in the following chemical species:

1) $NaOH$

2) HBr

3) VF_5

4) Na_2SO_4

5) NO_2^{-1}

6) $SiOF_2$

Determine which of the following reactions are redox reactions. For those that are, identify the element which has been oxidized and the one that has been reduced:

7) $2\ NaOH + CaCl_2 \rightarrow 2\ NaCl + Ca(OH)_2$

8) $3\ Li + FeCl_3 \rightarrow 3\ LiCl + Fe$

9) $CCl_4 + 2\ Br_2 \rightarrow CBr_4 + 2Cl_2$

10) $HNO_3 + NaHCO_3 \rightarrow H$ $_2 + NaNO_3$

COMPLETE THE FOLLOWING QUESTIONS ON OX.NUMBER AND REDOX REACTIONS:

1. Which of the following equations represents a redox reaction? For the redox reactions, identify

- element oxidized
- element reduced
- oxidizing agent
- reducing agent
- number of electrons transferred

a. $2\ Ag + Mg(NO_3)_2 \rightarrow 2\ AgNO_3 + Mg$

b. $MgO + SO_3 \rightarrow MgSO_4$

c. $2\ KMnO_4 + 8\ H_2SO_4 + 10\ KBr \rightarrow 5\ Br_2 + 6\ K_2SO_4 + 2\ MnSO_4 + 8\ H_2O$

d. $3\ Ag_2S + 8\ HNO_3 \rightarrow 6\ AgNO_3 + 2\ NO + 3\ S + 4\ H_2O$

e. $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

f. $BaCl_2 + Na_2SO_4 \rightarrow 2\ NaCl + BaSO_4$

g. $2\ Al + 3\ CuSO_4 \rightarrow 3\ Cu + Al_2(SO_4)_3$

h. $2\ NO + O_2 \rightarrow 2\ NO_2$

i. $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$

j. $2\ K_2Cr_2O_7 + 2\ H_2O + 3\ S \rightarrow 4\ KOH + 2\ Cr_2O_3 + 3\ SO_2$

k. $2\ MnO_4^- + 5\ Cu + 16\ H^+ \rightarrow 5\ Cu^{2+} + 2\ Mn^{2+} + 8\ H_2O$

l. $Na_2S + 2\ AgNO_3 \rightarrow Ag_2S + 2\ NaNO_3$

Balance the following using the Oxidation Number Method.

1. $HNO_3 + SO_2 + H_2O \rightarrow H_2SO_4 + NO$

2. $Al + CuSO_4 \rightarrow Al_2(SO_4)_3 + Cu$

3. $HNO_3 + H_2S \rightarrow NO + S + H_2O$

4. $HCl + HNO_3 \rightarrow HOCl + NO + H_2O$

5. $SnCl_4 + Fe \rightarrow SnCl_2 + FeCl_3$

6. $KMnO_4 + KCl + H_2SO_4 \rightarrow MnSO_4 + K_2SO_4 + Cl_2 + H_2O$

7. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

8. $KBr + H_2SO_4 \rightarrow K_2SO_4 + Br_2 + SO_2 + H_2O$

9. $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO + H_2O$

10. $HNO_3 + K_2CrO_4 + Fe(NO_3)_2 \rightarrow KNO_3 + Fe(NO_3)_3 + Cr(NO_3)_3 + H_2O$

Use the Half-Reaction Method to balance the following redox reactions in *acidic* solutions.

1. $MnO_4^- + Fe^{2+} \rightarrow Mn^{2+} + Fe^{3+}$

2. $BiO_3^- + MnO_2 \rightarrow Bi^{3+} + MnO_4^-$

3. $NO_3^- + SO_2 \rightarrow SO_4^{2-} + NO$

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

5. $S^{2-} + Cr_2O_7^{2-} \rightarrow S + Cr^{3+}$

6. $Br^- + SO_4^{2-} \rightarrow Br_2 + SO_2$

7. $Zn + NO_3^- \rightarrow Zn^{2+} + N_2O$

8. $C_2H_5OH + NO_3^- \rightarrow HC_2H_3O_2 + NO_2$

Use the Half-Reaction Method to balance the following redox reactions in a *basic* solution.

1. $Zn + NO_3^- \rightarrow ZnO_2^{2-} + NH_3$

2. $Bi_2O_3 + OCl^- \rightarrow BiO_3^- + Cl^-$

3. $Cl_2 + SO_3^{2-} \rightarrow SO_4^{2-} + Cl^-$

4. $CrO_2^- + S_2O_8^{2-} \rightarrow CrO_4^{2-} + SO_4^{2-}$

5. $AsO_4^{3-} + NO_2^- \rightarrow AsO_2^- + NO_3^-$

6. $ClO_3^- + N_2H_4 \rightarrow NO + Cl^-$

7. $Zn + BrO_4^- \rightarrow Zn(OH)_4^{2-} + Br^-$

8. $ClO_3^- + BH_4^- \rightarrow Cl^- + H_2BO_3^-$

Use the Standard Reduction Potentials Chart to answer the following questions.

1. Draw labeled electrochemical cells indicating the (i) electrodes, (ii) types of ions and their direction of motion, (iii) electrode reactions and (iv) the line notation or overall reaction for the cell:

   a) $3\ Mg(s) + 2\ Fe^{3+}(aq) \rightarrow 3\ Mg^{2+}(aq) + 2\ Fe(s)$

   b) $Al\ /\ Al^{3+}\ //\ Cd^{2+}\ /\ Cd$

2. Complete the following reactions using the Standard Electrode Potential table. Determine the net cell potential and state if the reaction will occur.

   a) $Zn(s) + Ag^{+}(aq) \rightarrow$

   b) $Sn(s) + Fe^{2+}(aq) \rightarrow$

   c) $Hg(l) + H^{+}(aq) \rightarrow Hg^{2+}(aq) +$

   d) $Mg(s) + Sn^{2+}(aq) \rightarrow$

   e) $Mn(s) + Cs^{+}(aq) \rightarrow$

3. Write the E° voltages for each half cell reaction, the net cell reaction and the net cell voltage. Indicate if the reaction will be spontaneous as written or not.
   a. $Ag^{+} + 1\ e^{-} \rightarrow Ag$
      $Cu \rightarrow Cu^{+} + 1\ e^{-}$
   b. $Au^{3+} + 3\ e^{-} \rightarrow Au$
      $Cr \rightarrow Cr^{3+} + 3\ e^{-}$
   c. $Cu^{+} + 1\ e^{-} \rightarrow Cu$
      $Al \rightarrow Al^{3+} + 3\ e^{-}$
   d. $Ca^{2+} + 2\ e^{-} \rightarrow Ca$
      $Fe \rightarrow Fe^{2+} + 2\ e^{-}$
   e. $NO_3^{-} + 2\ H^{+} + 1\ e^{-} \rightarrow NO_2(g) +$
      H_2O $2\ Ag\ (s) + S^{2-} \rightarrow Ag_2S + 2\ e^{-}$
   f. $2\ Br^{-} \rightarrow Br_2 + 2\ e^{-}$
      $Ni^{2+} + 2\ e^{-} \rightarrow Ni$
   g. $Cr_2O_7^{2-} + 14\ H^{+} + 6\ e^{-} \rightarrow 2\ Cr^{3+} + 7$
      H_2O $H_2Se(g) \rightarrow Se + 2\ H^{+} + 2\ e^{-}$

4. For each of the following electrochemical cells indicate

   (i) identity of the cathode
   (ii) identity of the anode
   (iii) the electrode reactions
   (iv) the net reaction
   (v) line notation
   (vi) the cell potential

   a) Nickel and silver electrodes

   b) Lead and zinc electrodes

   c) Magnesium and chlorine electrodes

   d) Sodium and manganese electrodes

5. A silver-lead cell is set up.
   a. In which half-cell does reduction occur?
   b. Write the half-cell reactions and the net reaction.
   c. Which metal is the anode?
   d. In which direction are the electrons moving?
   e. What is the cell's voltage?

Use the Reduction Potentials chart to answer the following questions.

1. In an experiment, strips of metals X, Y and Z were placed in beaker containing solutions of X^+, Y^+, Z^+. The following data was obtained.

$Y + Z^+ \rightarrow$ metallic Z deposited

$X^+ + Y \rightarrow$ metallic X deposited

$X + Z^+ \rightarrow$ No reaction

Arrange the ions above in order of increasing tendency to attract electrons.

2. Indicate whether or not the following reactions will occur based on the Standard Reduction Potential. If the reaction occurs, complete the equation.

a) $Cu^+ + Pb \rightarrow$

b) $Ag + Sn^{2+} \rightarrow$

c) $Sn + Cr^{3+} \rightarrow$

d) $NO_3^- + 4\ H^+ + Cr \rightarrow$

3. Answer the following questions about these substances: Au^{3+}, Cr, Sn^{2+}, Br^- *Note: more than one species can be used for some answers.*

a) most easily reduced

b) least affinity for electrons

c) least easily oxidized

d) most easily oxidized

e) which will reduce $Au^{3+} \rightarrow Au(s)$

4. Substances A, B, C, D & E are metals, which form positive ions. Ions of metal A react with metal E but not with metal C. However, metal C does react with solutions containing ions of metal D & B. Metal D will not react with ions of metal B.

List the metal ions placing the best oxidizing agent at the bottom.

5. In each of the following situations, write balanced net ionic equations to represent the reaction expected and indicate whether the reaction is spontaneous or not.

   a) A copper strip of metal is placed in a solution of zinc nitrate.

   b) A lead strip is placed in a solution of zinc nitrate.

   c) Hydrochloric acid is placed into a container of lead metal.

   d) Hydrochloric acid is placed into a container of zinc metal.

   e) A copper strip is placed in a solution of lead (II) nitrate.

   f) A Zinc strip is placed in a solution of copper (II) nitrate.

   g) A solution of hydrochloric acid is placed into a container of copper.

   h) Lead (II) nitrate solution is placed into a container of zinc metal.

   i) A lead strip is placed in a solution of copper (II) nitrate.

   j) Fluorine gas is bubbled into a sodium bromide solution.

Answer the following questions using Faraday's law. Be sure to show all your work.

1. If 7.85 amp flows through a molten solution of copper I chloride for 45.0 min,how many moles of electrons flow through the cell? **(0.220)**

2. How many seconds would be needed to generate 3.00 moles of electrons from 10.0 amp of current? **(28950)**

3. Using 2.50 moles of electrons, what mass of copper metal would be produced from molten copper (II) sulfate? **(79.4)**

4. If 9.00 amp flows for 10.0 min through an aqueous silver nitrate solution, what mass of silver metal would be formed? **(6.04)**

5. What time would be needed to deposit 42.50 g of Zinc metal from 5.00 amp of current through a molten solution of Zinc (II) bromide? **(25048)**

6. Calculate the mass of products generated at each electrode if 15.0 amp flows for 7.51 minutes through molten MgF_2. **(0.851, 1.33)**

7. How long would an aqueous gold (III) chloride cell need to operate to plate 2.5 g of gold on a bracelet with a current of 2.5 A? **(1470)**

8. If 10.0 g of sulfur is deposited on an electrode in an electrolytic cell, calculate the mass of silver deposited on the other electrode. **(67.2)**

9. How long will it take to use up all the Cr^{3+} ions in 400.0 mL of a 0.120 mol/L solution using a current of 1.50 A? **(9264)**

10. What products would be expected at each electrode? Write the half reactions and calculate the battery voltage needed to electrolyze the following:
    a) molten silver bromide **(0.26 V)**

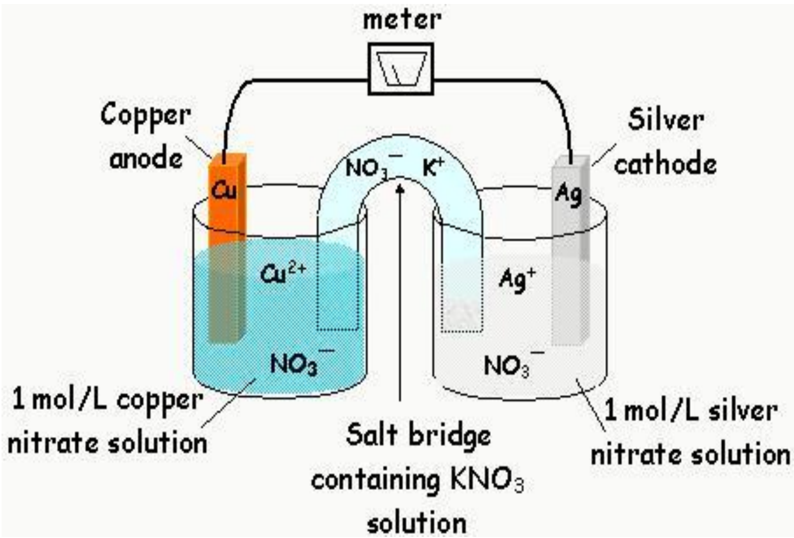
    b) molten magnesium chloride **(3.73 V)**

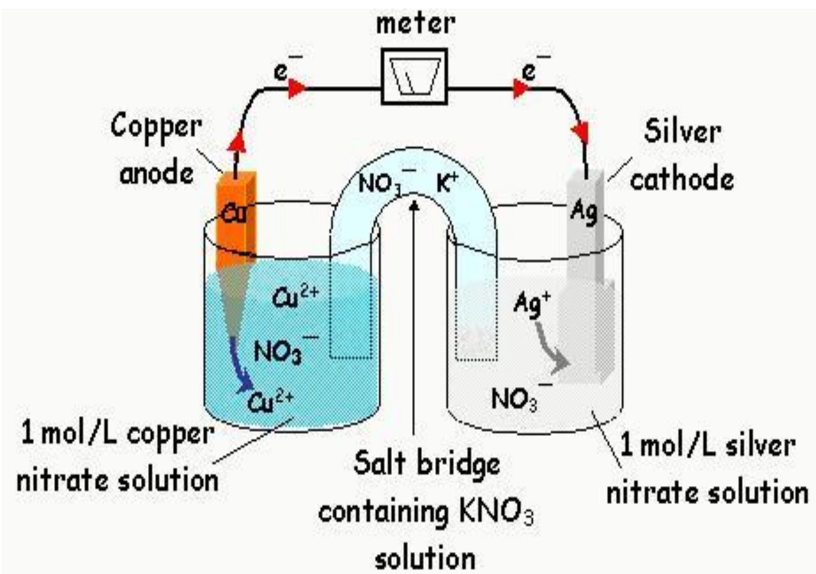
    c) ***aqueous*** copper (II) fluoride **(0.89 V)**

    d) ***aqueous*** aluminum chloride **(2.06 V)**

    e) ***aqueous*** iron (III) sulphate **(1.67 V)**

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





Standard Reduction Potentials

All values are for 1.0 M aqueous solutions at 25°C

Half-Reaction			E°(volts)
$F_2(g) + 2e^-$	$\rightarrow$	$2F^-$	+2.87
$H_2O_2 + 2H^+ + 2e^-$	$\rightarrow$	$2H_2O$	+1.77
$MnO_4^- + 8H^+ + 5e^-$	$\rightarrow$	$Mn^{2+} + 4H_2O$	+1.52
$Au^{3+} + 3e^-$	$\rightarrow$	$Au(s)$	+1.50
$Cl_2(g) + 2e^-$	$\rightarrow$	$2Cl^-$	+1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^-$	$\rightarrow$	$2Cr^{3+} + 7H_2O$	+1.33
$MnO_2(s) + 4H^+ + 2e^-$	$\rightarrow$	$Mn^{2+} + 2H_2O$	+1.28
$2O_2(g) + 4H^+ + 4e^-$	$\rightarrow$	$2H_2O$	+1.23
$Br_2(g) + 2e^-$	$\rightarrow$	$2Br^-$	+1.06
$NO_3^- + 4H^+ + 3e^-$	$\rightarrow$	$NO(g) + 2H_2O$	+0.96
$Ag^+ + e^-$	$\rightarrow$	$Ag(s)$	+0.80
$1/2\ Hg_2^{2+} + e^-$	$\rightarrow$	$Hg(l)$	+0.79
$Hg^{2+} + 2e^-$	$\rightarrow$	$Hg(l)$	+0.78
$NO_3^- + 2H^+ + e^-$	$\rightarrow$	$NO_2(g) + H_2O$	+0.78
$Fe^{3+} + e^-$	$\rightarrow$	Fe^{2+}	+0.77
$O_2(g) + 2H^+ + 2e^-$	$\rightarrow$	H_2O_2	+0.68
$I_2(s) + 2e^-$	$\rightarrow$	$2I^-$	+0.53
$Cu^+ + e^-$	$\rightarrow$	$Cu(s)$	+0.52
$Cu^{2+} + 2e^-$	$\rightarrow$	$Cu(s)$	+0.34
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightarrow$	$SO_2(g) + 2H_2O$	+0.17
$Cu^{2+} + e^-$	$\rightarrow$	Cu^+	+0.15
$Sn^{4+} + 2e^-$	$\rightarrow$	Sn^{2+}	+0.15
$S + 2H^+ + 2e^-$	$\rightarrow$	$H_2S(g)$	+0.14
$2H^+ + 2e^-$	$\rightarrow$	$H_2(g)$	0.00
$Fe^{3+} + 3e^-$	$\rightarrow$	$Fe(s)$	-0.04
$Pb^{2+} + 2e^-$	$\rightarrow$	$Pb(s)$	-0.13
$Sn^{2+} + 2e^-$	$\rightarrow$	$Sn(s)$	-0.14
$Ni^{2+} + 2e^-$	$\rightarrow$	$Ni(s)$	-0.25
$Co^{2+} + 2e^-$	$\rightarrow$	$Co(s)$	-0.28
$Cd^{2+} + 2e^-$	$\rightarrow$	$Cd(s)$	-0.40
$Se + 2H^+ + 2e^-$	$\rightarrow$	$H_2Se(g)$	-0.40
$Cr^{3+} + e^-$	$\rightarrow$	Cr^{2+}	-0.41
$Fe^{2+} + 2e^-$	$\rightarrow$	$Fe(s)$	-0.44
$Cr^{2+} + 2e^-$	$\rightarrow$	$Cr(s)$	-0.56
$Ag_2S + 2e^-$	$\rightarrow$	$2Ag(s) + S^{2-}$	-0.69
$Te + 2H^+ + 2e^-$	$\rightarrow$	$H_2Te(g)$	-0.72
$Cr^{3+} + 3e^-$	$\rightarrow$	$Cr(s)$	-0.74
$Zn^{2+} + 2e^-$	$\rightarrow$	$Zn(s)$	-0.76
$2H_2O + 2e^-$	$\rightarrow$	$2OH^- + H_2(g)$	-0.83
$Mn^{2+} + 2e^-$	$\rightarrow$	$Mn(s)$	-1.18
$Al^{3+} + 3e^-$	$\rightarrow$	$Al(s)$	-1.66
$Mg^{2+} + 2e^-$	$\rightarrow$	$Mg(s)$	-2.37
$Na^+ + e^-$	$\rightarrow$	$Na(s)$	-2.71
$Ca^{2+} + 2e^-$	$\rightarrow$	$Ca(s)$	-2.87
$Sr^{2+} + 2e^-$	$\rightarrow$	$Sr(s)$	-2.89
$Ba^{2+} + 2e^-$	$\rightarrow$	$Ba(s)$	-2.90
$K^+ + e^-$	$\rightarrow$	$K(s)$	-2.92

$Li^{+} + e^{-} \rightarrow Li(s)$ -3.00

POLYATOMIC IONS

NAME	*FORMULA*	*CHARGE*
ACETATE	CH_3COO^-	**-1**
AMMONIUM	NH_4^+	**+1**
HYDROGEN CARBONATE *(BICARBONATE)*	HCO_3^-	**-1**
CARBONATE	CO_3^{-2}	**-2**
CHLORATE	ClO_3^-	**-1**
CHLORITE	ClO_2^-	**-1**
CHROMATE	CrO_4^{-2}	**-2**
DICHROMATE	$Cr_2O_7^{-2}$	**-2**
DIHYDROGEN PHOSPHATE	$H_2PO_4^-$	**-1**
HYDROGEN PHOSPHATE	HPO_4^{-2}	**-2**
PHOSPHATE	PO_4^{-3}	**-3**
HYDROGEN SULFATE *(BISULFATE)*	HSO_4^-	**-1**
SULFATE	SO_4^{-2}	**-2**
HYDROGEN SULFITE *(BISULFITE)*	HSO_3^-	**-1**
SULFITE	SO_3^{-2}	**-2**
HYDRONIUM	H_3O^+	**+1**
HYDROXIDE	OH^-	**-1**
PERCHLORATE	ClO_4^-	**-1**
HYPOCHLORITE	ClO^- *(OCl^-)*	**-1**
NITRATE	NO_3^-	**-1**
NITRITE	NO_2^-	**-1**
PERMANGANATE	MnO_4^-	**-1**
THIOCYANATE	SCN^-	**-1**