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Electrochemistry

BATTERIES are portable power devices, essential in our every day life. They power cellphones and even cars. Batteries use the principles of chemistry to produce electricity. Galvanic cells are textbook batteries, not intended to generate electricity as they function reversibly and in equilibrium. In the eighteenth century, Luigi Galvani discovered that animals' muscles—in particular dead frogs—could be artificially moved by touching the muscles with rods of different metals. The generated electricity at first was believed to come from the muscles. However, Alessandro Volta proved that the source of electricity in muscle movement was indeed the metals. Volta created the first Voltaic pile by stacking metallic silver and zinc disks separated by paper soaked in saltwater. This early discovery jumpstarted electrochemistry, a new field of chemistry. Indeed, chemical reactions can produce electricity, and electricity can drive chemical reactions.



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GOALS

- 1 Identify anodes/cathodes
- 2 Calculate cell potentials
- 3 Interpret the line notation
- 4 Calculate cell potentials of concentration cells
- 5 Relate cell potential with ΔG°

1.1 Introduction to galvanic cells

Galvanic cells—also known as voltaic cells or piles—are electrochemical cells that generate electricity from spontaneous redox reactions. They contain two different metals immersed in electrolyte solutions. Each half-cell contains a metal in contact with a liquid solution of the same metal in ionic form. Both half-cells are either connected by a salt bridge or separated by a porous membrane. Galvanic cells differ from batteries. Batteries are composed of multiple single cells working out of equilibrium while producing electricity. Galvanic cells are textbook batteries. They are reversible devices unable to produce electricity.

⚙️ **Discussion:** What is the difference between a battery and a galvanic cell?

Components of a galvanic cell Galvanic cells are composed of two different electrodes, an anode, and a cathode, connected by means of a salt bridge or a membrane. The role of the salt bridge or membrane is to complete the electrical circuit. Anodes are sources of electrons, whereas cathodes are electron sinks. At the same time, anodes generate positive cations, whereas cathodes generate negative anions. The role of the salt bridge or membrane is also to allow the charge generated in the cathode to be compensated by the charges generated in the anode. Electrodes contain two different redox states of the same element in contact with each other. An example of an electrode would be a piece of metallic copper Cu in contact with a solution of $\text{Cu}_{(\text{aq})}^{2+}$ ions. However, electrodes are not always made of metals. For example, electrodes can contain gas in contact (e.g. H_2) with an electrolyte solution (e.g. $\text{H}_{(\text{aq})}^+$). Electrodes without a metal being directly involved in the redox reaction need to include an external metal to support the charge transfer. Metals such as Pt are normally used for this

purpose.

The electrodes: anode and cathode Every galvanic cell is composed of two electrodes, an anode, and a cathode. Electrodes produce ionic and electronic charges. The oxidation occurs on the anode which is indicated with a negative (−) sign. Electrons are being produced in the anode resulting from an oxidation reaction. The reduction occurs on the cathode, indicated with a positive (+) sign. Electrons are being consumed in the cathode resulting from a reduction reaction. These electrodes also produce ions, in particular cations and anions which have a tendency to migrate inside the cell. Anodes generate cations (and consume anions, depending on the chemical reaction involved), whereas cathodes generate anions (and consume cations, again, depending on the chemical reaction involved). As such, anions have a tendency to migrate to the anode, as their concentration is lower there, whereas cations have a tendency to migrate to the cathode. The excess ionic charge is compensated in the interface between the electrode and the salt bridge or the membrane. Mind that in a galvanic cell only electrons flow through the circuit, by means of the wire connecting both electrodes. The ions involved have a tendency to migrate inside each of the electrodes but do not leave the electrode. The name of the electrodes—anode and cathode—results from the ionic flow involved in the galvanic cell: anions have a tendency to migrate towards the anode and cations to the cathode.

Cell potential Water flows down a waterfall due to the difference of potential energy between the high and low parts of the waterfall. Similarly, heat flows between a hot and a cold reservoir due to the difference in temperature between both locations. The force that drives the flow is heat is temperature. Electricity flows through a galvanic cell resulting from the difference of cell potential $\Delta\mathcal{E}$ between both electrodes, the anode, and the cathode. The cell potential—also referred to as cell voltage, cell electromotive force, or cell *emf*—is the force that drives the flow of electrons. Anodes and cathodes have a characteristic cell potential associated with the electrochemical half-reaction happening in the electrode. The voltage of the anode (\mathcal{E}_{anode}) is always lower than the one from the cathode ($\mathcal{E}_{cathode}$). The combination of the anodic and cathodic voltage gives the overall cell potential measured in a galvanic cell. In particular, the overall voltage results from the voltage of the cathode with respect to the anode, so that the overall voltage of a galvanic cell is always positive.

Role of the salt bridge or the membrane The role of the salt bridge or the porous membrane is to compensate for the excess of ions generated in each electrode hence closing the electric circuit. Salt bridges contain saturated solutions of electrolytes containing ions with similar ionic mobility (KCl or NH_4NO_3). These electrolytes are also non-reactive with the chemicals involved in the galvanic cell. Each side of the salt bridge or membrane becomes charged due to ionic accumulation, with negative ions accumulating near the cathode and positive charges near the anode. Porous membranes impact the galvanic cell potential with an extra contribution called the liquid junction voltage. This voltage is due to the ion accumulation on both sides of the membrane. When using a salt bridge, the liquid junction voltage on the left side of the bridge compensates the salt bridge on the right side of the bridge so that overall the galvanic cell potential remains unaffected by the bridge.

A galvanic cell example Below we display a representation of the Daniell cell, a classical galvanic cell in which copper is oxidized by zinc. The name of the cell is in honor of John Daniell, a British chemist from the nineteenth century who was trying to develop an electric power supply to sustain telegraphy. He connected with a metallic

wire a zinc electrode in contact with a zinc sulfate solution to a copper electrode in contact with a copper (II) sulfate solution. In this cell, Copper(II) ions are converted into metallic copper in the cathode by means of the reaction $\text{Cu}_{(\text{aq})}^{2+} + 2\text{e}^{-} \longrightarrow \text{Cu}_{(\text{s})}$, whereas metallic zinc is converted into zinc ions in the anode following the reaction $\text{Zn}_{(\text{s})} \longrightarrow \text{Zn}_{(\text{aq})}^{2+} + 2\text{e}^{-}$. Copper(II) ions are being reduced whereas zinc is being oxidized. The cathode solution becomes negatively charged whereas the anode solution becomes positively charged. A porous membrane permeable to ions was used to avoid the charge buildup.

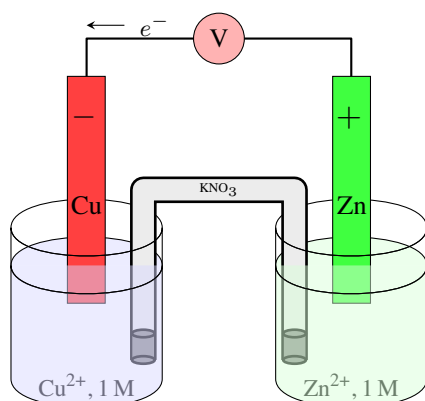
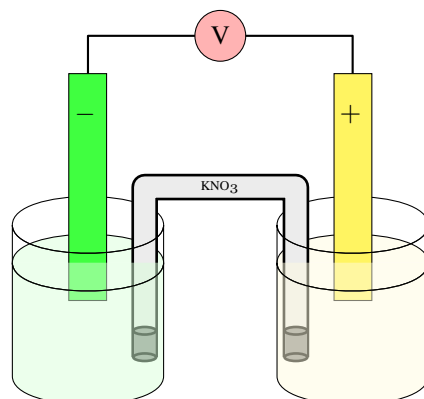


Figure 1.1 The Daniell galvanic cell

The potentiometer Voltmeters are devices used to measure cell potential, the number of Volts, in electric circuits. These devices work by drawing current through a known resistance. Voltmeter can not be used to precisely measure the voltage of a galvanic cell. As electricity flows into the voltmeter, frictional heating will occur and energy will be wasted. Hence the voltage measured would be lower than the real voltage. In order to avoid this problem in a lab setting we normally use potentiometers to measure galvanic cells. Potentiometers apply a counter-voltage to compensate the cell voltage without drawing any significant current. Due to their high internal resistivity, there is no electricity flow in a voltaic cell connected to a voltmeter. The cell remains in equilibrium, and the electrodes are not consumed. Still, an ammeter could be used to measure the tendency of the cell to generate a measurable intensity flow, the Amperes.

Sample Problem 1

For the galvanic cell below, indicate: (a) label the electrode as anode and cathode (b) identify the flow of electrons (c) identify the flow of cations and anions (d) identify the oxidation and reduction

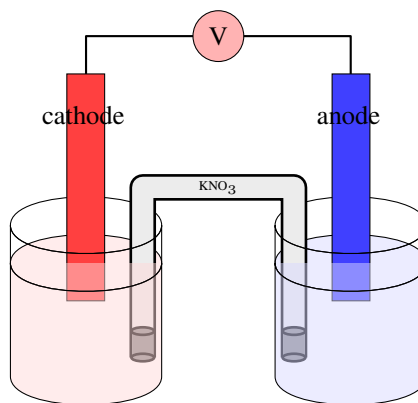


SOLUTION

(a) The electrode labeled with the $-$ sign located on the left is the anode, where the oxidation takes place. Electrons are being produced in the anode and cations are also being generated in the anode (or perhaps anions are being consumed, it depends on the redox reaction happening). The electrode labeled with the $+$ sign located on the right is the cathode, where the reduction takes place. Electrons are consumed in the cathode and anions are also being generated in the cathode. (b) The flow of electrons goes from the anode on the left to the cathode on the right (c) Cations are being produced in the anode and they have a tendency to migrate towards the cathode. Anions are being produced in the cathode and they have a tendency to migrate towards the anode. (d) The oxidation takes place on the anode on the left, whereas the reduction takes place on the cathode on the right.

STUDY CHECK

For the galvanic cell below, indicate: (a) label the signs of the electrodes (b) identify the flow of electrons (c) identify the flow of cations and anions (d) identify the oxidation and reduction



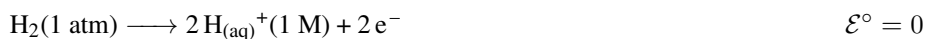
1.2 Standard reduction potentials

In a galvanic cell, each electrode—also called half-cell—has a given potential. When combining two electrodes, we obtain the measurable cell potential. This magnitude represents the force that pushed electrons from the anode to the cathode producing a measurable current. This section covers electrode potentials. We will define the concept of electrode potential, and we will identify the anode and cathode when two electrodes are in contact based on the electrode potentials.

Electrode potentials A galvanic cell consists of two electrodes, an anode, and a cathode. Each electrode contains two chemical species with different redox numbers in contact through an interface. For example, we can build up an electrode by soaking a piece of metallic copper on a solution of copper(II) sulfate. The interface consists of the liquid phase containing $\text{Cu}_{(\text{aq})}^{2+}$ in contact with the metal phase made of $\text{Cu}_{(\text{s})}$. Each electrode has a reduction potential— \mathcal{E} , expressed in volts (V)—that informs about the drive of the redox process in the electrode. The larger this value the stronger the tendency of the redox process to occur.

Standard conditions for reduction potentials The potential of a sin-

gle electrode cannot be accurately measured as electrodes only exist in the context of a two-electrode galvanic cell. Single electrode potentials are defined in a galvanic cell made of the electrode and a reference electrode with null potential under certain conditions. The standard hydrogen electrode (SHE) is universally accepted as the reference electrode in electrochemistry, hence having null potential. The hydrogen electrode contains gas hydrogen in contact with an acidic 1M HCl solution at 25°C with a wire made of platinum—an inert metal—that mediates the electron transfer. Below you can find the reaction involved in the hydrogen electrode:



The voltage of a galvanic cell made of an electrode combined with the reference hydrogen electrode will directly measure the electrode potential. Still, electrode potentials depend on concentration (or pressure for gases) and temperature conditions. The standard conditions for electrode potentials are a molar concentration of 1M for all electrolytes and pressure of 1atm. The electrode potential measured at these conditions is called standard potential \mathcal{E}° , where the symbol $^\circ$ represents standard state. At the same time, every redox reaction can be written as a reduction or an oxidation reaction. As we tabulate standard electrode potentials, all reactions will be written as reduction reactions and the standard potentials are called standard reduction potential \mathcal{E}° .

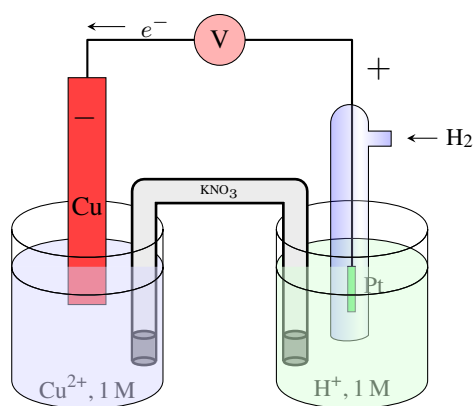


Figure 1.2 A galvanic cell with a hydrogen anode and a copper cathode. The hydrogen electrode is the reference electrode with null electrodic voltage. Hence the voltage of this cell will directly give the copper electrode voltage.

Anodes and cathodes The standard reduction potential of an electrode predicts the tendency of an electrode to act as an anode or cathode. Imagine we have two different electrodes involving the following reaction with the standard potentials indicated on the side, and we need to set up a galvanic cell:



How to determine which electrode will act as an anode and which will act as the cathode? The rule is the smaller the electrode potential the more tendency of the electrode to act as an anode. If we compare the copper and zinc electrodes, as the electrode potential of zinc is smaller—more negative—than the electrode potential of copper. Therefore, zinc will act as an anode and copper will act as a cathode. The

standard reduction potential is not affected by the stoichiometry of the reaction. For example, the standard reduction potential of cesium is -3.03V :



At the same time, the standard reduction potential of two moles of cesium is still -3.03V .



Differently, inverting the reduction reaction switches the sign of the reduction potential. For example, the standard reduction potential of Titanium(II) is -1.63V

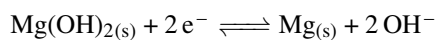


while the standard potential of oxidation of Titanium into Titanium(II) is $+1.63\text{V}$



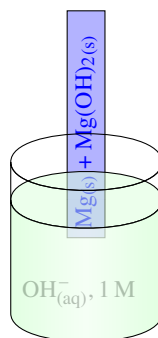
Sample Problem 2

Sketch a semi-cell for the semi reaction below, indicate if the electrode is a solid-liquid electrode or gas-liquid electrode. Assume standard conditions:



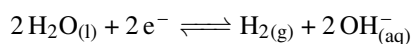
SOLUTION

This is a solid-liquid electrode. The solid piece would be made of Magnesium covered with magnesium hydroxide. The liquid phase would contain a base (OH^-). At standard conditions all molarities would be 1M .

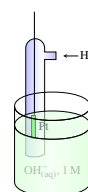


STUDY CHECK

Sketch a semi-cell for the semi reaction below, indicate if the electrode is a solid-liquid electrode or gas-liquid electrode. Assume standard conditions:



►Answer:

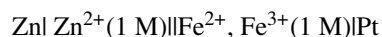


1.3 Line notation for galvanic cells

There is a quick and easy way to represent a galvanic cell without having to draw the whole cell set up. This is called the line notation and galvanic cells are represented in a single line, starting from left to right. The anode is presented in the left, starting from the metallic part and followed by the electrolyte. A single line represents the liquid-solid contact. A double line represents the salt bridge and the cathode is represented in the right, starting for the electrolyte and finishing by the metal. As you can see, the line notation respect all interphase present in the cells: from left to right we have solid, liquid in contact with the salt bridge which is in contact with the liquid part of the cathode and finally we have the solid part of the cathode. For example, the line notation of Daniell cell is:



In case there are several electrolytes in any of the electrodes, as all species are in liquid phase we separate them with just a comma. For example, in the galvanic cell below the cathode contains two different states of iron and uses Pt for the charge transfer:



1.4 Standard reduction potential table

The reduction potential of most of the electrodes are tabulated in standard conditions. This section will cover how to employ the table of standard potentials in order to obtain the electrode potential for a given electrode.

Standard electrode potentials are expressed as reduction

Electrodes potentials can be expressed as reduction or as oxidation. This is because all electrodes can behave as anode or cathode, depending on the conterelectrode used in the galvanic cell. For example, for the $\text{Zn(s)}/\text{Zn}_{(\text{aq})}^{2+}$ the reduction potential is:



whereas the oxidation potential is:



In the table of standard potentials all reactions are expressed only as reduction and as such all potentials are reduction potentials. As such, all reaction will have electrons on the left side of the arrow. The electrode potential you will find on the table for the $\text{Zn(s)}/\text{Zn}_{(\text{aq})}^{2+}$ would be



Table 1.1 Standard reduction potentials at 298K

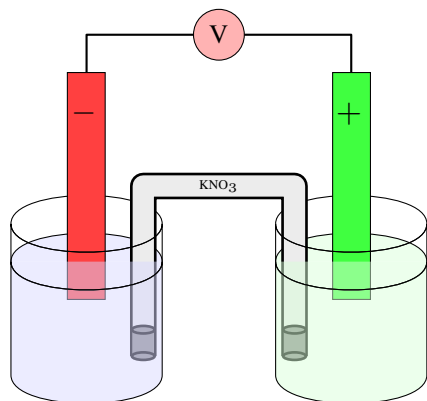
Element	Reaction	\mathcal{E}° (V)	Element	Reaction
Sr	$\text{Sr}^+ + \text{e}^- \rightleftharpoons \text{Sr}_{(\text{s})}$	-4.10	H	$2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_{2(\text{g})}$
Ca	$\text{Ca}^+ + \text{e}^- \rightleftharpoons \text{Ca}_{(\text{s})}$	-3.80	Ag	$\text{AgBr}_{(\text{s})} + \text{e}^- \rightleftharpoons \text{Ag}_{(\text{s})} + \text{Br}^-$
Li	$\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}_{(\text{s})}$	-3.04	S	$\text{S}_4\text{O}_6^{6-} + 2\text{e}^- \rightleftharpoons 2\text{S}_2\text{O}_3^{2-}$
Cs	$\text{Cs}^+ + \text{e}^- \rightleftharpoons \text{Cs}_{(\text{s})}$	-3.03	N	$\text{N}_{2(\text{g})} + 2\text{H}_2\text{O} + 6\text{H}^+ + 6\text{e}^- \rightleftharpoons 2\text{NH}_4\text{OH}_{(\text{aq})}$
Ca	$\text{Ca}(\text{OH})_2 + 2\text{e}^- \rightleftharpoons \text{Ca}_{(\text{s})} + 2\text{OH}^-$	-3.02	Hg	$\text{HgO}_{(\text{s})} + \text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{Hg}_{(\text{l})} + 2\text{OH}^-$
Ba	$\text{Ba}(\text{OH})_2 + 2\text{e}^- \rightleftharpoons \text{Ba}_{(\text{s})} + 2\text{OH}^-$	-2.99	C	$\text{C}_{(\text{s})} + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{CH}_{4(\text{g})}$
Rb	$\text{Rb}^+ + \text{e}^- \rightleftharpoons \text{Rb}_{(\text{s})}$	-2.98	Sn	$\text{Sn}^{4+} + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+}$
K	$\text{K}^+ + \text{e}^- \rightleftharpoons \text{K}_{(\text{s})}$	-2.93	Cu	$\text{Cu}^{2+} + \text{e}^- \rightleftharpoons \text{Cu}^+$
Ba	$\text{Ba}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ba}_{(\text{s})}$	-2.91	Fe	$3\text{Fe}_2\text{O}_3_{(\text{s})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{Fe}_3\text{O}_4_{(\text{s})} + \text{H}_2\text{O}$
Sr	$\text{Sr}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sr}_{(\text{s})}$	-2.90	Ag	$\text{AgCl}_{(\text{s})} + \text{e}^- \rightleftharpoons \text{Ag}_{(\text{s})} + \text{Cl}^-$
Sr	$\text{Sr}(\text{OH})_2 + 2\text{e}^- \rightleftharpoons \text{Sr}_{(\text{s})} + 2\text{OH}^-$	-2.88	Cu	$\text{Cu}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cu}_{(\text{s})}$
Ca	$\text{Ca}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ca}_{(\text{s})}$	-2.87	Fe	$\text{Fe}^+ + \text{e}^- \rightleftharpoons \text{Fe}_{(\text{s})}$
Li	$\text{Li}^+ + \text{C}_{6(\text{s})} + \text{e}^- \rightleftharpoons \text{LiC}_{6(\text{s})}$	-2.84	O	$\text{O}_{2(\text{g})} + 2\text{H}_2\text{O} + 4\text{e}^- \rightleftharpoons 4\text{OH}^-_{(\text{aq})}$
Na	$\text{Na}^+ + \text{e}^- \rightleftharpoons \text{Na}_{(\text{s})}$	-2.71	Cu	$\text{Cu}^+ + \text{e}^- \rightleftharpoons \text{Cu}_{(\text{s})}$
Mg	$\text{Mg}(\text{OH})_2 + 2\text{e}^- \rightleftharpoons \text{Mg}_{(\text{s})} + 2\text{OH}^-$	-2.69	C	$\text{CO}_{(\text{g})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{C}_{(\text{s})} + \text{H}_2\text{O}$
Mg	$\text{Mg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mg}_{(\text{s})}$	-2.37	I	$\text{I}_{2(\text{s})} + 2\text{e}^- \rightleftharpoons 2\text{I}^-$
H	$\text{H}_{2(\text{g})} + 2\text{e}^- \rightleftharpoons 2\text{H}^-$	-2.23	Mn	$\text{MnO}_4^- + 2\text{H}_2\text{O} + 3\text{e}^- \rightleftharpoons \text{MnO}_{2(\text{s})} + 4\text{OH}^-$
Sr	$\text{Sr}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sr}_{(\text{Hg})}$	-1.79	O	$\text{O}_{2(\text{g})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}_{2(\text{aq})}$
Al	$\text{Al}^{3+} + 3\text{e}^- \rightleftharpoons \text{Al}_{(\text{s})}$	-1.66	Fe	$\text{Fe}_2\text{O}_3_{(\text{s})} + 6\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{Fe}^{2+} + 3\text{H}_2\text{O}$
Ti	$\text{Ti}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ti}_{(\text{s})}$	-1.63	Fe	$\text{Fe}^{3+} + \text{e}^- \rightleftharpoons \text{Fe}^{2+}$
Ti	$\text{Ti}^{3+} + 3\text{e}^- \rightleftharpoons \text{Ti}_{(\text{s})}$	-1.37	Ag	$\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}_{(\text{s})}$
Ti	$\text{TiO}_{(\text{s})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{Ti}_{(\text{s})} + \text{H}_2\text{O}$	-1.31	Hg	$\text{Hg}_2^{2+} + 2\text{e}^- \rightleftharpoons 2\text{Hg}_{(\text{l})}$
Mn	$\text{Mn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Mn}_{(\text{s})}$	-1.18	N	$\text{NO}_3^-_{(\text{aq})} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{NO}_{2(\text{g})} + \text{H}_2\text{O}$
V	$\text{V}^{2+} + 2\text{e}^- \rightleftharpoons \text{V}_{(\text{s})}$	-1.13	Fe	$2\text{FeO}_2^{2-} + 5\text{H}_2\text{O} + 6\text{e}^- \rightleftharpoons \text{Fe}_2\text{O}_3_{(\text{s})} + 10\text{OH}^-$
Ti	$\text{TiO}^{2+} + 2\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{Ti}_{(\text{s})} + \text{H}_2\text{O}$	-0.93	Hg	$\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}_{(\text{l})}$
Si	$\text{SiO}_{2(\text{s})} + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{Si}_{(\text{s})} + 2\text{H}_2\text{O}$	-0.91	Mn	$\text{MnO}_4^- + \text{H}^+ + \text{e}^- \rightleftharpoons \text{HMnO}_4^-$
Fe	$\text{Fe}_2\text{O}_3_{(\text{s})} + 3\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons 2\text{Fe}(\text{OH})_{2(\text{s})} + 2\text{OH}^-$	-0.86	Hg	$2\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}_2^{2+}$
H	$2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{H}_{2(\text{g})} + 2\text{OH}^-$	-0.828	Pd	$\text{Pd}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pd}_{(\text{s})}$
Zn	$\text{Zn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Zn}_{(\text{s})}$	-0.762	N	$\text{NO}_3^-_{(\text{aq})} + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{NO}_{(\text{g})} + 2\text{H}_2\text{O}_{(\text{l})}$
Cr	$\text{Cr}^{3+} + 3\text{e}^- \rightleftharpoons \text{Cr}_{(\text{s})}$	-0.74	Fe	$\text{Fe}_3\text{O}_4_{(\text{s})} + 8\text{H}^+ + 2\text{e}^- \rightleftharpoons 3\text{Fe}^{2+} + 4\text{H}_2\text{O}$
Ni	$\text{Ni}(\text{OH})_{2(\text{s})} + 2\text{e}^- \rightleftharpoons \text{Ni}_{(\text{s})} + 2\text{OH}^-$	-0.72	Br	$\text{Br}_{2(\text{aq})} + 2\text{e}^- \rightleftharpoons 2\text{Br}^-$
Ag	$\text{Ag}_2\text{S}_{(\text{s})} + 2\text{e}^- \rightleftharpoons 2\text{Ag}_{(\text{s})} + \text{S}_{2-}^-(\text{aq})$	-0.69	Ag	$\text{Ag}_2\text{O}_{(\text{s})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{Ag}_{(\text{s})} + \text{H}_2\text{O}$
Pb	$\text{PbO}_{(\text{s})} + \text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{Pb}_{(\text{s})} + 2\text{OH}^-$	-0.58	Pt	$\text{Pt}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pt}_{(\text{s})}$
Fe	$\text{Fe}^{2+} + 2\text{e}^- \rightleftharpoons \text{Fe}_{(\text{s})}$	-0.44	Cl	$\text{ClO}^- + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{ClO}_3^- + \text{H}_2\text{O}$
Cr	$\text{Cr}^{3+} + \text{e}^- \rightleftharpoons \text{Cr}_2^{3+}$	-0.42	O	$\text{O}_{2(\text{g})} + 4\text{H}^+ + 4\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$
Cd	$\text{Cd}^{2+} + 2\text{e}^- \rightleftharpoons \text{Cd}_{(\text{s})}$	-0.40	Cl	$\text{Cl}_{2(\text{g})} + 2\text{e}^- \rightleftharpoons 2\text{Cl}^-$
Cu	$\text{Cu}_2\text{O}_{(\text{s})} + \text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons 2\text{Cu}_{(\text{s})} + 2\text{OH}^-$	-0.36	Br	$\text{BrO}_3^- + 5\text{H}^+ + 4\text{e}^- \rightleftharpoons \text{HBrO}_{(\text{aq})} + 2\text{H}_2\text{O}$
Pb	$\text{PbSO}_4_{(\text{s})} + 2\text{e}^- \rightleftharpoons \text{Pb}_{(\text{s})} + \text{SO}_4^{2-}$	-0.36	Br	$2\text{BrO}_3^- + 12\text{H}^+ + 10\text{e}^- \rightleftharpoons \text{Br}_{2(\text{l})} + 6\text{H}_2\text{O}$
Pb	$\text{PbSO}_4_{(\text{s})} + 2\text{e}^- \rightleftharpoons \text{Pb}(\text{Hg}) + \text{SO}_4^{2-}$	-0.35	Cl	$2\text{ClO}_3^- + 12\text{H}^+ + 10\text{e}^- \rightleftharpoons \text{Cl}_{2(\text{g})} + 6\text{H}_2\text{O}$
Co	$\text{Co}^{2+} + 2\text{e}^- \rightleftharpoons \text{Co}_{(\text{s})}$	-0.28	Mn	$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightleftharpoons \text{Mn}^{2+} + 4\text{H}_2\text{O}$
Ni	$\text{Ni}^{2+} + 2\text{e}^- \rightleftharpoons \text{Ni}_{(\text{s})}$	-0.25	Au	$\text{Au}^{3+} + 3\text{e}^- \rightleftharpoons \text{Au}_{(\text{s})}$
As	$\text{As}_{(\text{s})} + 3\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{AsH}_{3(\text{g})}$	-0.23	Pb	$\text{Pb}^{4+} + 2\text{e}^- \rightleftharpoons \text{Pb}^{2+}$
Ag	$\text{AgI}_{(\text{s})} + \text{e}^- \rightleftharpoons \text{Ag}_{(\text{s})} + \text{I}^-$	-0.15	Mn	$\text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{MnO}_{2(\text{s})} + 2\text{H}_2\text{O}$
Sn	$\text{Sn}^{2+} + 2\text{e}^- \rightleftharpoons \text{Sn}_{(\text{s})}$	-0.13	Ag	$\text{AgO}_{(\text{s})} + 2\text{H}^+ + \text{e}^- \rightleftharpoons \text{Ag}^+ + \text{H}_2\text{O}$
Pb	$\text{Pb}^{2+} + 2\text{e}^- \rightleftharpoons \text{Pb}_{(\text{s})}$	-0.126	O	$\text{H}_2\text{O}_{2(\text{aq})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{H}_2\text{O}$
C	$\text{CO}_{2(\text{g})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{HCOOH}_{(\text{aq})}$	-0.11	Au	$\text{Au}^+ + \text{e}^- \rightleftharpoons \text{Au}_{(\text{s})}$
C	$\text{CO}_{2(\text{g})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{CO}_{(\text{g})} + \text{H}_2\text{O}$	-0.11	Ag	$\text{Ag}^{2+} + \text{e}^- \rightleftharpoons \text{Ag}^+$
Fe	$\text{Fe}_3\text{O}_4_{(\text{s})} + 8\text{H}^+ + 8\text{e}^- \rightleftharpoons 3\text{Fe}_{(\text{s})} + 4\text{H}_2\text{O}$	-0.08	Mn	$\text{HMnO}_4^- + 3\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{MnO}_{2(\text{s})} + 2\text{H}_2\text{O}$
Fe	$\text{Fe}^{3+} + 3\text{e}^- \rightleftharpoons \text{Fe}_{(\text{s})}$	-0.04	Fe	$\text{FeO}_2^{4-} + 8\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{Fe}_3^{3+} + 4\text{H}_2\text{O}$
H	$2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_{2(\text{g})}$	0.00	F	$\text{F}_{2(\text{g})} + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons 2\text{HF}_{(\text{aq})}$

Increasing oxidizing strength (decreasing reducing strength)

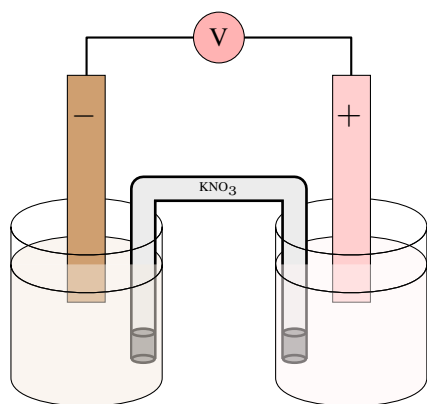
CHAPTER 1

INTRODUCTION TO GALVANIC CELLS

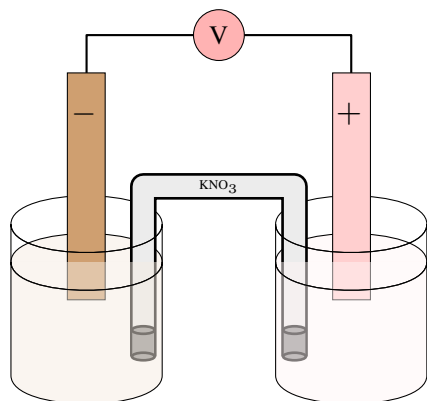
1.1 For the galvanic cell below, indicate: (a) the direction of flow of electrons (b) the direction of flow of cations (c) the direction of flow of anions



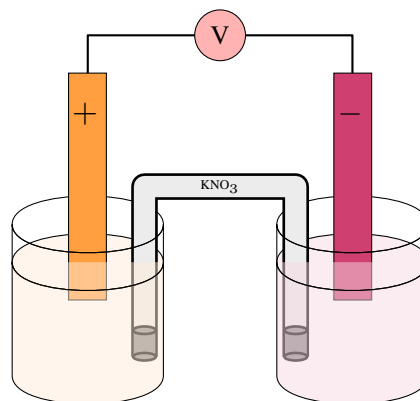
1.2 For the galvanic cell below, indicate: (a) the direction of flow of electrons (b) the direction of flow of cations (c) the direction of flow of anions



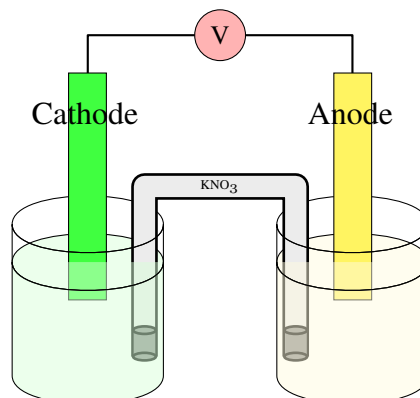
1.3 For the galvanic cell below, indicate: (a) label the anode (b) label the cathode



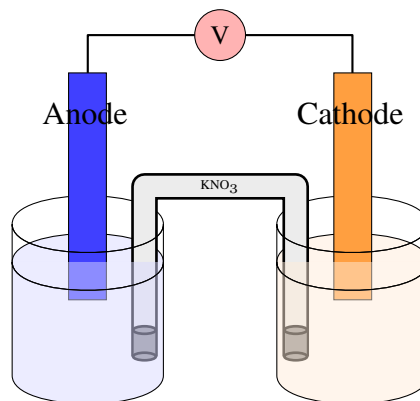
1.4 For the galvanic cell below, indicate: (a) label the anode (b) label the cathode



1.5 For the galvanic cell below, indicate: (a) label the sign (– or +) of each electrode (b) identify the flow of electrons

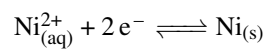


1.6 For the galvanic cell below, indicate: (a) label the sign (– or +) of each electrode (b) identify the flow of electrons

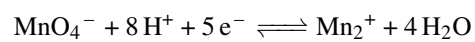


STANDARD REDUCTION POTENTIALS

1.7 Sketch a semi-cell for the semi reaction below, indicate if the electrode is a solid-liquid electrode or gas-liquid electrode. Assume standard conditions:



1.8 Sketch a semi-cell for the semi reaction below, indicate if the electrode is a solid-liquid electrode or gas-liquid electrode. Assume standard conditions:



Answers 1.1 (a) left to right (b) left to left side of salt bridge (c) right to right side of salt bridge **1.2** (a) right to left (b) right to right side of salt bridge (c) left to left side of salt bridge **1.3** (a) left (b) right **1.4** (a) right (b) left **1.5** (a) left + , right – (b) from right to left **1.6** (a) right + , left – (b) from left to right **1.7** A solid-liquid electrode

1.8 A solid-liquid electrode

