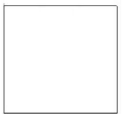
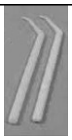
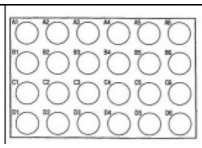
AP Lab 16Galvanic Cells

Lab Equipment Chromatography paper or cotton string  		Chemicals in kit: 0.1 *M* copper(II) nitrate, $Cu(NO_3)_2$ 0.1 *M* zinc nitrate, $Zn(NO_3)_2$ 0.1 *M* nickel(II) nitrate, $Ni(NO_3)_2$ 0.1 *M* sodium acetate, $NaC_2H_3O_2$ metals strips: Cu, Ni, Zn fine sandpaper	
 toothpick	 24-well reaction plate	Student supplied:  lemon or lime sliced in half (optional)	 any 1.5 V battery
 alligator clips	 Digital voltmeter	 Tesla Model S (optional)	

GOALS

1. What voltages can galvanic cells generate? Cu^{2+}
2. How do the reduction potentials of substances compare to hydrogen's reduction potential?

Part I Background

Your lab kit has several solutions with cations that are capable of reducing to metals in aqueous solutions.

0.1 *M* copper(II) nitrate, $Cu(NO_3)_2$, copper	$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
0.1 *M* zinc nitrate, $Zn(NO_3)_2$, zinc	$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$
0.1 *M* nickel(II) nitrate, $Ni(NO_3)_{2,}$, nickel	$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$

The other metal ion solutions in your kit cannot be reduced to metals since any metal that would be reduced would immediately oxidize in back into its ion as it reduced with water to make hydrogen gas.

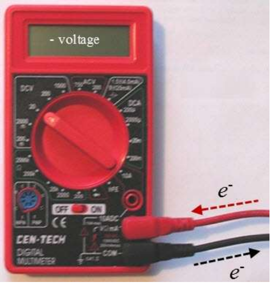
E.g. If any Na^+ reduced to Na, the Na would react with water: $2\ Na + H_2O \rightarrow Na^+ + H_2 + 2\ OH^-$

You will create a series of galvanic cells to find out how the reduction potentials of the three metal ions compare.

CAVEAT: Small electrochemical cells produce so little current that it is very difficult to get accurate voltage readings without specialized voltmeters. Don't expect prefect results with this experiment done on a micro scale.

Part I Procedure:

(1) You will use a Digital Multimeter, DMM, to determine any potential difference between the half-cells. If they are not already attached, clip the alligator clips over the probe ends of the DMM. Insert the black probe wire plug into the COM socket and the red wire probe wire plug should be inserted to the VΩmA socket[1]. With the wires set up this way you can determine the direction of electron flow based on the sign of the voltage[2].

	
positive voltage reading red probe is connected to the cathode black probe is connected to the anode	negative voltage reading red probe is connected to the anode black probe is connected to the cathode

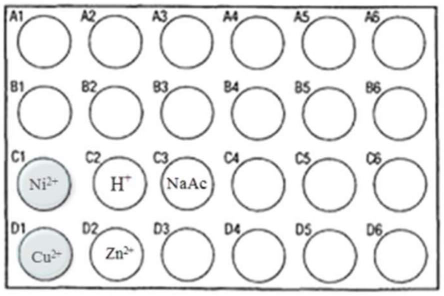
(2) Turn the dial's pointer to the DCV range to measure direct current voltages. The 20 V range will have more than enough sensitivity for this experiment.

(3) Get a 1.5 V battery (AA or any other). With the dial turned to the 20 V DCV range, flip the DMM main off-on switch to on. Touch the red probe to the (+ cathode) marked pole of the battery and the black to the (- anode) pole of the battery. The voltage reading should be positive indicating that electrons are moving from the (-) terminal of the battery through the DMM to the (+) pole. The voltage should be close to the voltage rating of the battery. Turn the off-on switch of the DMM off to conserve the internal battery of the meter.

DO NOT turn the dial to the DCA range while testing the battery. The DCA range is used to measure current. With no resistance, the battery could easily drive enough current through the DMM to overload the fuse of the DMM.

If your DMM doesn't' read close to 1.5 V then either you don't have the DMM dial set correctly, your battery is dead, or your DMM is defective (or the internal battery of the DMM is low or dead).

(4) Place 15 drops of 0.10 *M* copper(II) nitrate, $Cu(NO_3)_2\,(aq)$, in cell D1.
Place 15 drops of 0.10 *M* zinc nitrate, $Zn(NO_3)_2\,(aq)$, in cell D2.
Place 15 drops of 0.10 *M* nickel(II) nitrate, $Ni(NO_3)_2\,(aq)$, in cell C1.
Place 15 drops of 0.10 *M* hydrochloric acid, $HCl(aq)$, in cell C2.
Place 30 drops of 0.10 *M* sodium acetate, $NaC_2H_3O_2\,(aq)$, in cell C3.



(5) Clean the surface of the small strips of Cu, Zn, and Ni in your kit with the fine sandpaper to remove any oxide coatings on the metal surfaces. Use a fresh spot on sandpaper for each metal to prevent transfer of metal particles from one metal to another.

(6) Cut five strips of chromatography paper, or the coffee filter paper about 0.3 cm wide and 4 cm long, or cut five pieces of cotton string about 4 cm long.

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[1] The DMM will work just as well with the colors reversed, but following this convention will help with determining the direction of electrons.

[2] The sign and color conventions of current are unusual because the direction of current used in physics and electronics texts is opposite that of the actual flow the electrons.

(7) Soak a paper strip or string in the sodium acetate solution in cell C3. The sodium acetate will not react with any of the solutions and will act as your salt bridge.

(8) Drape the salt bridge across the copper(II) nitrate zinc nitrate cells to allow for the flow of ions between the two cells. Use a toothpick to poke the paper or string into the solutions to be sure that the salt bridge is in both solutions.

(9) Attach the copper strip to the red wire alligator clip and the zinc metal strip to the black wire alligator clip and dip the wires into their matching ion solutions. There should a voltage reading. If not, make sure that the salt bridge is submerged in both solutions. Record the most stable reading of the Cu-Zn cell along with its sign. Also indicate direction of electron flow based on the sign.

(10) Drape a soaked salt bridge from the copper(II) solution to the nickel (II) solution. You don't have to remove the salt bridge from the copper(II)-zinc solution.

(11) Replace the zinc metal on the black alligator clip with the nickel metal and record the voltage and sign of the Cu-Ni cell.

(12) Drape a soaked salt bridge from the nickel(II) solution to the zinc solution.

(13) Replace the Cu metal on the red alligator clip with the zinc metal and leave the nickel metal on the black alligator clip probe. Record the voltage and sign of the Ni-Zn cell.

(14) remove the salt bridges and turn of the master switch of your DMM.

Based on the data determine the relative reduction potential of the metal ions and the net ionic reaction for each galvanic cell.

Part II Background

The hydrogen ion reduction reaction is the standard which sets the reduction potentials of other substances as positive for materials that reduce better than hydrogen ions and negative for substances that are not as good as hydrogen ions at reducing. Of the elements we tested, zinc(II) has the most negative voltage relative to hydrogen.

You will test the zinc oxidation by acid using any nonreactive metal electrode as the cathode in the acid. The hydrogen ions from the acid are reduced on the surface of the cathode. The reduction reaction taking place is

$$2\ H^+\ (aq) + 2\ e^- \rightarrow H_2(g$$

Hydrogen is a product from the acid. The cathode metal itself is not involved in the reaction. Therefore, composition of the cathode will not be relevant.

Part II Procedure:

(1) Drape a soaked salt bridge from the zinc solution to the hydrochloric acid solution.

(2) Attach a strip of zinc to the red alligator clip and a copper strip on the black alligator clip. The copper will act as an inert electrode in the acid. Gold or platinum would be a better electrode and you may use those if you prefer.

(3) Place the zinc-red probe in the NaAc solution in (C3) and the copper (or gold) -black probe in the hydrogen ion solution in (C2). Record the voltage and sign.

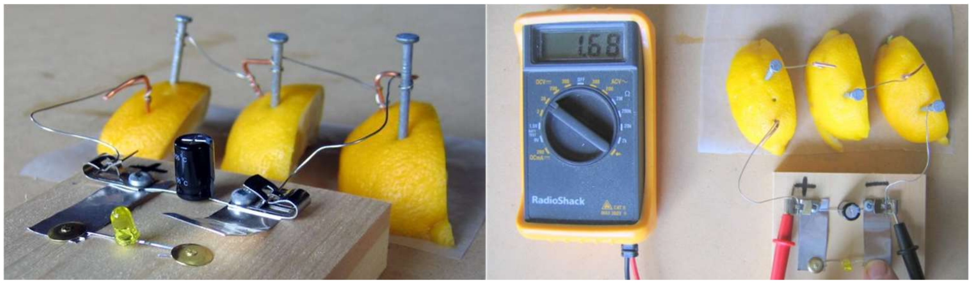
Experiment	Red probe	Black Probe	Voltage (V)
1	Zn | 0.0 *M* Zn^{2+} (*aq*) in (D2)	Cu | H^+ in (C2)	

Rather than rinsing your spot plate under a running faucet, submerge the spot plate into a pail or large container with water and a tablespoon or two of baking soda. This will neutralize the acid and also prevent the solutions from splattering out. You can then dispose of the diluted solutions directly down a drain. It’s not advisable to pour into a stainless-steel sink since the copper(II) ions are corrosive.

Part III Optional : Lemon Battery

Once done with the experiment discard the lemons. Since it was used in the experiment is should not be used for food.

https://www.sciencebuddies.org/stem-activities/lemon-battery

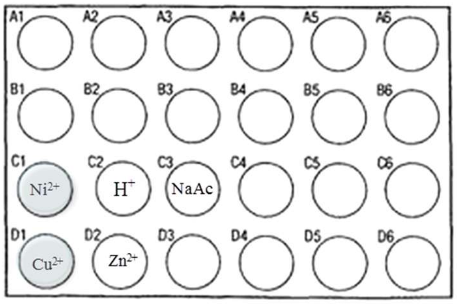


EXP. NUMBER	EXPERIMENT/SUBJECT	DATE
AP L16	Galvanic Cells	

I put metal three nitrate solutions, a hydrochloric acid solution and a sodium acetate solution in a separate wells of a 24 well spot plate. All solutions were 0.1 M.

The nitrate solutions were copper(II) nitrate, nickel(II) nitrate and zinc nitrate.

I soaked some cotton string in the 0.10 M sodium acetate solution and used it as the salt bridge to complete the circuit of my cells.



After draping the salt bridge to connect the cells, I tested for electron potential using copper, zinc and metal strips attached to the alligator clips of a digital multimeter set to the 20 V range.

The red wire of the DMM was connected to the VΩmA socket and the black wire was connected to the COM.

Using this DMM setup a (+) voltage indicates that the red probe is connected to cathode.
A (-) voltage indicates that the black probe is connected to the cathode.

Part I Experiment	Red probe	Electron flow	Black Probe	DVM reading (V)
1	$Cu(s) \mid Cu^{2+}(aq)$	←	$Zn^{2+}(aq) \mid Zn(s)$	+0.86
2	$Cu(s) \mid Cu^{2+}(aq)$	←	$Ni^{2+}(aq) \mid Ni(s)$	+0.12
3	$Zn^{2+}(aq) \mid Zn(s)$	→	$Ni^{2+}(aq) \mid Ni(s)$	-0.53

I also tested for zinc's relative potential to the hydrogen ion using hydrochloric acid.

Part II Experiment	Red probe	electron flow	Black Probe	DVM reading (V)
1	$Zn \mid Zn^{2+}$	→	$H^{+} \mid Cu$	-0.10

As an optional experiment, I set up a series battery by chaining together copper metal-zinc nail electrodes in the acidic citric acid (pH = 2) of a lemon to get a voltage of 1.8 V

Part I

Based on the experimental results these would be the net reactions for each experiment.

Part I	*Net reactions*
1	$Cu^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cu(s)$
2	$Cu^{2+}(aq) + Ni(s) \rightarrow Ni^{2+}(aq) + Cu(s)$
3	$Ni^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Ni(s)$

<table>
<tr><td><i>Experiment 1</i></td><td>
<p><i>The voltage was + which meant that the Cu was the cathode.</i></p>
<p><i>Cu^{2+} reduced to Cu and Zn was the anode as Zn metal oxidized into Zn^{2+}.</i></p>
<p><i>∴ Cu^{2+} must have a higher reduction potential than Zn^{2+}</i></p>
<table>
<tr><td>Copper</td><td>$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$</td></tr>
<tr><td>Zinc</td><td>$Zn^{2+}(aq) + 2e^- \leftarrow Zn(s)$</td></tr>
</table>
</td></tr>
<tr><td><i>Experiment 2</i></td><td>
<p><i>The voltage was + again which meant that the Cu was the cathode</i></p>
<p><i>Cu^{2+} reduced and Ni was oxidized into Ni^{2+}.</i></p>
<p><i>∴ Cu^{2+} must have a higher reduction potential than nickel ions.</i></p>
<table>
<tr><td>Copper</td><td>$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$</td></tr>
<tr><td>Nickel</td><td>$Ni^{2+}(aq) + 2e^- \leftarrow Ni(s)$</td></tr>
</table>
<p><i>The voltage was smaller than the Zn (0.12 V vs 0.86 V) so I would speculate that Ni was more difficult to oxidize than Zn.</i></p>
<p><i>∴ Ni^{2+} would have a higher reduction potential than Zn^{2+}.</i></p>
<p><i>So the order of reduction potential would be:</i></p>
<table>
<tr><td>Copper</td><td>$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$</td></tr>
<tr><td>Nickel</td><td>$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$</td></tr>
<tr><td>Zinc</td><td>$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$</td></tr>
</table>
</td></tr>
<tr><td><i>Experiment 3</i></td><td>
<p><i>The voltage was + which meant that the Ni was the cathode and Ni^{2+}must have reduced and theZn was oxidized.</i></p>
<table>
<tr><td>Nickel</td><td>$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$</td></tr>
<tr><td>Zinc</td><td>$Zn^{2+}(aq) + 2e^- \leftarrow Zn(s)$</td></tr>
</table>
<p><i>∴ Ni^{2+} must have a higher reduction potential than Zn^{2+}.</i></p>
<p><i>This reinforces the smaller difference in voltage for this cell than the Cu cell.</i></p>
<p><i>The order of reduction potentials based on these three experiments would be:</i></p>
<table>
<tr><td>Copper</td><td>$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$</td></tr>
<tr><td>Nickel</td><td>$Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$</td></tr>
<tr><td>Zinc</td><td>$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$</td></tr>
</table>
</td></tr>
</table>

Part II

The hydrogen half cells was different in that hydrogen ions could only be reduced in its half cell.

The -0.10 V measurement meant that the Cu (black wire) was the cathode.

H^+ reduced to H_2 on the non-reacting copper electrode and Zn was the anode as Zn metal oxidized into Zn^{2+}.

$\therefore$ H^+ must have a higher reduction potential than Zn^{2+}

(Copper) in 0.10 M HCl	$2\,H^+(aq) + 2e^- \rightarrow H_2(g)$
Zinc in 0.10 M $Zn^{2+}(aq)$	$Zn^{2+}(aq) + 2e^- \leftarrow Zn(s)$

Optional Experiment

Setting up a series cell with copper and zinc in lemons allowed the hydrogen ions in the acidic lemon juice to be reduced to hydrogen and zinc metal to oxidize into zinc ions. Making a less acidic lemon with extra zinc ions.

In series, the voltage increased in an additive manner. I used 4 lemons to get 3 V

Using approx 8 million lemons (and copper cathodes and zinc anodes) I could recharge my 60 KW Tesla electric car,