Experiment 7: Electrochemistry

Part One - Depositing a metal from solution

Results

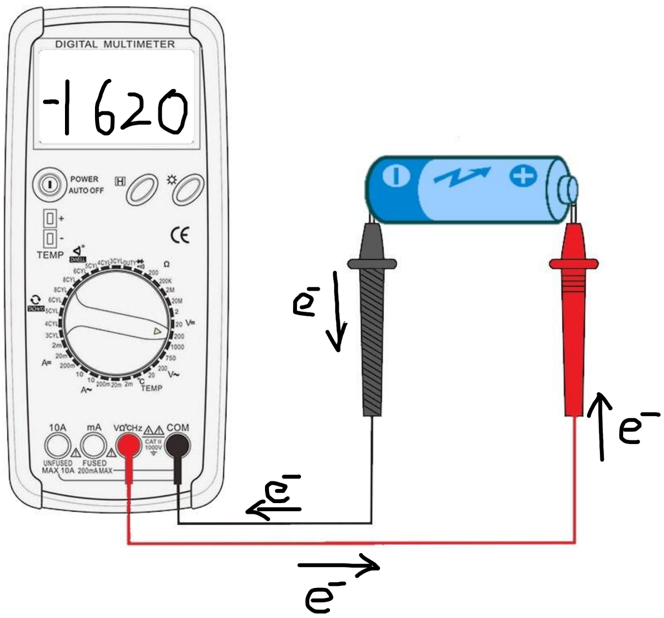
Reaction phenomenon:
2^{nd} test tube: add some solid zinc
Brown copper is formed on the surface of the zinc sheet, zinc sheet produces a few bubbles, brown precipitate appeared in the solution and the solution changes from blue to colorless.
3^{rd} test tube: add some solid iron
Reddish brown copper is formed on the surface of the iron nail, and the solution changes from blue to pale green.
4^{th} test tube: add some solid magnesium
Green copper is formed on the surface of the magnesium strip, magnesium strip produces a few bubbles, green precipitate appeared in the solution and the solution changes from blue to green.

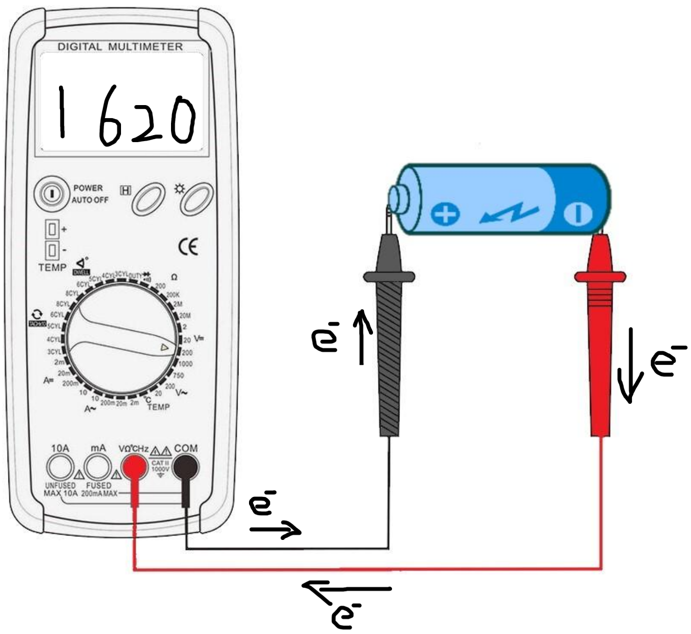
The half-equation in which a species gains electrons is called reduction.
$Zn + Cu2+ \rightarrow Zn2+ + Cu$
$Fe + Cu2+ \rightarrow Fe2+ + Cu$
$Mg + Cu2+ \rightarrow Mg2+ + Cu$

If we put a piece of copper metal into the test-tube of copper sulphate, copper and Zn/Fe/Mg forming galvanic cell. Zn: The solution becomes lighter in color, the copper surface is covered with a red substance, and the zinc sheet becomes thinner. Fe: The solution changes from blue to pale green, the copper surface is covered with a red substance, and the iron nail becomes thinner. Mg: The solution changes from blue to green, the copper surface is covered with a red substance, magnesium strip becomes thinner and produces bubbles.

Part Two - Measuring the Direction of Electron Flow Using a Multimeter

Results



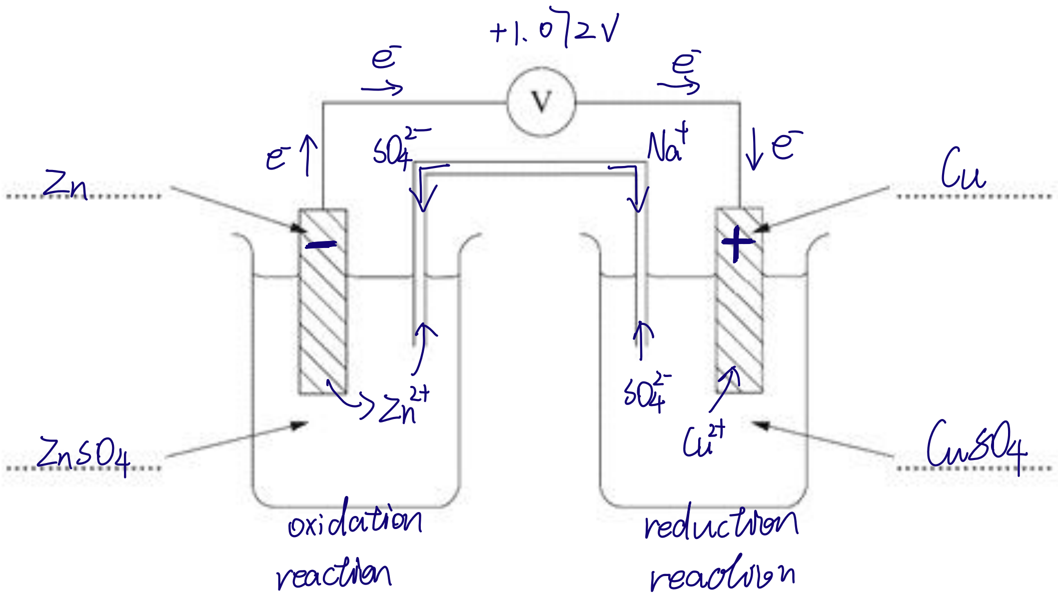


Part Three - Separating the half reactions in an electrochemical cell

Results

Table 1: the results of voltage, metal which electrons flows from, oxidation reaction and reduction reaction in different electrochemical cell

Multimeter terminals connected to the following electrodes					
Red	Black	Sign and magnitude of voltage	Electrons flowed from which metal?	What is the oxidation reaction in this cell?	What is the reduction reaction in this cell?
Cu	Fe	+0.595 V	Fe	$Fe - 2e^- \rightarrow Fe^{2+}$	$Cu^{2+} + 2e^- \rightarrow Cu$
Cu	Zn	+1.072 V	Zn	$Zn - 2e^- \rightarrow Zn^{2+}$	$Cu^{2+} + 2e^- \rightarrow Cu$
Cu	Mg	+1.649 V	Mg	$Mg - 2e^- \rightarrow Mg^{2+}$	$Cu^{2+} + 2e^- \rightarrow Cu$
Fe	Zn	+0.489 V	Zn	$Zn - 2e^- \rightarrow Zn^{2+}$	$Fe^{2+} + 2e^- \rightarrow Fe$
Fe	Mg	+1.081 V	Mg	$Mg - 2e^- \rightarrow Mg^{2+}$	$Fe^{2+} + 2e^- \rightarrow Fe$
Zn	Mg	+0.587 V	Mg	$Mg - 2e^- \rightarrow Mg^{2+}$	$Zn^{2+} + 2e^- \rightarrow Zn$



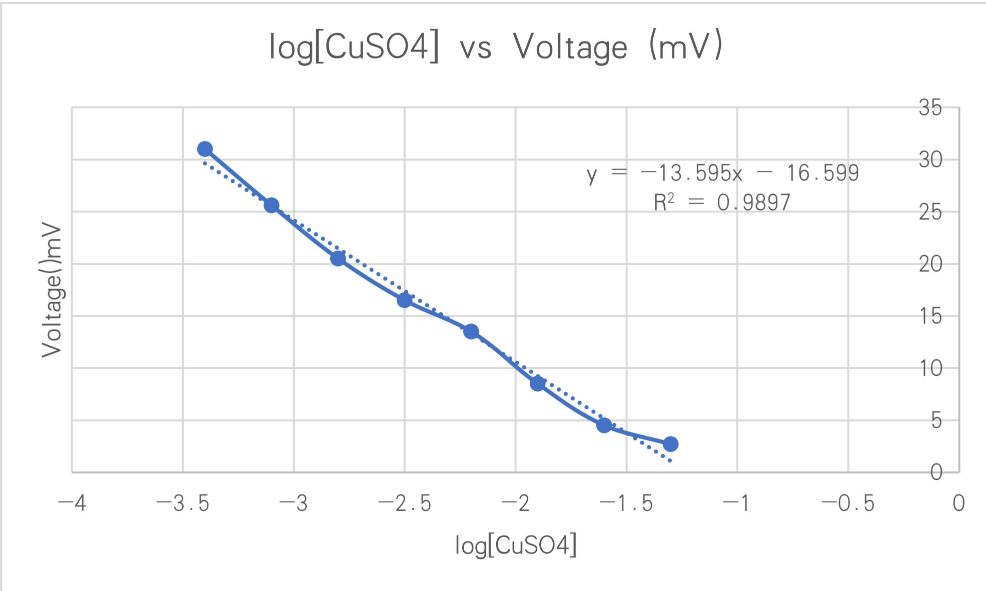
Remove the salt bridge and the voltage becomes 0. Salt Bridges cannot be replaced by wire, because the salt Bridges are used to transport ions, not electrons.
The most readily gains electrons → the least readily gains electron:
$Cu^{2+} > Fe^{2+} > Zn^{2+} > Mg^{2+}$

Part Four - Voltages due to concentration differences

Results

Table 2: results of the the voltage and log [$CuSO_4$] of each of the concentration cells

[$CuSO_4$] (M)	Voltage (mV)	log [$CuSO_4$]
0.050	2.7	-1.3
0.025	4.5	-1.6
0.013	8.5	-1.9
0.0063	13.5	-2.2
0.0031	16.5	-2.5
0.0016	20.5	-2.8
0.00078	25.6	-3.1
0.00039	31.0	-3.4



Graph 1

log [$CuSO_4$] increase, voltage decrease.
trendline equation: y = -13.595x -16.599
measure an unknown concentration of copper(II) ions in solution:
Plug the voltage into the formula,
log [unknown concentration Cu^{2+}] = (voltage + 16.599)/ (-13.595).
unknown concentration Cu^{2+} in solution=10^ (log [unknown concentration Cu^{2+}])

Part Five – Plating

Results

Table 3: results of reaction between different metal ion solution and solid metal

Metal ion solution / Solid metal	Copper	Iron	Zinc	Magnesium
Copper	-	redox reaction	redox reaction	redox reaction
Iron	Spontaneous reaction	-	redox reaction	redox reaction
Zinc	Spontaneous reaction	Spontaneous reaction	-	redox reaction
Magnesium	Spontaneous reaction	Spontaneous reaction	Spontaneous reaction	-

red cathode: zinc electrode
black anode: iron electrode
iron nail is covered with a white solid,
Anode reduction reaction: $Zn^{2+} + 2e^{-} = Zn$
Cathode oxidation reaction: $Zn - 2e^{-} = Zn^{2+}$

red cathode: iron electrode
black anode: zinc electrode
Cathode reduction reaction: $Zn^{2+} + 2e^{-} = Zn$
Anode oxidation reaction: $Zn - 2e^{-} = Zn^{2+}$