Electrochemistry

Part Three: Electrochemical Cells (Batteries)

Electrochemistry

Electrochemical Cells

- The world's first truly handheld mobile phone was invented in 1973 by Martin Cooper of Motorola.
- The phone weighed 1.1 kg and measured 23 cm long, 13 cm deep and 4.45 cm wide.
- It was a prototype offering 30 minutes of talk time and took 10 hours to recharge.

Electrochemistry
Electrochemical Cells

- Many of the improvements that have been made to mobile phones since 1973 have only been possible due to developments in the batteries that they use.
- Scientists are constantly trying to develop batteries that are smaller, lighter, charge faster, last for longer, and are environmentally friendly.

Electrochemistry

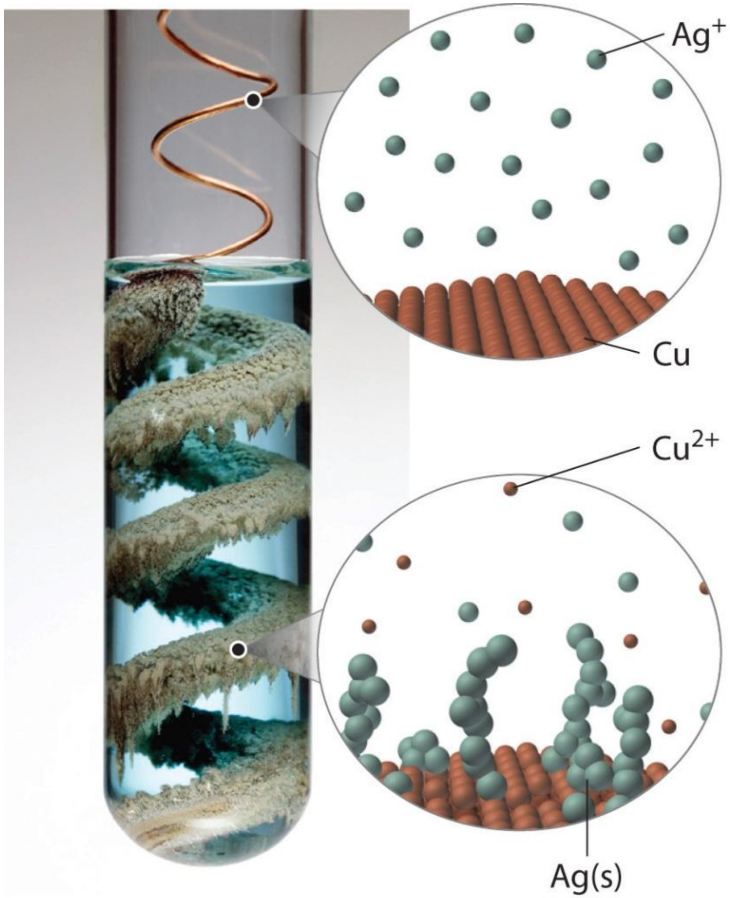
Electrochemical Cells

What happens in an electrochemical cell?
How do they generate electricity?

Electrochemistry
Electrochemical Cells

- What would you observe when a piece of copper wire is placed in a test tube containing an aqueous solution of silver nitrate?

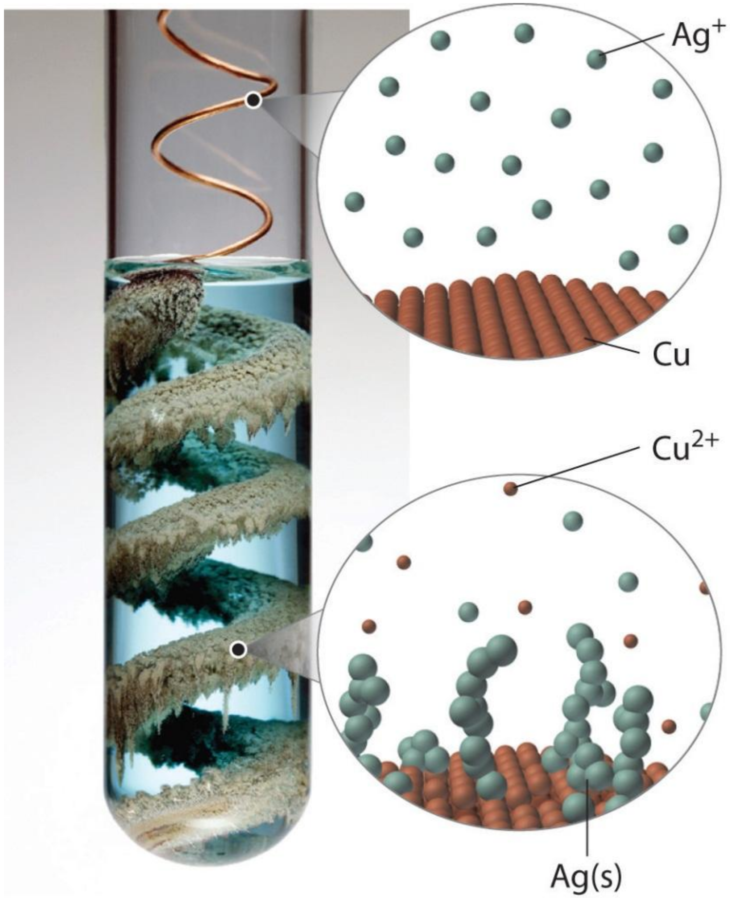
Electrochemistry
Electrochemical Cells



- Copper is more reactive than silver, and is therefore placed above silver in the reactivity series.
- Copper will displace silver from the silver nitrate.

Electrochemistry

Electrochemical Cells



- This is a *redox reaction*:

$$Cu(s) + 2AgNO_3(aq) \downarrow Cu(NO_3)_2(aq) + 2Ag(aq)$$

- The ionic equation is:

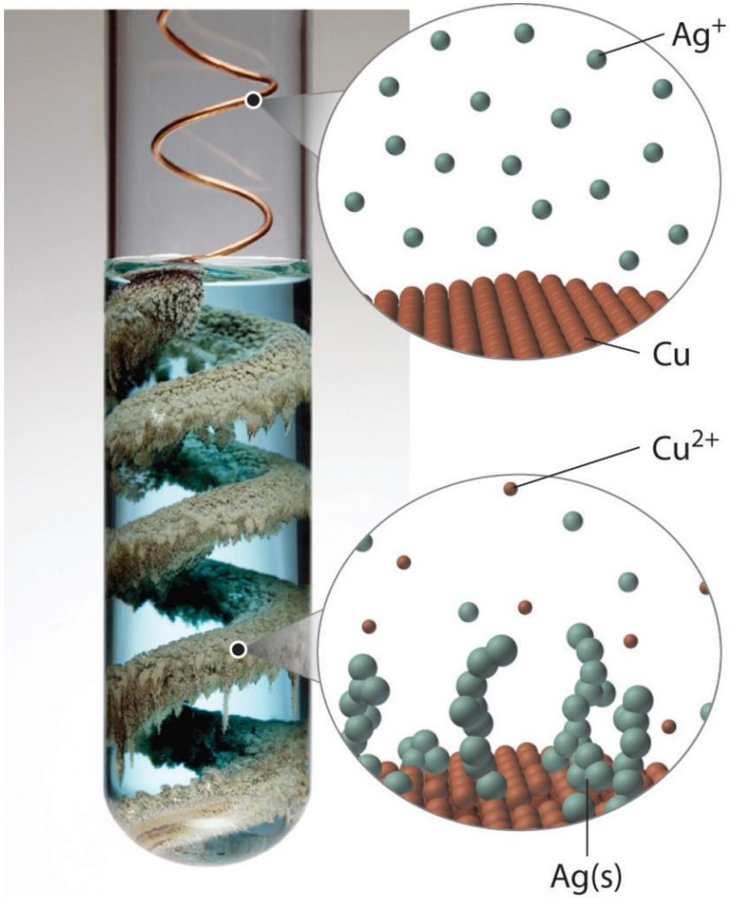
$$Cu(s) + 2Ag^+(aq) \downarrow Cu^{2+}(aq) + 2Ag(s)$$

- The ionic half-equations are:

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

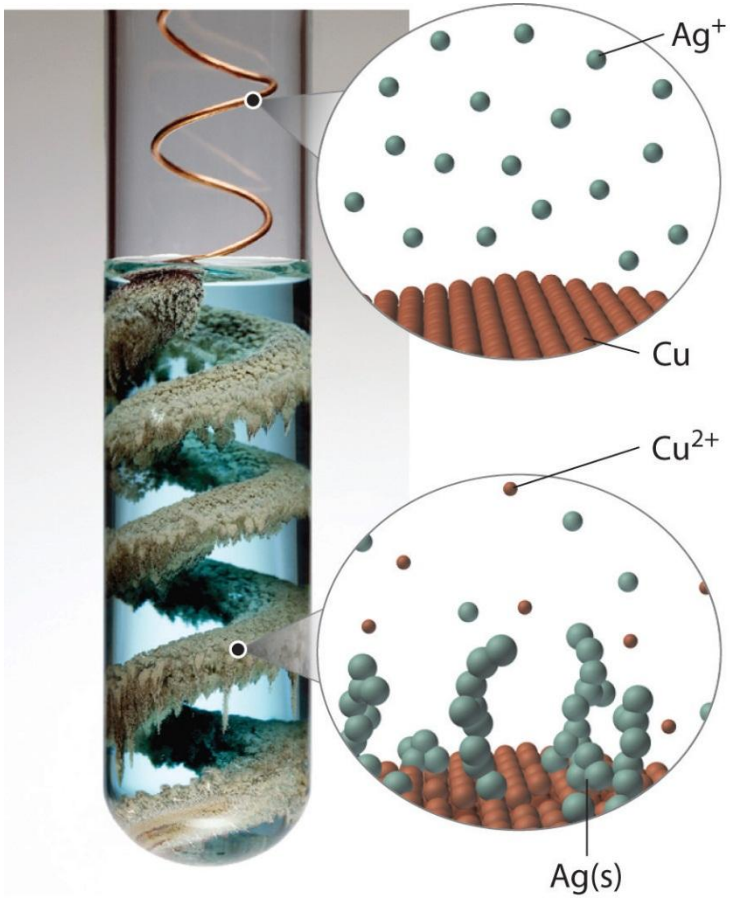
$$2Ag^+(aq) + 2e^- \rightarrow 2Ag(s)$$

Electrochemistry
Electrochemical Cells



- Copper atoms are *oxidised* to copper(II) ions (loss of electrons, increase in oxidation state from 0 to +2).
- Copper is the *reducing agent*, it reduces the silver ions to elemental silver.
- As copper(II) ions move into the aqueous solution, the solution turns *blue*.

Electrochemistry
Electrochemical Cells



- Silver ions are *reduced* to silver atoms (gain of electrons, decrease in oxidation state from +1 to 0).
- Silver is the *oxidising agent*. It oxidises the copper atoms to copper(II) ions.
- Crystals of silver form over the surface of the copper.

Electrochemistry

Electrochemical Cells

- Now imagine that the reaction for the copper, and the reaction for the silver, take place in *two separate beakers*.

- Electrons can no longer be transferred directly from the copper atoms to the silver ions. Instead, the *electrons must travel through wires* to get from one metal to the other.

- This is a *simple battery*.

Combination of Copper and Silver Half-cells

- Atoms of the more reactive metal (copper) are oxidized:

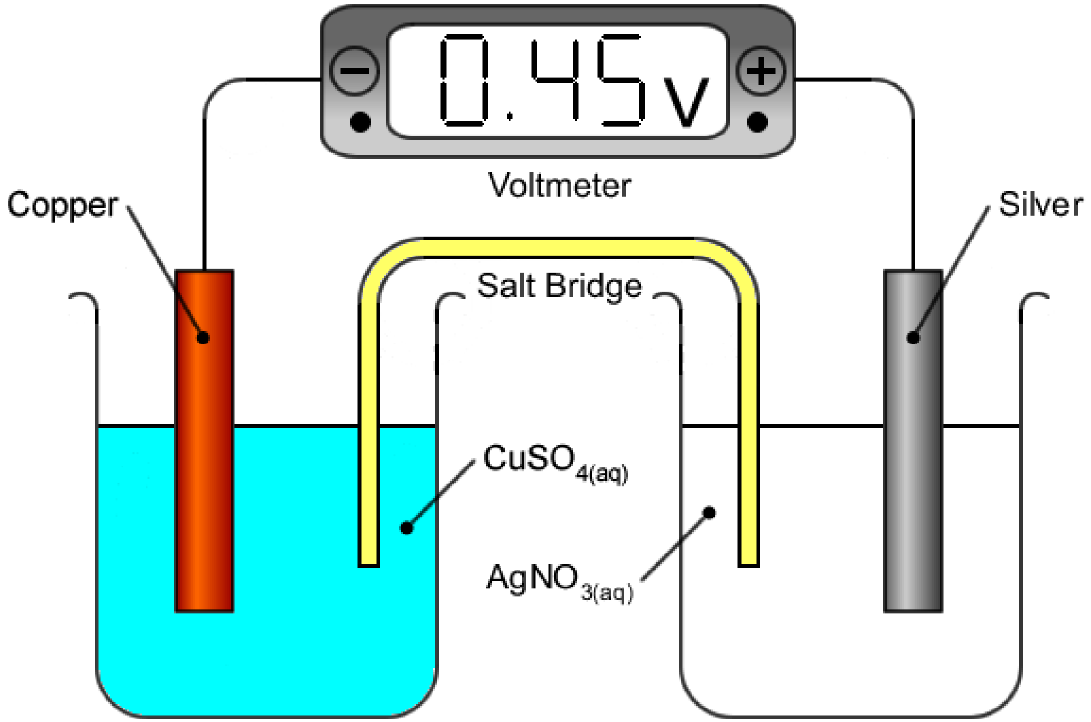
$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$$

- Ions of the less reactive metal (silver ions) are reduced:

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

Anode

The anode is defined as the electrode at which oxidation takes place.



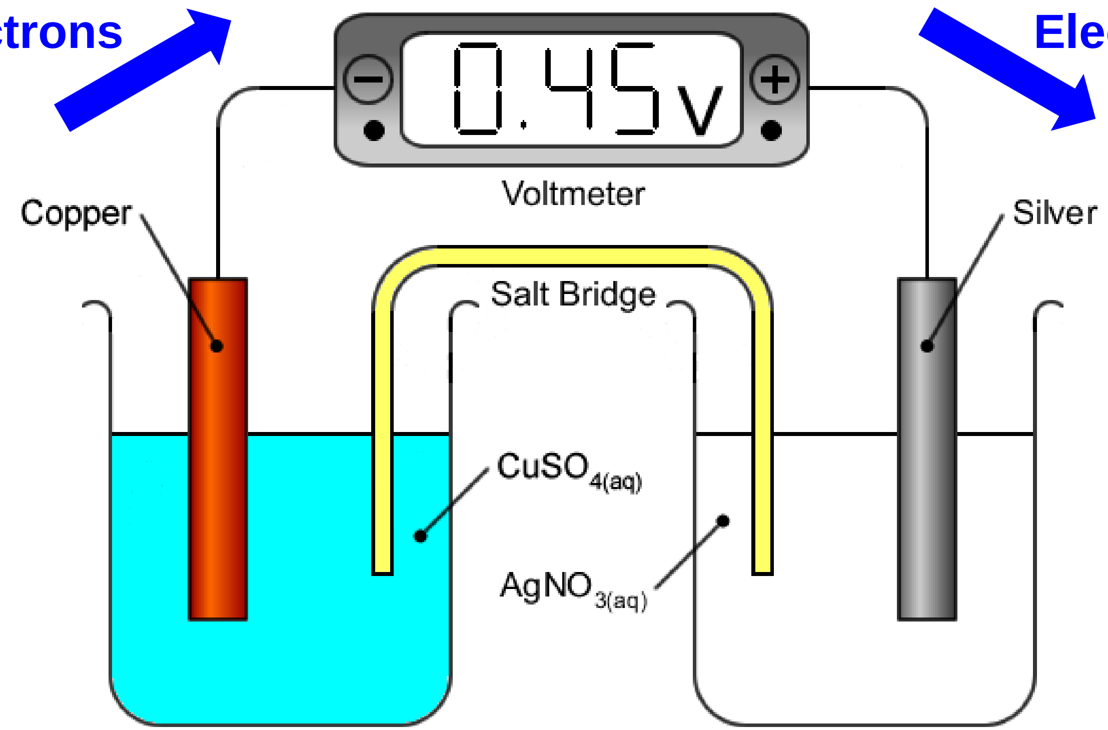
Cathode

The cathode is defined as the electrode at which reduction takes place.

Combination of Copper and Silver Half-cells

- Electrons flow through the external circuit from the more reactive metal (anode: site of oxidation) to the less reactive metal (cathode: site of reduction). Because negatively charged electrons will flow from negative to positive, this means that the anode is the negative electrode while the cathode is the positive electrode!

Electrons → **Electrons**



Negative
Anode

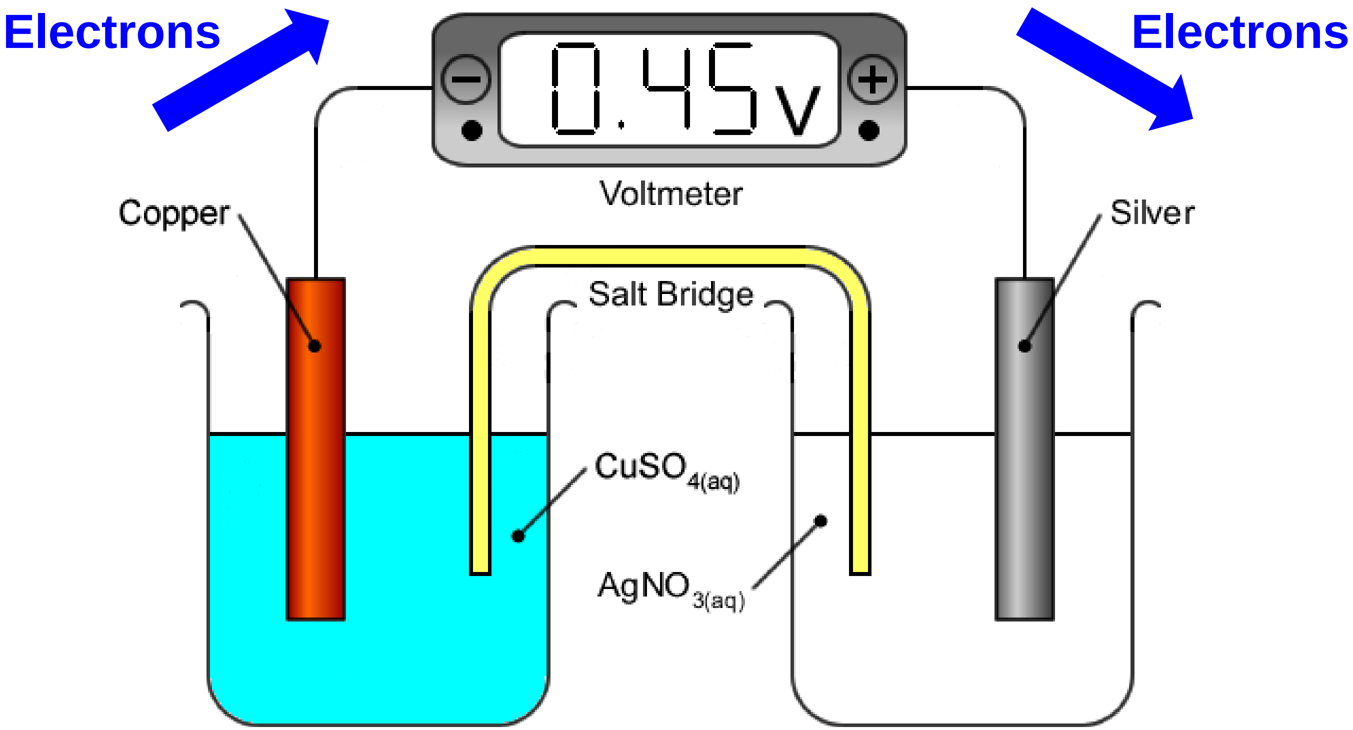
The anode is defined as the electrode at which oxidation takes place.

Positive
Cathode

The cathode is defined as the electrode at which reduction takes place.

Combination of Copper and Silver Half-cells

- **Note:** The reactions that take place in an electrochemical cell are *exothermic*.
- Electrical *energy is produced* by the chemical system.



Combination of Zinc and Copper Half-cells

- Atoms of the more reactive metal (zinc) are oxidized:

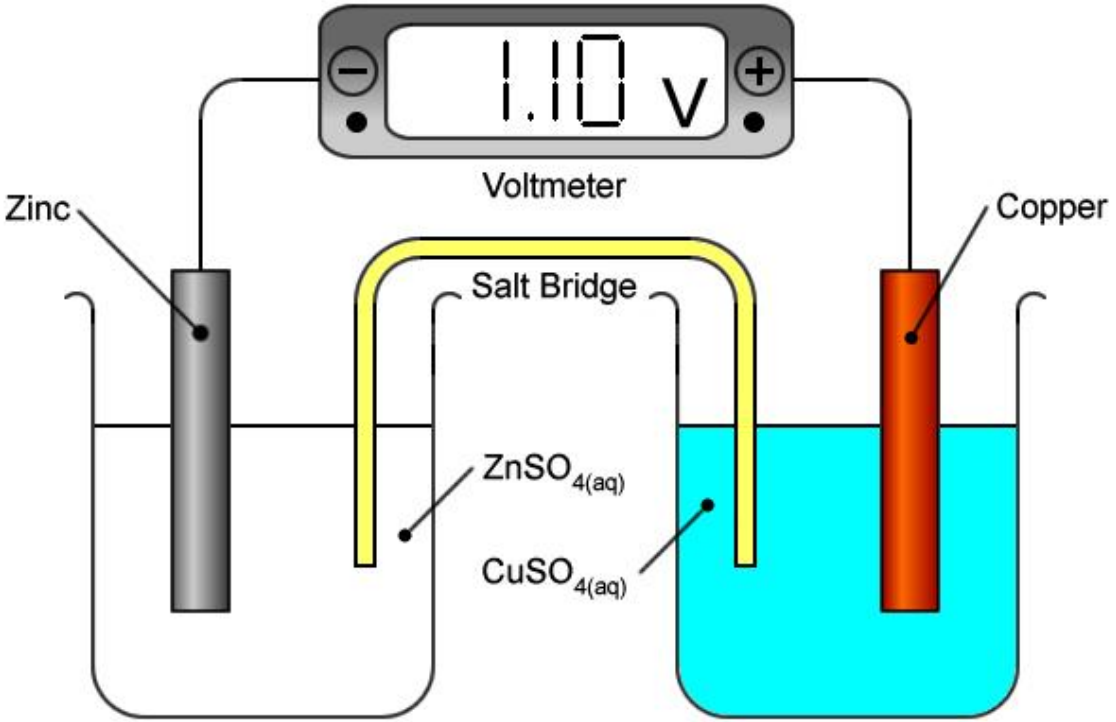
$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$$

- Ions of the less reactive metal (copper(II) ions) are reduced:

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

Anode

The anode is defined as the electrode at which oxidation takes place.



Cathode

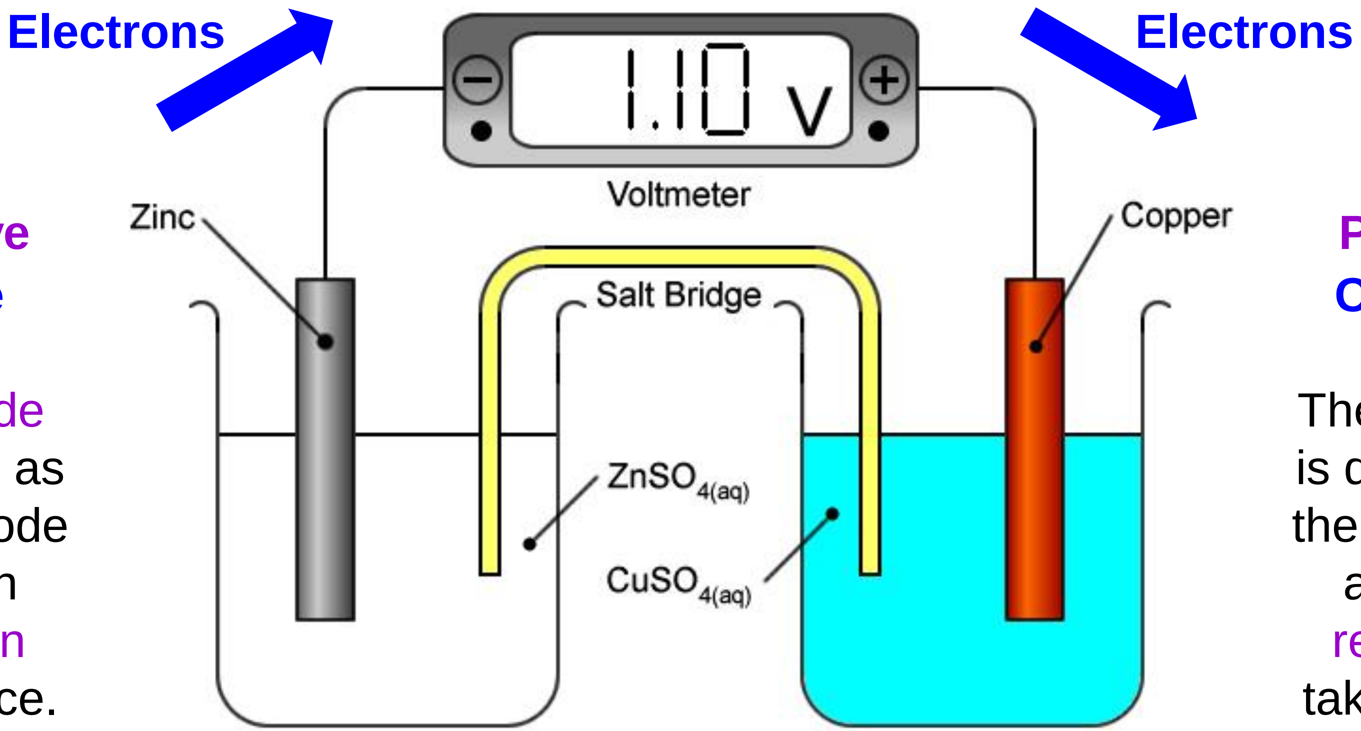
The cathode is defined as the electrode at which reduction takes place.

Combination of Zinc and Copper Half-cells

- Electrons flow through the external circuit from the more reactive metal (anode: site of oxidation) to the less reactive metal (cathode: site of reduction). Because negatively charged electrons will flow from negative to positive, this means that the anode is the negative electrode while the cathode is the positive electrode!

Negative
Anode

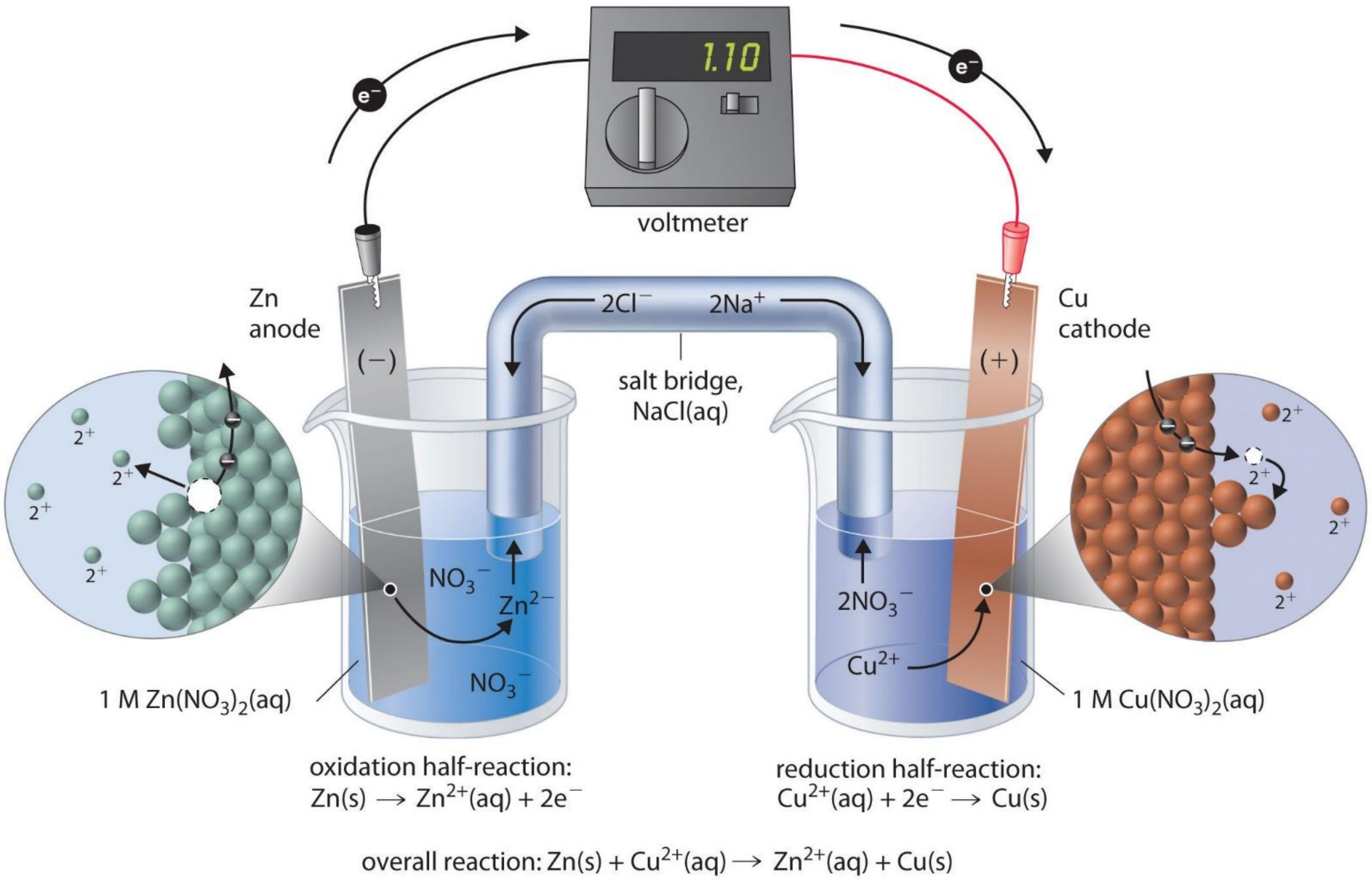
The anode is defined as the electrode at which oxidation takes place.



Positive
Cathode

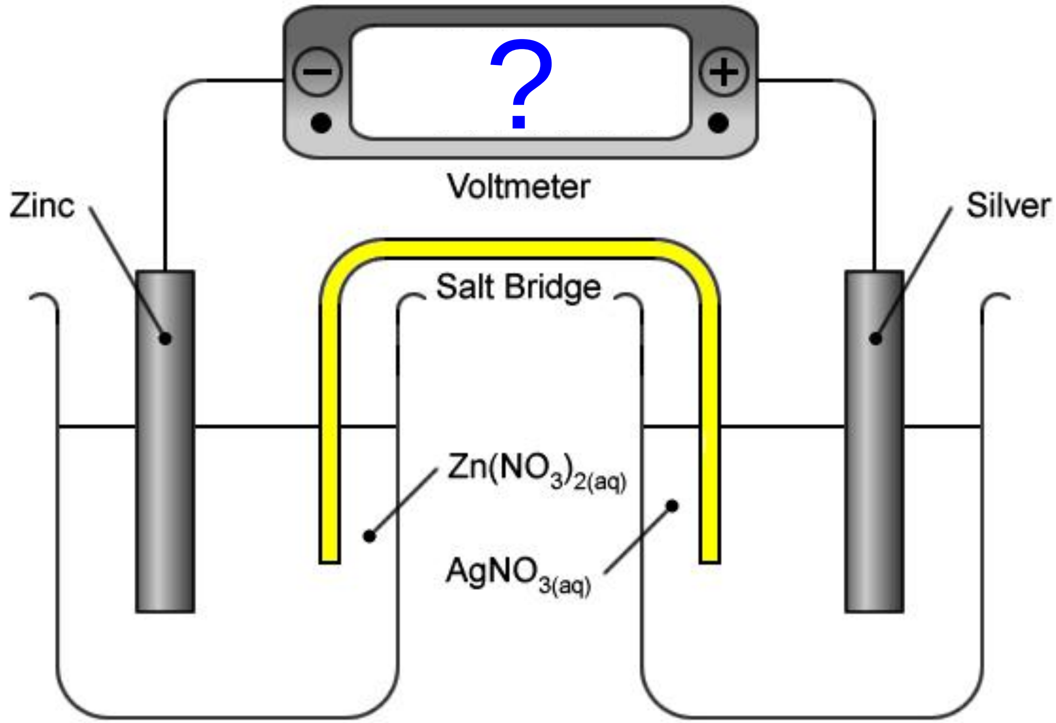
The cathode is defined as the electrode at which reduction takes place.

Combination of Zinc and Copper Half-cells



Combination of Zinc and Silver Half-cells

- The greater the separation of the metals in the electrochemical series, the greater the potential difference that is generated between their half-cells, and vice-versa.



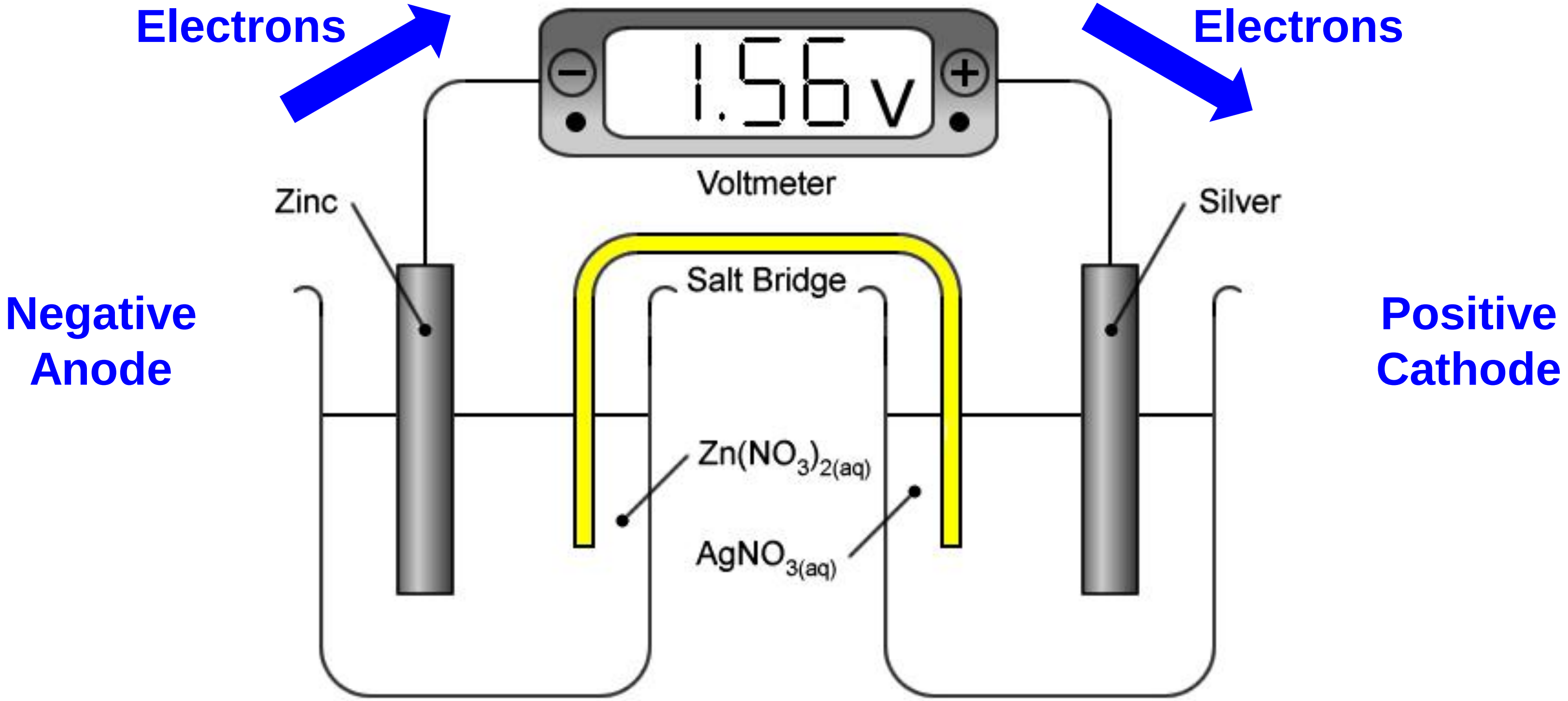
Combination of Zinc and Silver Half-cells

- Atoms of the more reactive metal (zinc) are oxidized:

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

- Ions of the less reactive metal (silver ions) are reduced:

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



Electrochemistry
Electrochemical Cells

- Potassium – K / K^+
- Sodium – Na / Na^+
- Calcium – Ca / Ca^{2+}
- Magnesium – Mg / Mg^{2+}
- Aluminium – A*l* / Al^{3+}
- Zinc – Zn / Zn^{2+}
- Iron – Fe / $Fe^{2+ / 3+}$
- Lead – Pb / Pb^{2+}
- Copper – Cu / Cu^{2+}
- Silver – Ag / Ag^+

- The greater the separation of the two metals in the electrochemical series, the greater the potential difference generated by the cell.

Electrochemistry
Electrochemical Cells

- Potassium – K / K^{+}
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- Calcium – Ca / Ca^{2+}
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- Zinc – Zn / Zn^{2+}
- Iron – Fe / $Fe^{2+ / 3+}$
- Lead – Pb / Pb^{2+}
- Copper – Cu / Cu^{2+}
- Silver – Ag / Ag^{+}

- The greater the separation of the two metals in the electrochemical series, the greater the potential difference generated by the cell.

e.g. Fe / Cu = 0.80 V

Electrochemistry

Electrochemical Cells

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- Zinc – Zn / Zn^{2+}
- Iron – Fe / $Fe^{2+\,/\,3+}$
- Lead – Pb / Pb^{2+}
- Copper – Cu / Cu^{2+}
- Silver – Ag / Ag^{+}

- The greater the separation of the two metals in the electrochemical series, the greater the potential difference generated by the cell.

Electrochemistry

Electrochemical Cells

- Potassium – K / K^{+}
- Sodium – Na / Na^{+}
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- Magnesium – Mg / Mg^{2+}
- Aluminium – A*l* / Al^{3+}
- Zinc – Zn / Zn^{2+}
- Iron – Fe / $Fe^{2+\,/\,3+}$
- Lead – Pb / Pb^{2+}
- Copper – Cu / Cu^{2+}
- Silver – Ag / Ag^{+}

- The greater the separation of the two metals in the electrochemical series, the greater the potential difference generated by the cell.

e.g. Mg / Cu = 2.70 V

Electrochemistry

- Simple batteries in a nutshell…

→ The more reactive metal is oxidised. This is the anode.

→ Ions of the less reactive metal are reduced. This is the cathode.

→ Electrons flow through the external circuit from the more reactive metal to the less reactive metal.

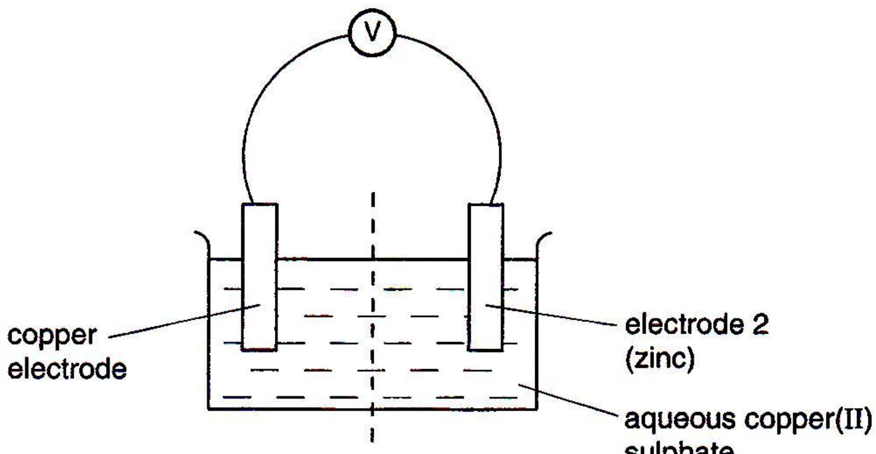
→ The further apart the metals are in the electrochemical series, the greater the voltage.

Electrochemistry

Electrochemical Cells

Could I please have some questions to check my understanding of electrochemical cells?

Electrochemistry
Electrochemical Cells



Question.

- The voltage of the cell was measured when the following metals were used for electrode 2.

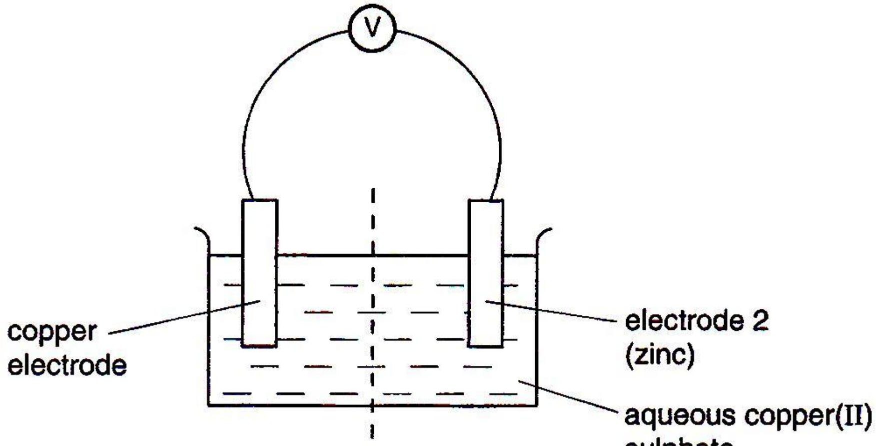
Copper Iron Lead Zinc

Complete the table by entering the metals in the correct order.

Meter Reading / V	Metal
1.10	
0.78	
0.21	
0.00	

Electrochemistry

Electrochemical Cells



Question.

- The voltage of the cell was measured when the following metals were used for electrode 2.

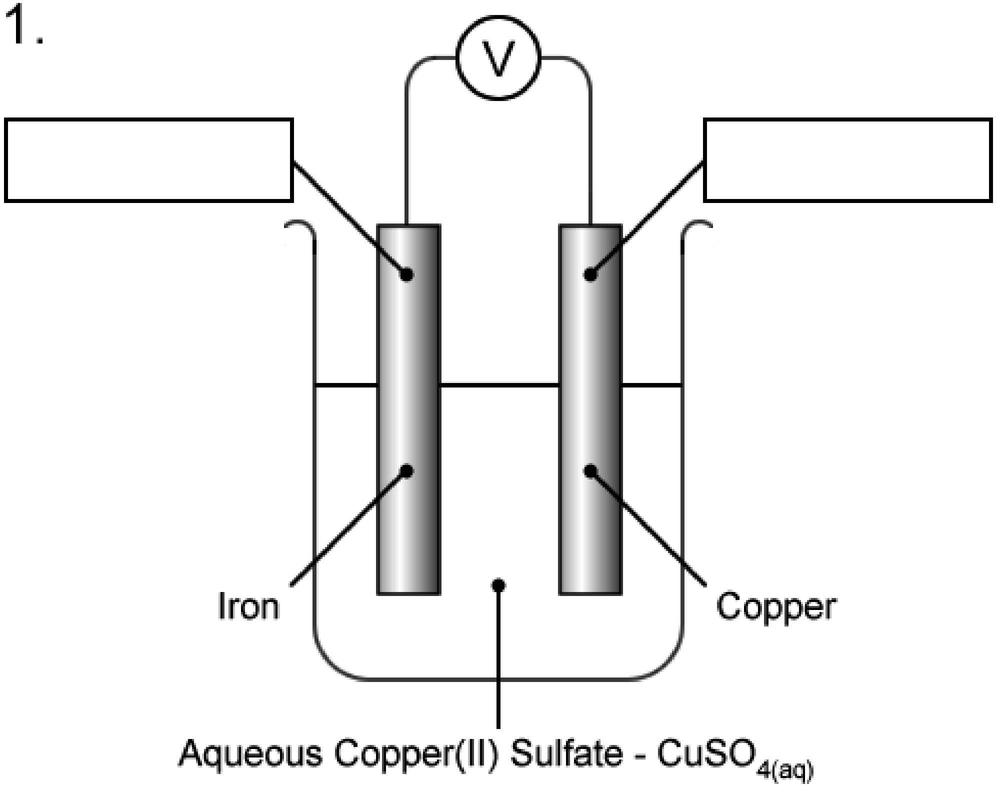
Copper Iron Lead Zinc

Complete the table by entering the metals in the correct order.

Meter Reading / V	Metal
1.10	Zinc
0.78	Iron
0.21	Lead
0.00	Copper

Electrochemistry
Electrochemical Cells

1.



(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

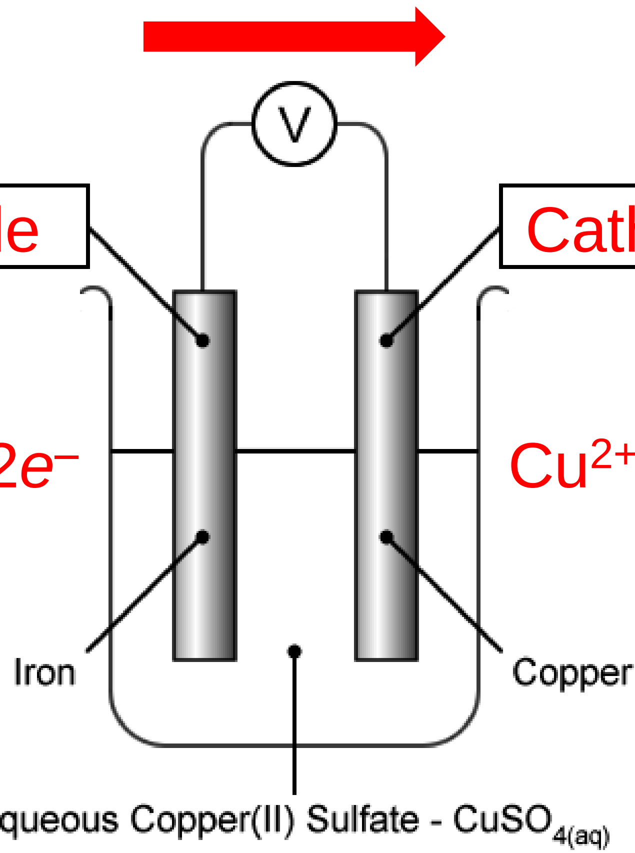
Electrochemistry

Electrochemical Cells

1.

(Oxidation) Anode

Cathode (Reduction)



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

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

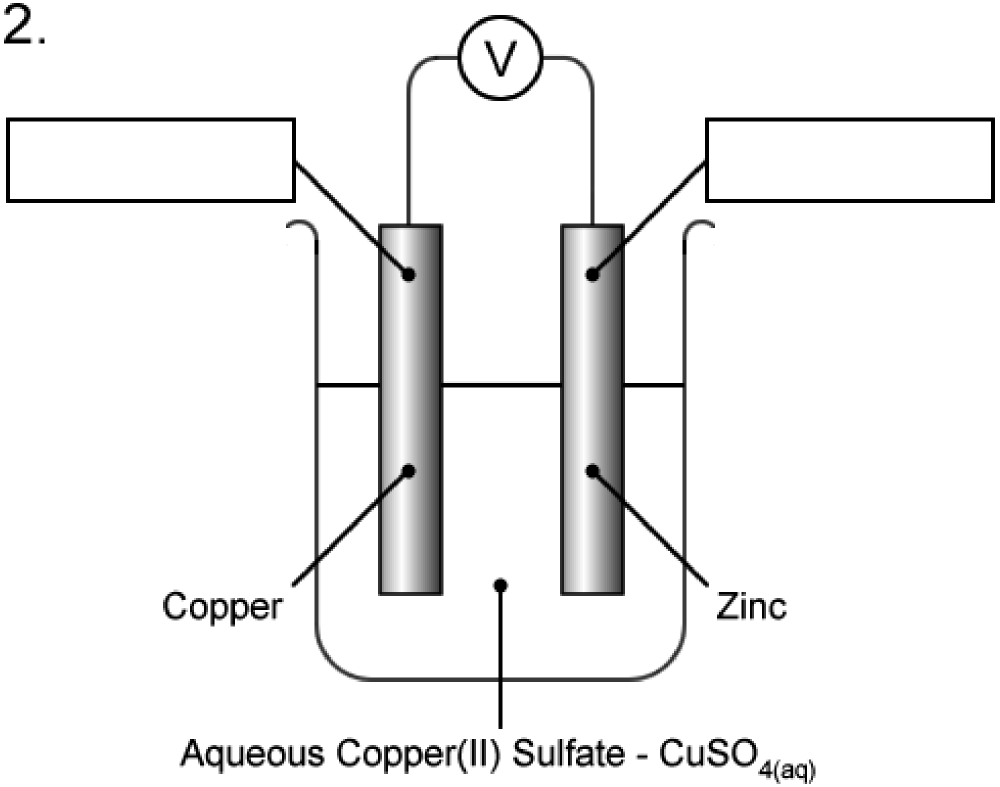
(a) Label the anode and the cathode.

(b) Write ionic half-equations for the reactions at the anode and cathode.

(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry
Electrochemical Cells

2.

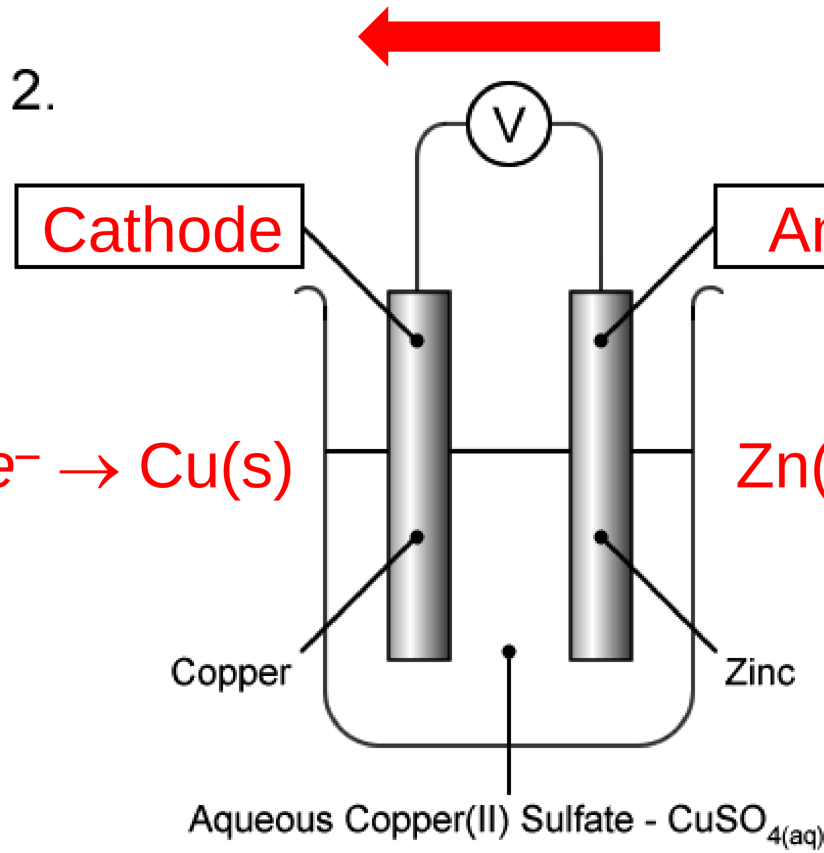


(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry

Electrochemical Cells

(Reduction)



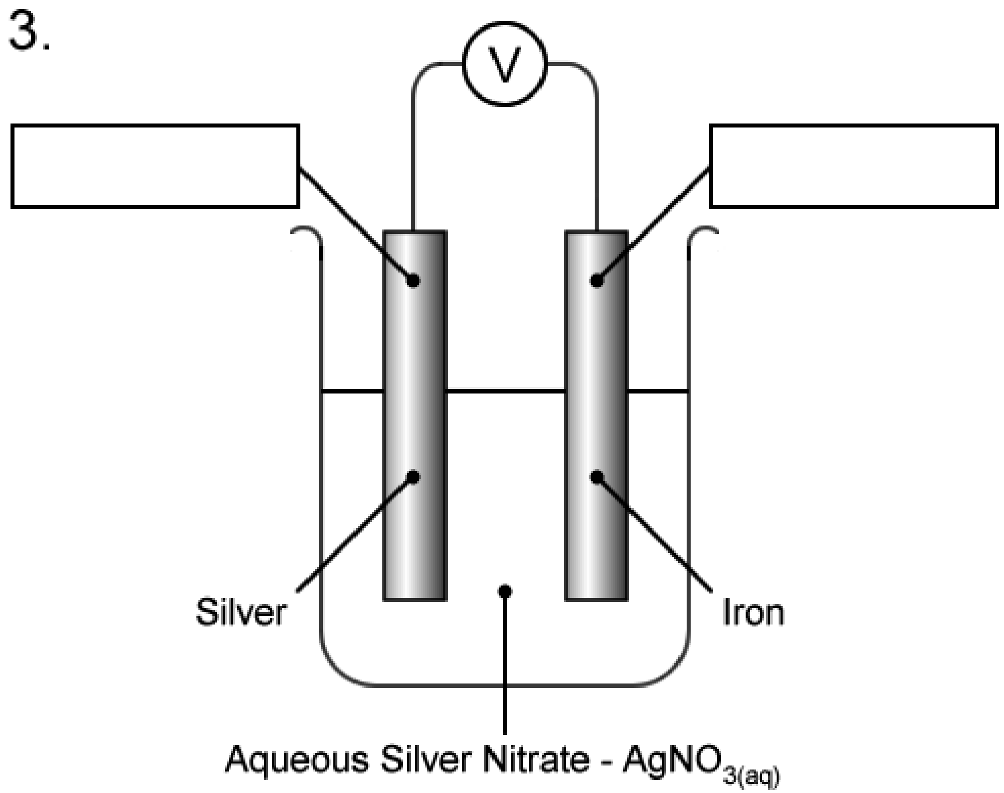
(Oxidation)

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

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

(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

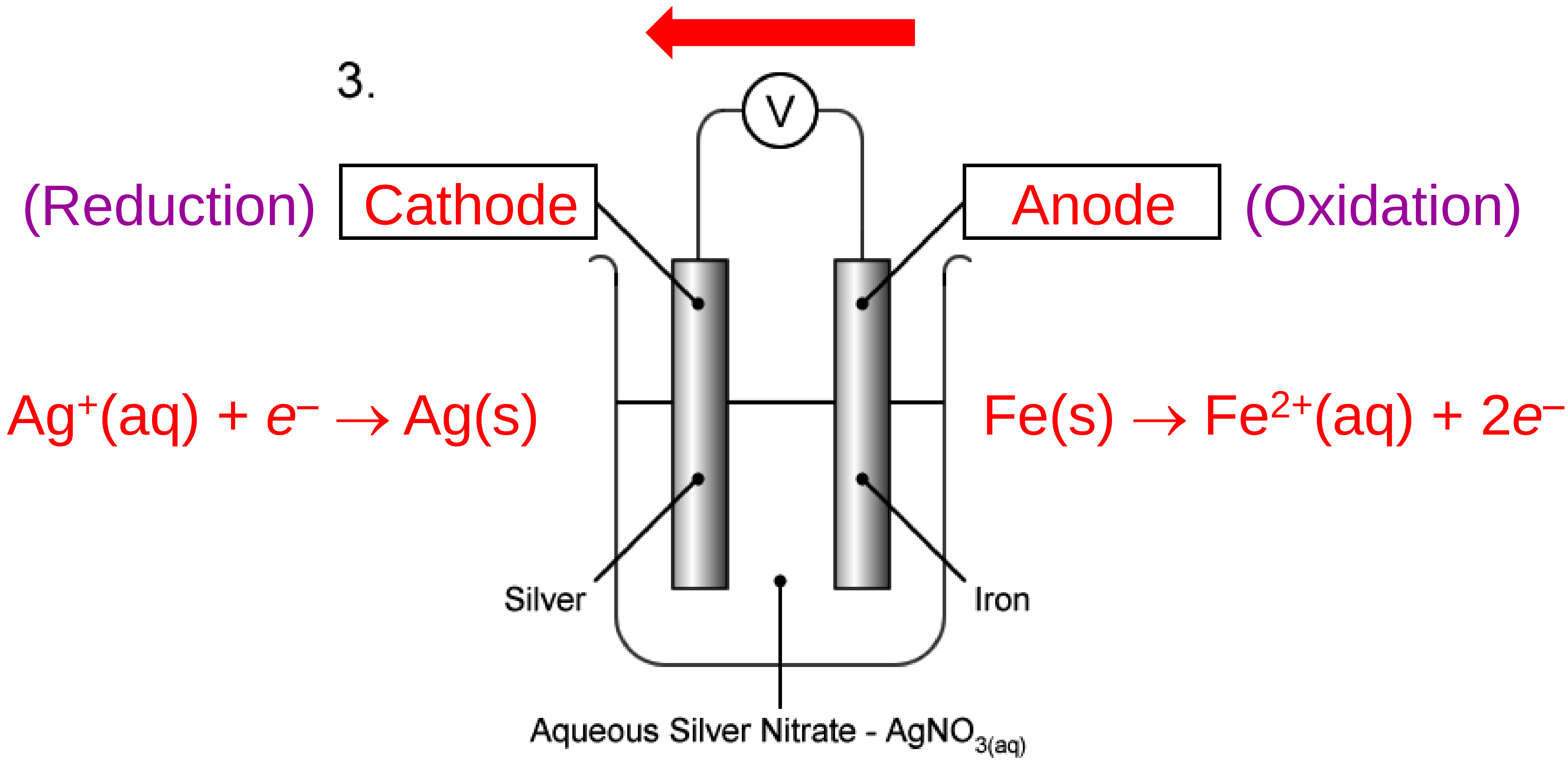
Electrochemistry
Electrochemical Cells



(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry
Electrochemical Cells

3.



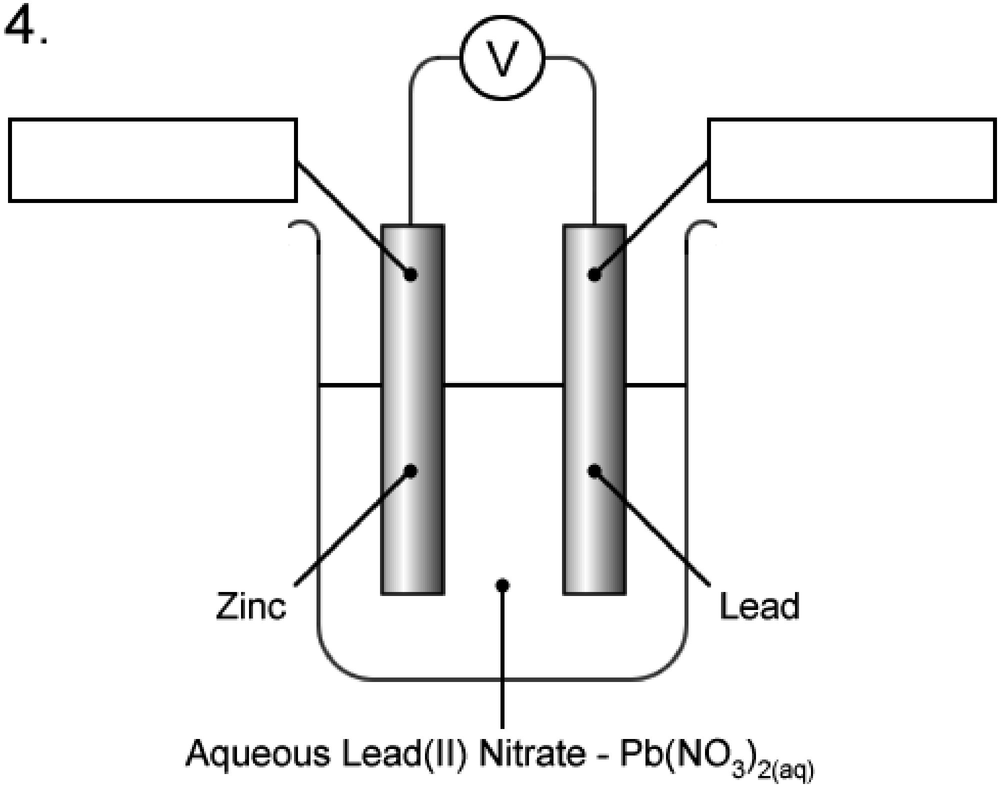
(a) Label the anode and the cathode.

(b) Write ionic half-equations for the reactions at the anode and cathode.

(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry
Electrochemical Cells

4.



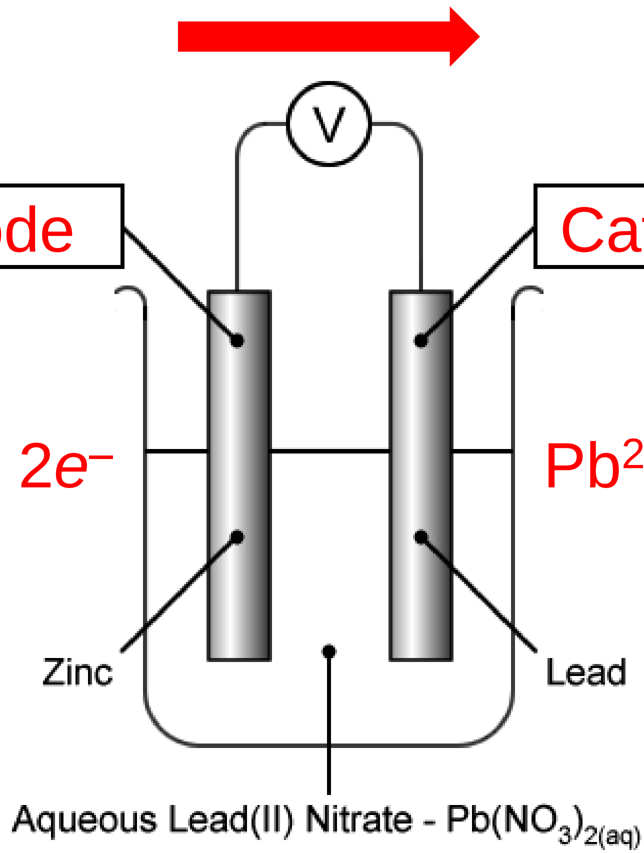
(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry
Electrochemical Cells

4.

(Oxidation) Anode

Cathode (Reduction)



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

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

(a) Label the anode and the cathode.
(b) Write ionic half-equations for the reactions at the anode and cathode.
(c) Draw an arrow to show the flow of electrons through the wires.

Electrochemistry

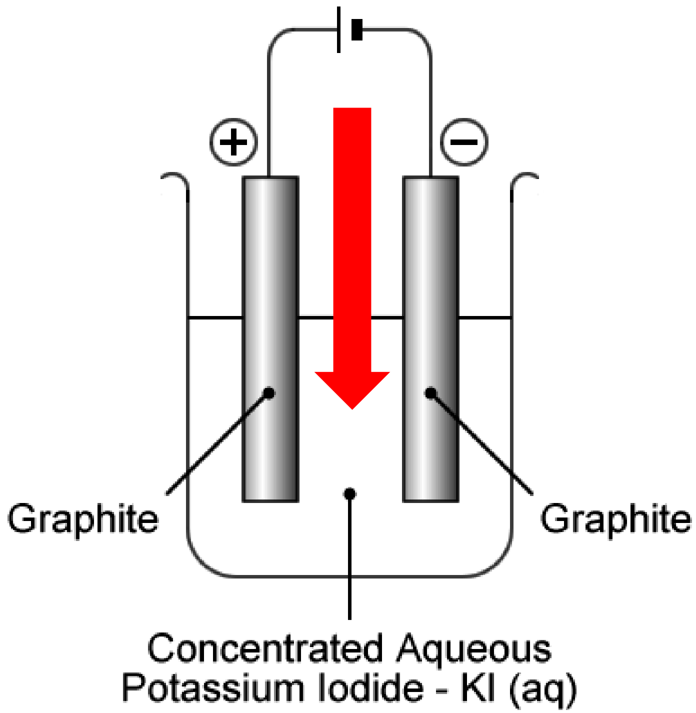
- Enduring understanding in a nutshell…

→ If a circuit diagram contains a battery, then the system is electrolysis. Electrical energy is used to decompose a chemical (endothermic).

→ If a circuit diagram does not contain a battery, but does contain two different metals connected to a voltmeter, then the system is a simple cell. The chemical system produces electricity (exothermic).

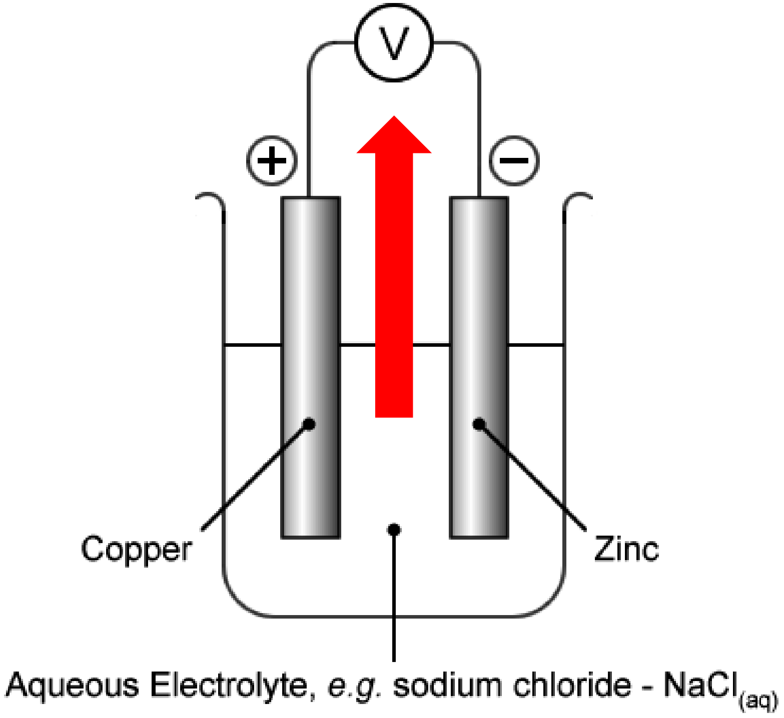
Electrochemistry

- Electrolysis



Electrical energy goes into the system (endothermic).
Electrical energy causes a chemical change.

- Simple Cell (Battery)



The system produces electrical energy (exothermic).
A chemical change produces electrical energy.

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

What are standard electrode potentials?

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

Electrode Reaction	Standard Electrode Potential ($E^\ominus$) / V
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.92
$Na^+(aq) + e^- \rightleftharpoons Na(s)$	–2.71
$Mg^{2+}(aq) + 2e^- \rightleftharpoons Mg(s)$	–2.38
$Zn^{2+}(aq) + 2e^- \rightleftharpoons Zn(s)$	–0.76
$H^+(aq) + e^- \rightleftharpoons {}^1/_2H_2(g)$	0.00
$Cu^{2+}(aq) + 2e^- \rightleftharpoons Cu(s)$	+0.34
$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

Electrode Reaction	Standard Electrode Potential ($E^{\ominus}$) / V
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.92
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$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

- The standard electrode potential is defined as the *potential* or tendency of a redox system to lose, or gain, electrons when compered to the *standard hydrogen electrode* (SHE) – which is assigned a value of 0.00 V.

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

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$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

- The $\rightleftharpoons$ symbol is used in place of the $\rightarrow$ symbol to show that the reaction is *reversible*.

- The ionic half-equations that are used to describe standard electrode potentials are always written showing the species on the left-hand-side being *reduced*, *i.e.* gaining electrons.

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

Electrode Reaction	Standard Electrode Potential ($E^{\ominus}$) / V
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$Ag^+(aq) + e^- \rightleftharpoons Ag(s)$	+0.80

- Redox systems with the reduced side (right-hand-side) *more reactive* than hydrogen have a *negative* electrode potential, *i.e.* they lose electrons *more readily* than hydrogen.

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

Electrode Reaction	Standard Electrode Potential ($E^\ominus$) / V
$K^+(aq) + e^- \rightleftharpoons K(s)$	–2.92
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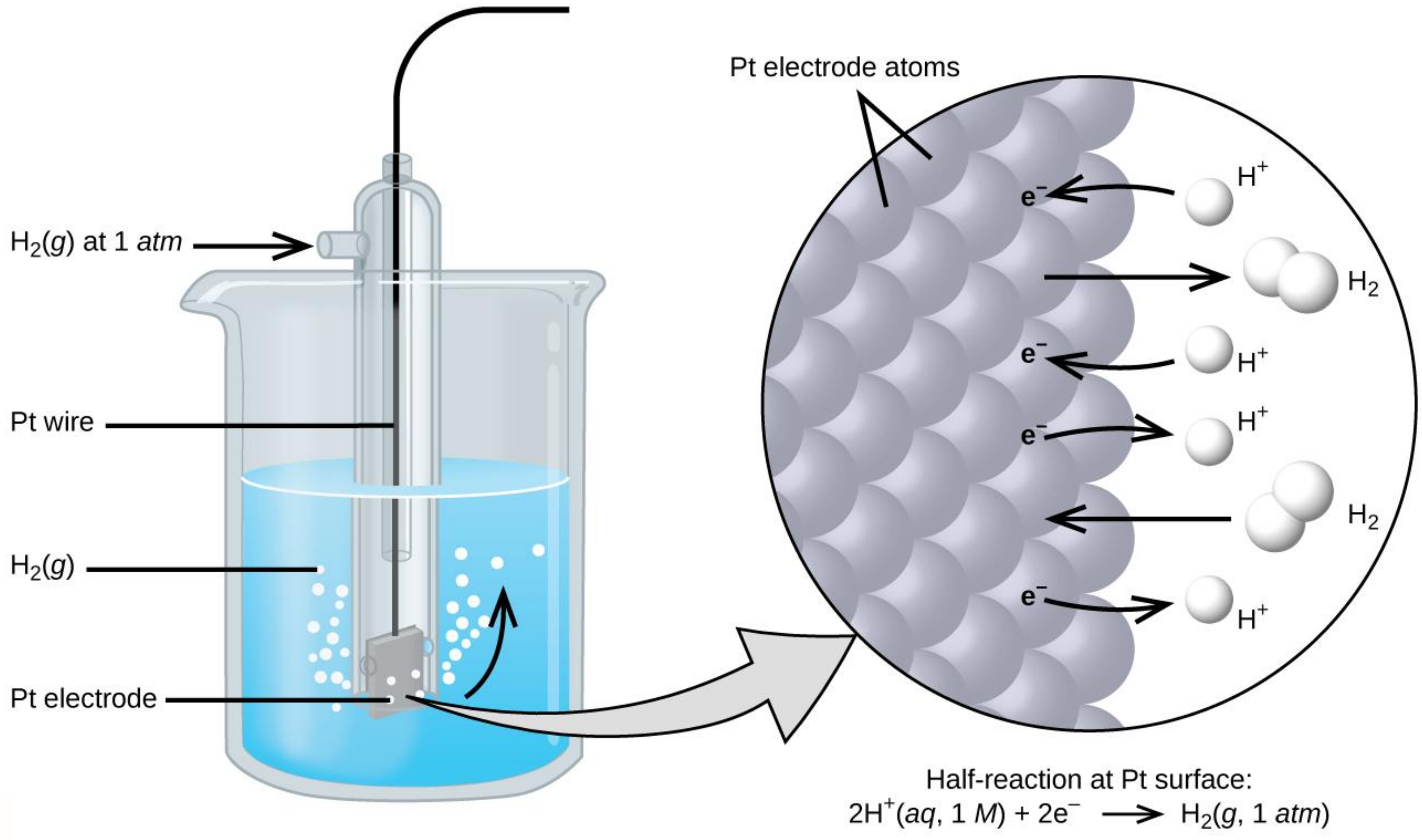
- Redox systems with the reduced side (right-hand-side) *less reactive* than hydrogen have a *positive* electrode potential, *i.e.* they lose electrons *less readily* than hydrogen.

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

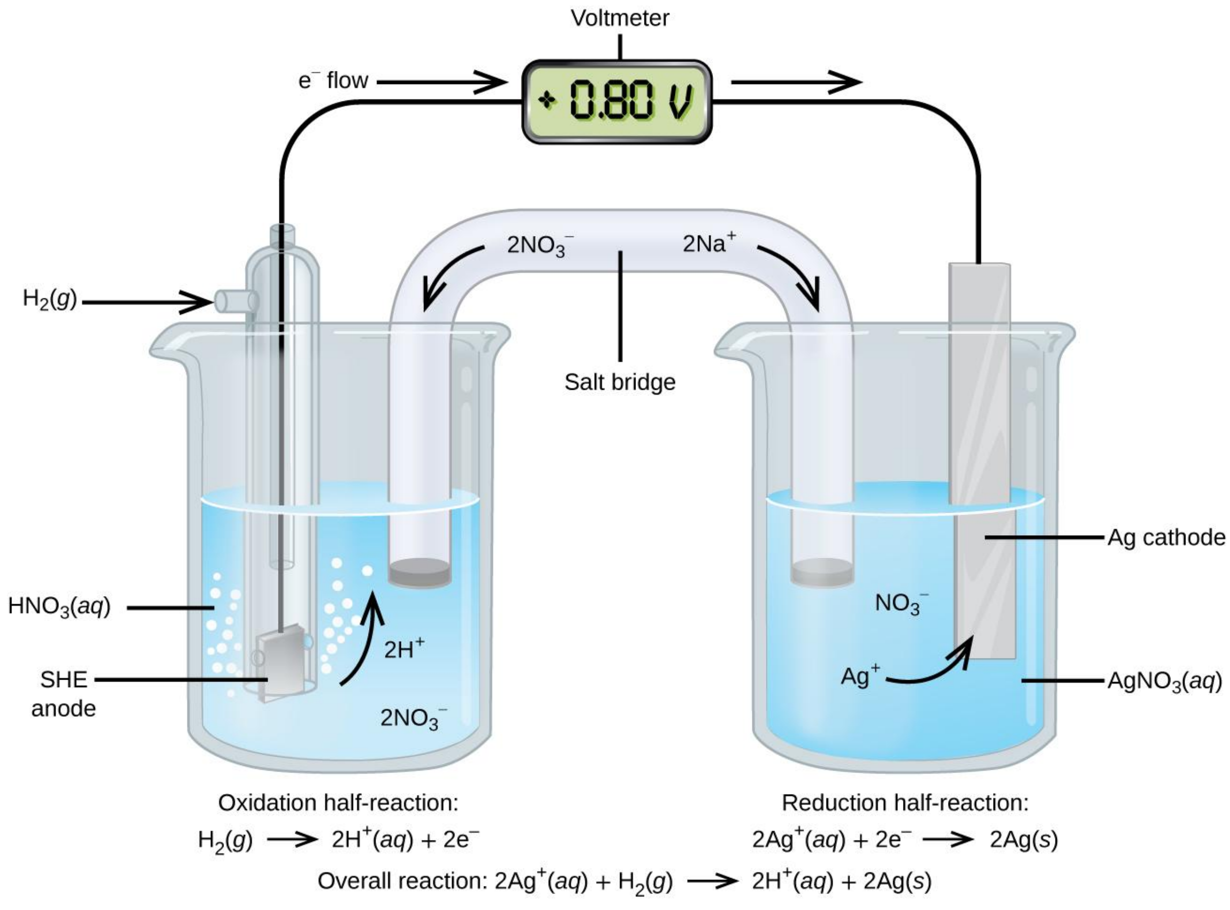
The Standard Hydrogen Electrode (SHE)



Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

Measuring the Potential of a Silver Half-Cell Relative to the Standard Hydrogen Electrode



Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

- Calculate the potential difference between a zinc half-cell and a copper half-cell:

→ Write the ionic half-equations for the two half-cells, the equation for the half-cell with the *less positive* (more negative) value being written *above* the equation for the half-cell with the *more positive* (less negative) value.

$$Zn^{2+}(aq) + 2e^{-} \rightleftharpoons Zn(s) \qquad E^{\ominus} = -0.76\ V$$

$$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s) \qquad E^{\ominus} = +0.34\ V$$

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

- Calculate the potential difference between a zinc half-cell and a copper half-cell:

→ Cycle through the two ionic-half equations in an *anti-clockwise* direction:

$$Zn^{2+}(aq) + 2e^{-} \rightleftharpoons Zn(s) \qquad E^{\ominus} = -0.76\ V$$

$$Cu^{2+}(aq) + 2e^{-} \rightleftharpoons Cu(s) \qquad E^{\ominus} = +0.34\ V$$

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

- Calculate the potential difference between a zinc half-cell and a copper half-cell:

→ The reaction for the *zinc* moves from *right-to-left* and the reaction for the *copper* moves from *left-to-right*.

$$Zn^{2+}(aq) + 2e^{-} \leftarrow Zn(s) \qquad E^{\ominus} = -0.76\ V$$

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s) \qquad E^{\ominus} = +0.34\ V$$

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

- Calculate the potential difference between a zinc half-cell and a copper half-cell:

→ Re-write the ionic half-equation for zinc so that the reaction takes place in the correct direction. Change the sign of $E^{\ominus}$ for zinc from *negative* to *positive* to reflect that the reaction is now taking place in the opposite direction.

$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-} \qquad E^{\ominus} = +0.76\ V$$

$$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s) \qquad E^{\ominus} = +0.34\ V$$

Electrochemistry

Electrochemical Cells – Standard Electrode Potentials

- Calculate the potential difference between a zinc half-cell and a copper half-cell:

→ Combine the ionic half-equations for the two half-cells together to produce the overall reaction that takes place in the cell.

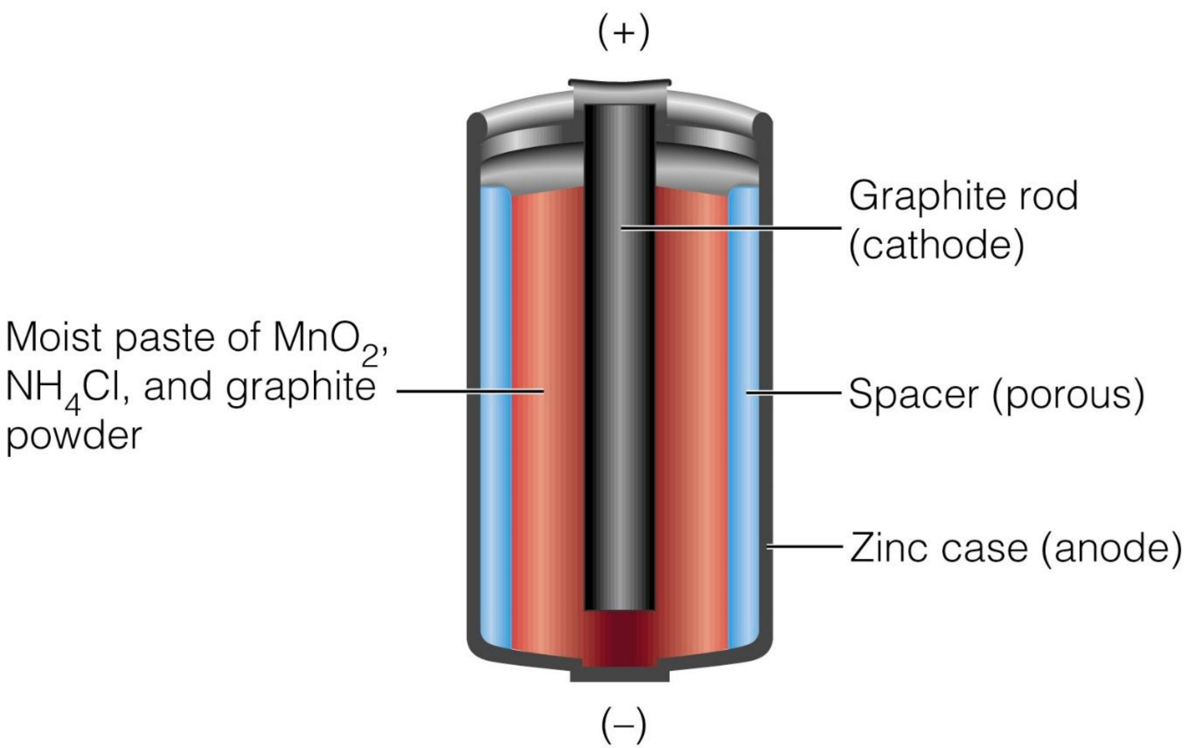
$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$$

→ The sum of the $E^{\ominus}$ values for the two half-cells gives the overall potential difference (voltage) produced by the cell:

$$(+0.76) + (+0.34) = 1.10 \text{ V}$$

Electrochemistry

Electrochemical Cells – Zinc Carbon Battery



Anode – Oxidation:

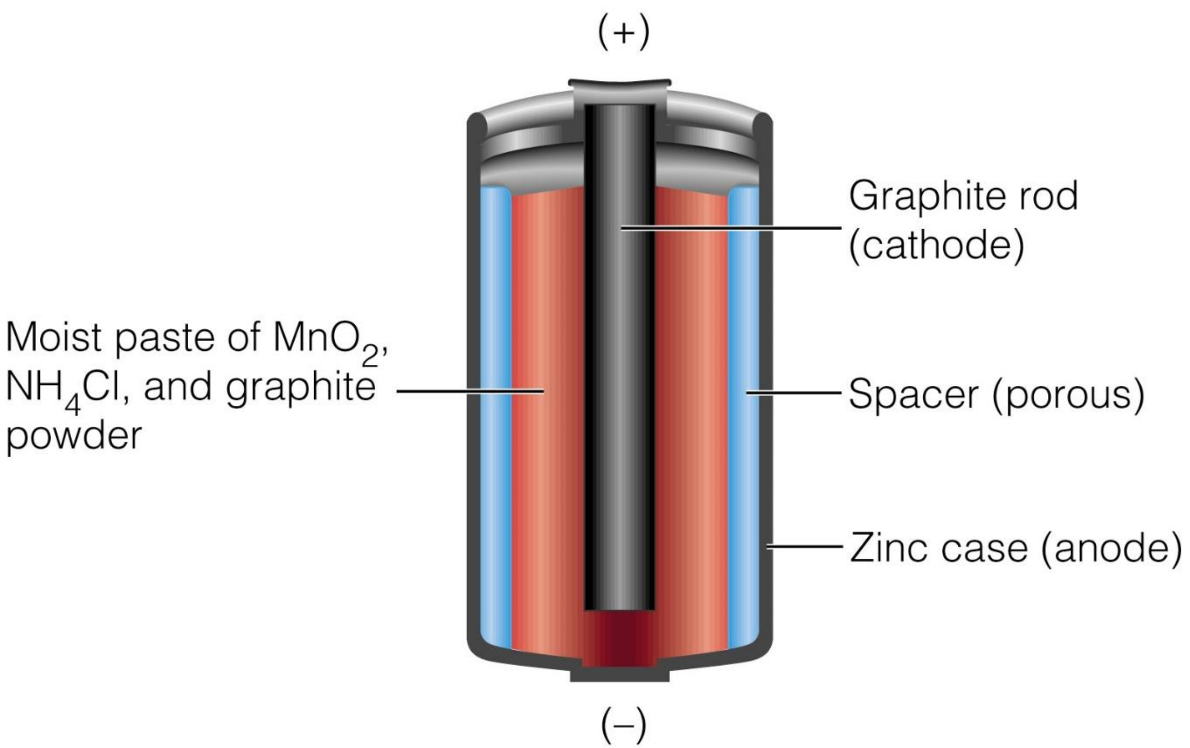
$$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$$

Cathode – Reduction:

$$2NH_4^{+}(aq) + 2e^{-} \rightarrow 2NH_3(g) + H_2(g)$$

Electrochemistry

Electrochemical Cells – Zinc Carbon Battery



Side Reactions:

$NH_3(g) + H_2O(l) \rightarrow NH_4OH(aq)$

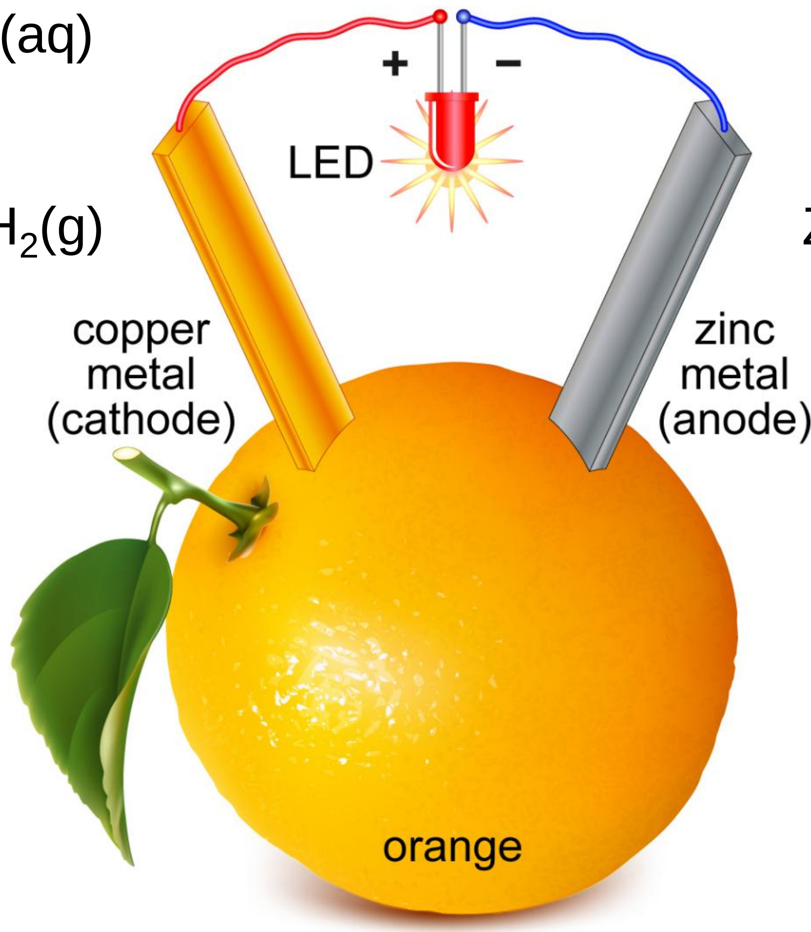
$MnO_2(s) + H_2(g) \rightarrow MnO(s) + H_2O(l)$

Electrochemistry

Electrochemical Cells – Citrus Fruit Battery

- Reduction of $H^+(aq)$ at the copper cathode:
  $2H^+(aq) + 2e^- \rightarrow H_2(g)$

- Oxidation of Zn(s) at the zinc anode:
  $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$



- The electrolyte is the orange juice which contains citric acid, a source of $H^+(aq)$

- Electrons flow through the external circuit from the zinc electrode to the copper electrode.

Electrochemistry

Presentation on

Electrochemical Cells (Batteries)

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8th February 2016