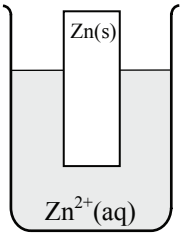
Inv 3.1 Electrochemical cells 1

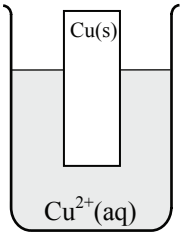
All redox reactions involve electron transfer. This can happen even when the two reactants are physically separated, say, electrons transferring from one to another through a wire. This is what happens in an electrochemical cell.

1 Clean the surface of the metal strips you use.

2 Prepare a Zn^{2+}/Zn half cell by placing a strip of zinc into a beaker of zinc sulfate solution so that the solution covers three-quarters of the zinc strip.



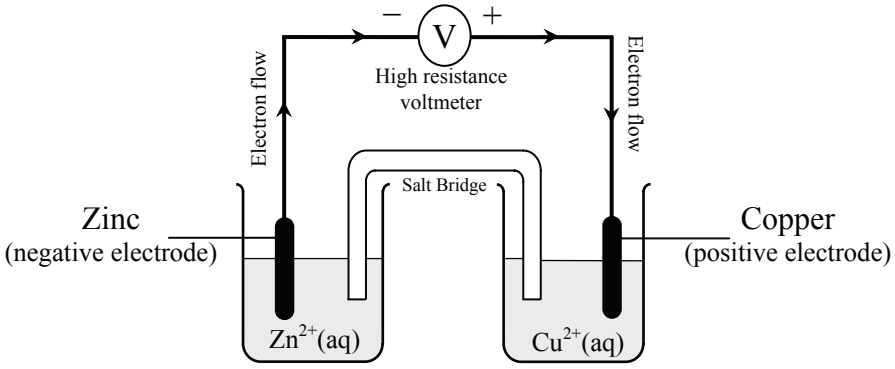
3 Prepare a Cu^{2+}/Cu half cell in a similar manner using a copper strip and copper sulfate.



4 Connect a wire from the zinc electrode (strip) to the negative terminal of a multimeter set on voltage. Record the reading on the meter. 0.00 V

5 Connect a wire from the copper electrode (strip) to the positive terminal of the meter. Record the reading on the meter. 0.0 V

6 Place the half cells next to each other and connect them via a salt bridge.



Record the reading on the meter. 1.12 V

7 Which way do electrons flow in the circuit? L to R from Zn to Cu

8 What is the cathode? RHE (copper)

9 What is the anode? LHE (zinc)

10 Write the half equation for the reaction occurring in the Zn^{2+}/Zn half cell.

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

11 Is this oxidation or reduction? oxidation

continues...

12 Write the half equation for the reaction occurring in the Cu^{2+}/Cu half cell.

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

13 Is this oxidation or reduction? reduction

14 Combine your half-equations from **11** and **13** above to give the overall cell reaction.

$Zn(s) + Cu2+(aq) \rightarrow Zn2+(aq) + Cu(s)$

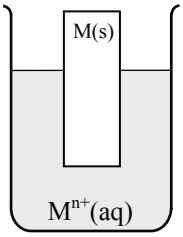
15 What is the EMF of the above cell? 1.12 V

Inv 3.2 Electrochemical cells 2

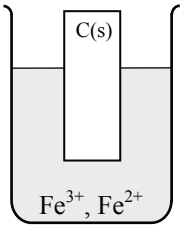
We can use the electrode potentials of redox systems to determine the relative strengths of oxidants and reductants. Your teacher will give you a range of oxidants and reductants from which you can prepare different half-cells. Electrochemical cells are made by combining two half-cells.

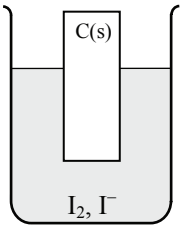
1 Clean the surface of the metal strips you use.

2 Prepare M^{n+}/M half cells by placing a strip of metal into a beaker of the metal sulfate solution so that the solution covers three-quarters of the metal strip.

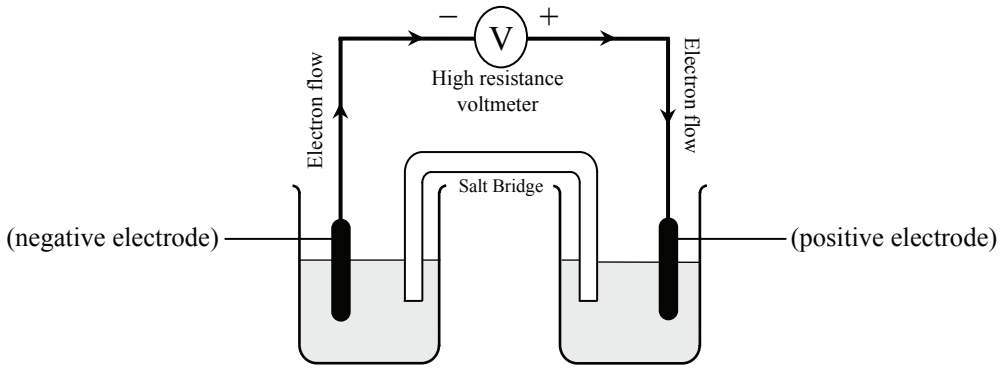


3 Prepare a Fe^{3+}/Fe^{2+} half cell using a carbon electrode in equal volumes of iron(II) sulfate and iron(III) nitrate and an I_2/I^- half cell using a carbon electrode in equal volumes of iodine solution and potassium iodide solution.





4 Prepare cells by combining two half cells, connecting the two electrodes with a voltmeter and the solutions with a salt bridge.



5 For each cell measure the voltage (EMF) and record it in the table shown along with the direction of electron flow and the oxidant and reductant for each cell.

Cell	Cell voltage		Electron flow		Oxidant	Reductant
	measured	calculated	from	to		
$Zn/Zn^{2+}/Cu^{2+}/Cu$	1.11 V	1.1 V	LHE	RHE	Cu^{2+}	Zn
$Mg/Mg^{2+}// Zn^{2+}/Zn$	0.85 V	1.6 V	LHE	RHE	$Zn2+$	Mg
$Mg/Mg^{2+}//Cu^{2+}/Cu$	1.98 V	2.7 V	LHE	RHE	Cu^{2+}	Mg
$Zn/Zn^{2+}//Fe^{3+},Fe^{2+}/C$	1.49 V	1.53 V	LHE	RHE	Fe^{2+}	Zn
$Zn/Zn^{2+}//I_2,I^-/C$	1.10 V	1.38 V	LHE	RHE	I_2	Zn

6 Calculate the cell voltages from $E°$ values from data tables and compare with the measured values. Suggest any reasons for differences.

The values will be different from the standard values if the conditions were not standard -

including carbon electrodes instead of platinum.

7 Arrange the oxidants in order of increasing ability to be reduced.

$Zn^{2+} < Cu^{2+} < I_2 < Fe^{3+}$

Inv 3.3 The lead-acid cell

One of the most useful practical electrochemical cells is the lead-acid cell used in car batteries. It is so useful because its electrode reactions can be reversed by passing an electric current back through the cell.

1 Clean the surface of two lead plates and place them in a 250 mL beaker. Add about 150 mL of about 2 mol L^{-1} of sulfuric acid.

2 Connect a voltmeter between the plates and record the voltage.

Voltage reading: 0.0 V

3 Remove the voltmeter and replace it with a DC power supply. Adjust the supply to about 3 V, switch on, wait and observe for about 10 minutes. Note any changes that take place.

Observations: Bubbles of gas are seen at the negative electrode, while the positive electrode is coated in a very dark brown substance.

4 Replace the DC supply with the voltmeter and measure the voltage. (Make sure of the polarity.)

Voltage reading: 1.99 V

5 Replace the voltmeter with a 1.5 V bulb and observation:

The bulb lights up.

During the charging of the cell the lead, Pb(s), of the cathode is converted to lead (IV) oxide, $PbO_2(s)$.

6 Write equations for the reactions that take place at each electrode during discharge.

Anode: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$

Cathode: $PbO_2(s) + 4H^+(aq) + 2e^- \rightarrow Pb^{2+}(aq) + 2H_2O(l)$