I've got the Power! The Voltaic Cell

Why?

How exactly does a battery work? The tendency for electrons to flow from one substance to another is something that can be channeled and controlled. A voltaic (or electrochemical) cell provides a mechanism for the conversion of chemical energy into electrical energy. The cell potential, or voltage, and current that are produced by an electrochemical cell are determined by the chemicals that comprise the cell. In this activity, you will see that the amount of voltage produced by a cell depends upon the chemical composition.

Learning Objectives

- Understand how the cell potential is related to composition of the cell.
- See how measurements of cell potentials are used to construct an activity series
- Determine whether a redox reaction is spontaneous or non-spontaneous

Success Criteria

- Demonstrate an understanding of the operation of a voltaic cell.

Prerequisites

- Ability to write oxidation and reduction half-reactions
- Introduction to Voltaic Cell diagrams and terminology

Requirements

- Activity Series reference table
- Computer and internet access for the following site:

http://www.blackgold.ab.ca/ict/Division4/Science/Div.%204/Voltaic%20Cells/Voltaic.htm

Vocabulary

- Reduction
- Voltaic cell
- Anode
- Salt bridge
- Oxidation
- Half-cell
- Cathode

Information

Cell potentials for spontaneous voltaic cells are positive. A voltmeter reading can be positive or negative depending on which wires are connected to which electrodes. If a voltmeter gives a negative reading for voltage, switching the connections between the wires and the electrodes will produce a positive voltmeter reading.

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Model

Connect to the web site:
http://www.blackgold.ab.ca/ict/Division4/Science/Div.%204/Voltaic%20Cells/Voltaic.htm
Scroll down to find the diagram of a voltaic cell.

Select the metals and solutions for the cell to set up the simulation for each voltaic cell as shown in the model below. Once you have the electrodes and solutions selected for the cell, press the switch "on" to start the flow of electrons. [Note: the cell potential should be positive. If the voltmeter reading is negative, reverse the electrodes and try again.] Record what happens to the electrons shown in the model, and record the voltage produced by each cell.

Cell	Electrode	Solution (1M)	Electrode	Solution (1M)	Cell Potential
A	Ag	$AgNO_3$	Cu	$Cu(NO_3)_2$	
Electrons flow from _____half cell to the _____ half cell.					

Refresh the site on your browser and repeat the steps, this time using the combination shown in the following table for Cell (B).

Cell	Electrode	Solution (1M)	Electrode	Solution (1M)	Cell Potential
B	Cu	$Cu(NO_3)_2$	Zn	$Zn(NO_3)_2$	
Electrons flow from _____half cell to the _____ half cell.					

Refresh the site on your browser and then repeat steps, this time using the combination shown in the following table for Cell (C).

Cell	Electrode	Solution (1M)	Electrode	Solution (1M)	Cell Potential
C	Ag	$AgNO_3$	Zn	$Zn(NO_3)_2$	
Electrons flow from _____half cell to the _____ half cell.					

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Key Questions

1. As each cell in the model goes to completion, which half-cell shows an increase in the concentration of metallic ions?

Cell A	
Cell B	
Cell C	

2. As each cell in the model goes to completion, which half-cell shows a decrease in concentration of metallic ions?

Cell A	
Cell B	
Cell C	

3. Which set of half-cells (those in Key Question 1 or those in Key Question 2) indicate that oxidation is taking place in the half-cell?

4. What is occurring in the set of half-cells that is not selected as your answer to Key Question 3?

5. Which of the cells produces the largest positive voltage? Which produces the lowest positive voltage?

6. Select one of the cells and set it up to produce a negative voltage. Describe the cell that you set up, and describe what you observe as the cell runs. Compare the observations of this cell to the observations made when the cell was set up to produce a positive voltage. Why does this occur?

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Exercises

1. For each of the half-cells named in Key Question 1, write the half-reaction that describes what is happening in each of the half-cells.

Cell A	
Cell B	
Cell C	

2. For each of the half-cells named in Key Question 2, write the half-reaction that describes what is happening in each of the half-cells.

Cell A	
Cell B	
Cell C	

3. In a voltaic cell, where does oxidation occur, anode or cathode?

4. In a voltaic cell, where does reduction occur, anode or cathode?

5. In a voltaic cell, what is the direction of electron flow?

6. Based on your data for *cell potential and the direction of electron flow* in the cells recorded in the Model, create a list that ranks the elements copper, silver, and zinc from most active to least active.

7. Explain how you decided on the relative ranking of the metals.

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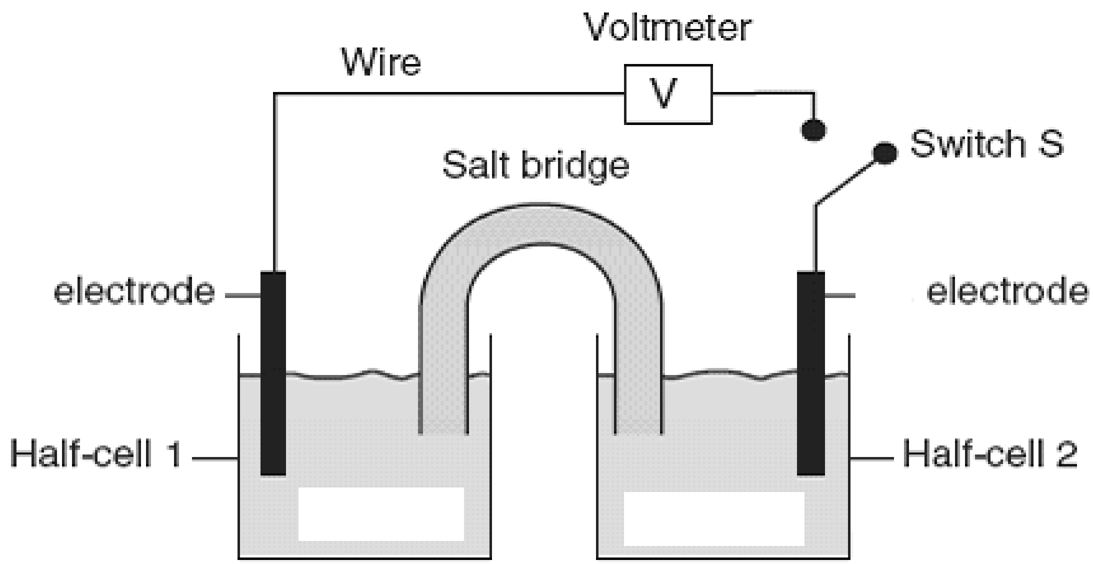
8. The cells in the Model contain metals that are found on Activity Series-Reference Table J. Describe how your relative ranking compares to the placement of the metals on the Activity Series.

Problems

The following two half-cells are prepared: a piece of silver metal is placed in a 1 M solution of silver nitrate and a piece of nickel metal is placed in a 1 M solution of nickel nitrate. The metal strips are connected by a wire, and the solutions are connected by a salt bridge. The salt bridge contains a saturated potassium nitrate solution. When the switch is closed the circuit is complete.

1. In the cell below, label each electrode, indicate the location of each 1 M solution and use arrows to indicate the flow of electrons.



2. The possible reduction equations are: $Ag^{+} + e^{-} \rightarrow Ag$ and $Ni^{2+} + 2e^{-} \rightarrow Ni$. Which of the metals in this cell is more likely to be reduced? How do you know?

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3. Write the equation for the half-reaction occurring at the anode.

4. Write the equation for the half-reaction occurring at the cathode.

5. Write the overall redox reaction.

6. Identify the species being oxidized and the species being reduced.

7. For each of the following statements, indicate if it is true or not. If true, explain why it is true; if false, explain why it is false and suggest how the statement can be changed to make it true.

   (a) Electrons flow from the nickel electrode to the silver electrode.

   (b) The silver electrode increases in mass as the cell operates.

   (c) Ag^{+} ions move through the salt bridge to the nickel half-cell.

   (d) Negative ions move through the salt bridge to the silver half-cell.

   (e) Ni^{2+} ions move from the nickel half-cell toward the silver half-cell.

8. What would happen to the voltage if the salt bridge were removed from a voltaic cell? Why would this occur?

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STANDARD ELECTRODE POTENTIALS	
Ionic Concentrations 1 M Water At 298 K, 1 atm	
Half-Reaction	E^0 (volts)
$F_2(g) + 2e^- \rightarrow 2F^-$	+2.87
$8H^+ + MnO_4^- + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51
$Au^{3+} + 3e^- \rightarrow Au(s)$	+1.50
$Cl_2(g) + 2e^- \rightarrow 2Cl^-$	+1.36
$14H^+ + Cr_2O_7^{2-} + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	+1.23
$4H^+ + O_2(g) + 4e^- \rightarrow 2H_2O$	+1.23
$4H^+ + MnO_2(s) + 2e^- \rightarrow Mn^{2+} + 2H_2O$	+1.22
$Br_2(\ell) + 2e^- \rightarrow 2Br^-$	+1.09
$Hg^{2+} + 2e^- \rightarrow Hg(\ell)$	+0.85
$Ag^+ + e^- \rightarrow Ag(s)$	+0.80
$Hg_2^{2+} + 2e^- \rightarrow 2Hg(\ell)$	+0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77
$I_2(s) + 2e^- \rightarrow 2I^-$	+0.54
$Cu^+ + e^- \rightarrow Cu(s)$	+0.52
$Cu^{2+} + 2e^- \rightarrow Cu(s)$	+0.34
$4H^+ + SO_4^{2-} + 2e^- \rightarrow SO_2(aq) + 2H_2O$	+0.17
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	+0.15
$2H^+ + 2e^- \rightarrow H_2(g)$	0.00
$Pb^{2+} + 2e^- \rightarrow Pb(s)$	−0.13
$Sn^{2+} + 2e^- \rightarrow Sn(s)$	−0.14
$Ni^{2+} + 2e^- \rightarrow Ni(s)$	−0.26
$Co^{2+} + 2e^- \rightarrow Co(s)$	−0.28
$Fe^{2+} + 2e^- \rightarrow Fe(s)$	−0.45
$Cr^{3+} + 3e^- \rightarrow Cr(s)$	−0.74
$Zn^{2+} + 2e^- \rightarrow Zn(s)$	−0.76
$2H_2O + 2e^- \rightarrow 2OH^- + H_2(g)$	−0.83
$Mn^{2+} + 2e^- \rightarrow Mn(s)$	−1.19
$Al^{3+} + 3e^- \rightarrow Al(s)$	−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
$Ba^{2+} + 2e^- \rightarrow Ba(s)$	−2.91
$Cs^+ + e^- \rightarrow Cs(s)$	−2.92
$K^+ + e^- \rightarrow K(s)$	−2.93
$Rb^+ + e^- \rightarrow Rb(s)$	−2.98
$Li^+ + e^- \rightarrow Li(s)$	−3.04

Table J
Activity Series**

Metals (Most → Least)	Nonmetals (Most → Least)
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
**H_2	
Cu	
Ag	
Au	

**Activity Series based on hydrogen standard
Note: H_2 is *not* a metal

Standard Electrode Potentials: http://nysedregents.org/testing/reftable/archreftable/chemref.pdf
Table J: http://nysedregents.org/testing/reftable/archreftable/ChemRef1-7.pdf

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