Redox Reactions and Voltaic Cells

Introduction:

The vast majority of chemistry involves the movement and redistribution of electrons. In many systems, this happens without any consequence, but if the system can be constructed such that the moving electrons are passed through an external circuit, that electron movement can be harnessed to do useful work. This is the basis of all batteries and other power sources. In this experiment, we will be exploring the behavior of various electrochemical half-cells and using our observations to rank their activity.

Safety Concerns:

The solutions used in this experiment should be handled with caution. Wear your goggles at all times in the lab. Rinse your hands thoroughly if any solution splashes or spills on you. Wipe up any spills immediately and rinse the spilled area with water. Do not pour any solutions down the drain, but instead, place them in the appropriately labeled waste containers in the hood.

Experimental Procedure:

I. Reactions of metals with metal ions

Obtain a strip of Zn, Cu, Pb, and Ag metals. If necessary, clean the surface of pieces of metal using sandpaper or steel wool. {***Do not clean electrodes directly on the benchtop!***} On each metal strip, place a single drop of each metal ion solution from the provided dropper bottles (don't let them touch!) and determine whether or not a reaction has happened. It may be useful to put a drop of water on each strip to serve as a reference.

Q: Which <u>metal</u> is most reactive? Least reactive? Describe how you determine their reactivity and rank all 4 metals from most reactive to least reactive.

Q: Which <u>metal ion</u> is most reactive? Least reactive? Describe how you determine their reactivity and rank all 4 metal ions from most reactive to least reactive.

Q: Are your rankings related? Offer an explanation of any relationship you observe.

II. Voltaic Cells

IIA. From the stock bottles on the side of the lab, fill four clean small test tubes half full of the solutions of the same four metal ions studied in part I. Fill another small test tube half full of a $Ni(NO_3)_2$ solution. Label the tubes appropriately. Place a clean strip of each metal in the solution of its corresponding ion. Each of these tubes is a M^{+x}|M-type half-cell. Coupling two half-cells with a salt bridge forms an electrochemical cell with a multimeter completing the external circuit.

IIB. Measure the cell potential of each electrochemical cell shown in the table. Connect the multimeter leads firmly on the strips of metal. Do not simply touch the leads to the metal electrodes but connect them firmly to the metal. It often helps to clean the surface of the electrode first for better electrical contact. The current passing through these cells is so low that there is no notable danger involved.

Cell	Cell
$Zn \mid Zn^{2+} \parallel Cu^{2+} \mid Cu$	$Pb \mid Pb^{2+} \parallel Ni^{2+} \mid Ni$
$Zn \mid Zn^{2+} \parallel Ni^{2+} \mid Ni$	$Pb \mid Pb^{2+} \parallel Ag^{+} \mid Ag$
$Zn \mid Zn^{2+} \parallel Ag^{+} \mid Ag$	$Ag \mid Ag^{+} \parallel Cu^{2+} \mid Cu$
$Zn \mid Zn^{2+} \parallel Pb^{2+} \mid Pb$	$Ag \mid Ag^{+} \parallel Ni^{2+} \mid Ni$
$Pb \mid Pb^{2+} \parallel Cu^{2+} \mid Cu$	$Ni \mid Ni^{2+} \parallel Cu^{2+} \mid Cu$

It is a good practice to connect the voltmeter so that the multimeter negative (black lead) is connected to the left-hand side electrode of the cells shown in the table. Being consistent will make it MUCH easier to interpret your results.

Q: What does it mean if the measured potential is negative? What does this tell you about the cells as written in the table?

IIC. Pick one of the electrochemical cells from the previous section and reassemble it with the multimeter hooked up backwards. That is, connect the negative black lead to the *right* hand electrode.

Q: How does the new measurement compare to the potential you measured in Part IIB? Explain any difference.

IID. Starting with your ***most*** reactive metal as determined in Part I, note the measured potential between it and the other three metals used in Part I. You can re-measure the potentials if you are unsure of your interpretation of the results.

Q: Is there a relationship between the ranking you determined in Part I and the potentials you measured for each pair of metals?

IIE. Compare the measured cell potentials between various pairs of metals, both those that are adjacent to each other in the ranking from Part I and those that are not. Again, you can re-measure the potentials if you are unsure of your interpretation of the results. For example, if M1 is more reactive than M2, and M2 is more reactive than M3, look at your measured cell potentials for M1-M2, M2-M3 and M1-M3.

Q: Is there a relationship between the measured cell potentials for adjacent metals and the measured cell potentials for non-adjacent pairs?

Q: Using all of your measured cell potentials, where would you place nickel in your ranking? Explain how your data supports this ranking.

III. Concentration Effects

IIIA. In this part of the experiment the $Cu^{+2}(aq)|Cu$ half-cell will be coupled with several $Ag^{+}(aq)|Ag$ half-cells in which the concentration of $Ag^{+}(aq)$ is varied systematically. Dilute 1.0 mL of the 0.1 M $Ag^{+}(aq)$ solution to 10.0 mL with distilled water to prepare a 0.01 M $Ag^{+}(aq)$ solution. You can use a graduated cylinder for this dilution; mix the dilution well and transfer to a large test tube. Perform two more serial dilutions in the same manner in order to prepare 0.001 M $Ag^{+}(aq)$ and 0.0001 M $Ag^{+}(aq)$ solutions. Place a few milliliters of each $Ag^{+}(aq)$ solution in separate clean, small test tubes.

IIIB. In the tube containing the most dilute solution {0.0001 M $Ag^{+}(aq)$} place a clean strip of Ag. Couple this half-cell to the Cu^{+2} | Cu half-cell and **measure the cell potential of the spontaneous reaction**.

IIIC. Remove the salt bridge and Ag metal strip from the $Ag^{+}(aq)$ solution, dab the salt bridge on a paper towel and then place the salt bridge and Ag metal strip into the 0.0010 M $Ag^{+}(aq)$ solution. Couple this half-cell with the copper half-cell, and measure and record the cell voltage.

IIID. Measure the cell potential with the 0.010 M $Ag^{+}(aq)$, and 0.10 M $Ag^{+}(aq)$ solutions in the same manner.

IIIE. How does the cell potential change when the concentration of $Ag^{+}(aq)$ in the half-cells changes? Prepare graphs of E_{cell} vs. $[Ag^{+}]$ and E_{cell} vs. $\log[Ag^{+}]$. Use E_{cell} for the ***spontaneous reaction*** in your graphs.

Q: How would you describe the relationship (shape of the graph) between cell potential and concentration?

Q: How would you describe the relationship (shape of the graph) between cell potential and log(concentration)?

Q: Which of these graphs shows a clearer relationship between cell voltages and concentration of $Ag^{+}(aq)$ ions?