Electrochemistry

Chapter 20

Chemical vs Electro-chemical Reactions

- What is the difference between an "ordinary" Chemical Reaction and an Electro-chemical Reaction ?

"Ordinary" Reactions produce <u>Heat</u>

Electrochemical ones produce <u>Voltage</u>

Electrochemistry

20.1 Oxidation-Recuction (Redox) Reactions

20.1 Balancing Redox Reactions

20.3 Voltaic Cells (Batteries)

20.4 EMF

20.5 Spontaneity of Redox Reactions

20.9 Electrolysis

Two Parts to Electrochemistry

A CHEMICAL REACTION EITHER

1. **GENERATES Voltage**

   Therefore A BATTERY

   or

2. **USES Voltage**

   WHICH IS ELECTROLYSIS

Two Parts to Electrochemistry

1. **The Energy Released By A SPONTANEOUS Chemical Reaction**

Is Converted to Electricity

Or In Which

2. **Elecrical Energy is USED**

To CAUSE A NONSPONTANEOUS Chemical Reaction To Occur

Electrochemical Cells

- **Electrodes:** are where **Oxidation Reduction** reactions occur.
- **Anode:** is the electrode where oxidation takes place.
- **Cathode:** is the electrode where reduction takes place.

Review

Chapter 4

Page 128 - 129

Oxidation & Reduction Reactions

Oxidation Numbers

OXIDATION – REDUCTION (redox) REACTIONS

***electron*-TRANSFER REACTIONS**

***TWO STEP* PROCESS**

OR

TWO ***HALF-REACTIONS*** INVOLVED

1. **<u>OXIDATION</u>** – **<u>LOSS</u>** of electrons

2. <u>REDUCTION</u> – <u>GAIN</u> of electrons

- ***AN OXIDIZING AGENT "CAUSES" OXIDATION***

- ***A REDUCING AGENT "CAUSES" REDUCTION***

OXIDATION REDUCTION (*Redox*) REACTIONS

Examples

$$HgO(s) \rightarrow Hg(liq) + O_2\ (gas)$$

$$Fe(s) + O_2(gas) \rightarrow Fe_2O_3(s)$$

$$Zn(s) + HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

Rules for Assigning Oxidation Numbers

MUST KNOW OXIDATION "RULES"

See Page 128 – 129 Text

OXIDATION NUMBERS

1. **EACH ATOM in a PURE ELEMENT Has An OXIDATION NUMBER OF *ZERO***

2. **For IONS Consisting of a SINGLE ATOM, the Ox Num IS EQUAL TO THE CHARGE ON THE ION**

3. **F *ALWAYS* Has an Ox Num of –1 IN ALL OF ITS COMPOUNDS**

OXIDATION NUMBERS

4. **Cl, Br, and I are – 1 EXCEPT when combined with O or F**

5. **H is + 1 and O is – 2 EXCEPT** in

6. **Hydrides (CaH_2) & Peroxides (H_2O_2)**

7. **The ALGEBRAIC SUM of the Oxidation Numbers in a NEUTRAL Compound MUST be ZERO**

Prob 4.40 Determine the oxidation number for

1. Ti in TiO_2 **+4**
2. Sn in $SnCl_2$ **+2**
3. C in $C_2O_4^{2-}$ **+3**
4. N in $(NH_4)_2SO_4$ **-3**
5. N in HNO_3 **+5**
6. Cr in $Cr_2O_7^{2-}$ **+6**

Give OXIDATION NUMBER of Underlined Atom

1 $H\underline{Cl}O_4$	+7
2. $\underline{Cl}O_3^-$	+5
3. $\underline{Cl}F$	+1
4. $\underline{Cl}_2$	0

Redox Reactions

- Redox reaction are those involving the oxidation and reduction of species.
    - **OIL** – **O**xidation **I**s **L**oss of electrons.
    - **RIG** – **R**eduction **I**s **G**ain of electrons.
- Oxidation and reduction *must* occur together.

OXIDATION – REDUCTION (redox) REACTIONS

$$Zn(s) + HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

1. How do you know that this is an Oxidation Reduction (Redox) Reaction?

2. Does this reaction occur spontaneously?

Oxidation Reduction

- **What is The OXIDATION NUMBER For EACH Element Present ?**
- **What is OXIDIZED ?**
- **What is REDUCED ?**
- **What is The OXIDIZING AGENT ?**
- **What is The REDUCING AGENT ?**

$$Zn(s) + HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

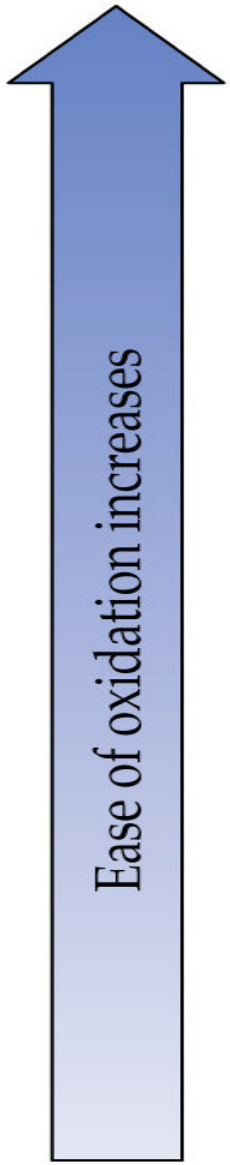
$Zn^{0} \rightarrow Zn^{2+}$ **OXIDATION**

$Cl^{-} \rightarrow Cl^{-}$ **NO CHANGE**

$H^{+} \rightarrow H_2^{O}$ **REDUCTION**

TABLE 4.5 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$Li(s) \longrightarrow Li^+(aq) + e^-$
Potassium	$K(s) \longrightarrow K^+(aq) + e^-$
Barium	$Ba(s) \longrightarrow Ba^{2+}(aq) + 2e^-$
Calcium	$Ca(s) \longrightarrow Ca^{2+}(aq) + 2e^-$
Sodium	$Na(s) \longrightarrow Na^+(aq) + e^-$
Magnesium	$Mg(s) \longrightarrow Mg^{2+}(aq) + 2e^-$
Aluminum	$Al(s) \longrightarrow Al^{3+}(aq) + 3e^-$
Manganese	$Mn(s) \longrightarrow Mn^{2+}(aq) + 2e^-$
Zinc	$Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$
Chromium	$Cr(s) \longrightarrow Cr^{3+}(aq) + 3e^-$
Iron	$Fe(s) \longrightarrow Fe^{2+}(aq) + 2e^-$
Cobalt	$Co(s) \longrightarrow Co^{2+}(aq) + 2e^-$
Nickel	$Ni(s) \longrightarrow Ni^{2+}(aq) + 2e^-$
Tin	$Sn(s) \longrightarrow Sn^{2+}(aq) + 2e^-$
Lead	$Pb(s) \longrightarrow Pb^{2+}(aq) + 2e^-$
Hydrogen	$H_2(g) \longrightarrow 2H^+(aq) + 2e^-$
Copper	$Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$
Silver	$Ag(s) \longrightarrow Ag^+(aq) + e^-$
Mercury	$Hg(l) \longrightarrow Hg^{2+}(aq) + 2e^-$
Platinum	$Pt(s) \longrightarrow Pt^{2+}(aq) + 2e^-$
Gold	$Au(s) \longrightarrow Au^{3+}(aq) + 3e^-$

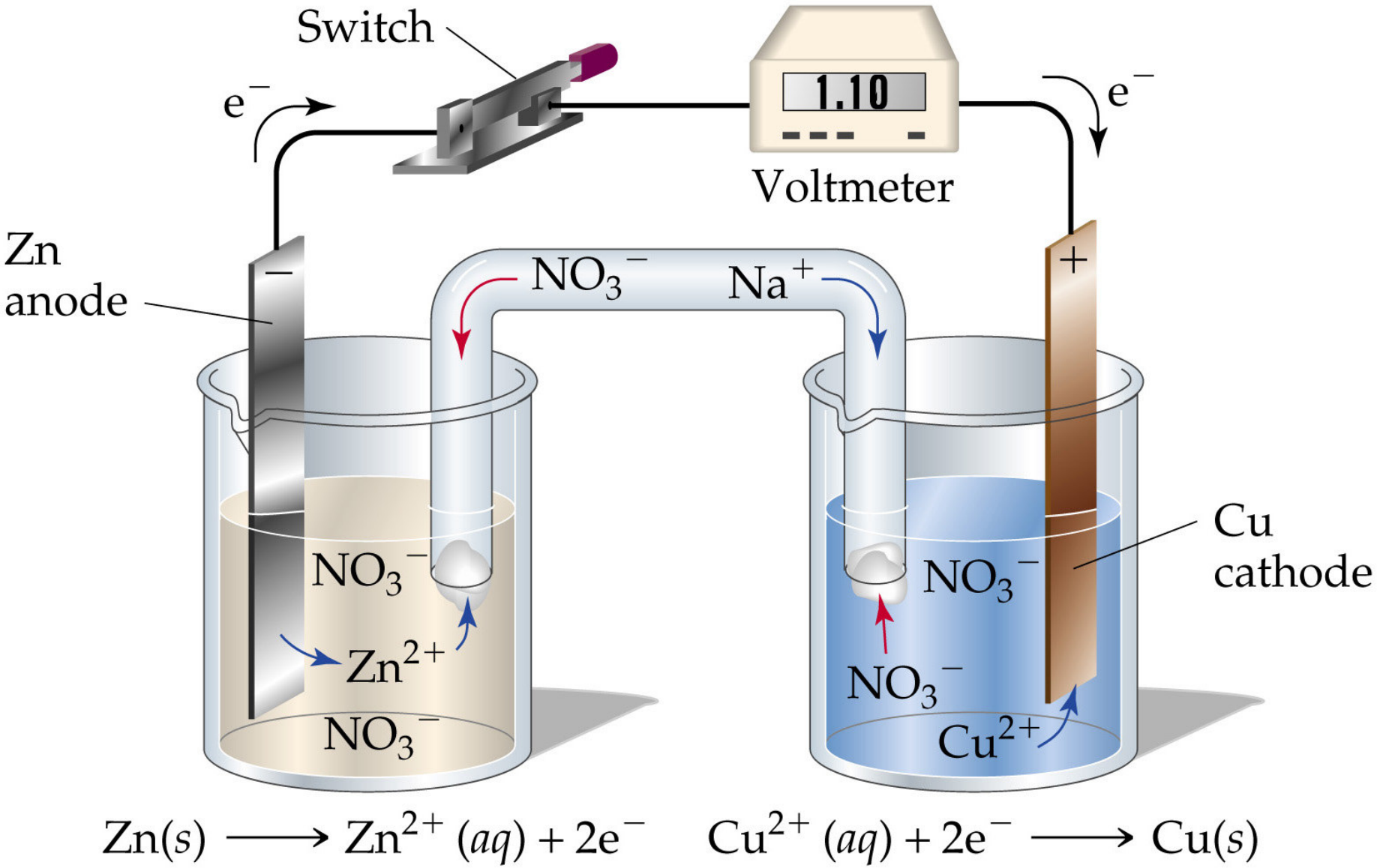


Back to Electrochemistry

Electrochemical Cells

Construct An Electrochemical Cell

1. Two Beakers
2. One containing a solution of Zinc Nitrate with a strip of Zinc Metal in it
3. The other beaker containing a solution of Copper II Nitrate with a strip of Cu Metal
4. A liquid connection between beakers


Switch
e⁻
1.10
Voltmeter
e⁻
Zn
anode
−
+
NO3⁻
Na⁺
NO3⁻
Zn2+
NO3⁻
NO3⁻
NO3⁻
Cu2+
Cu
cathode
Zn(s) ⟶ Zn2+ (aq) + 2e⁻
Cu2+ (aq) + 2e⁻ ⟶ Cu(s)
Movement of cations
Movement of anions

Electrochemical Cells

Electrodes: are usually metals strips/wires

Salt Bridge: is a U–shaped tube that contains a salt. Used to connect solutions

Anode: electrode where oxidation takes place

Cathode: electrode where reduction takes place

Electrochemical Cells

Electrode

OXIDATION REDUCTION REACTIONS

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

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

Net Cell Reaction

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

Electrochemical cell Convention

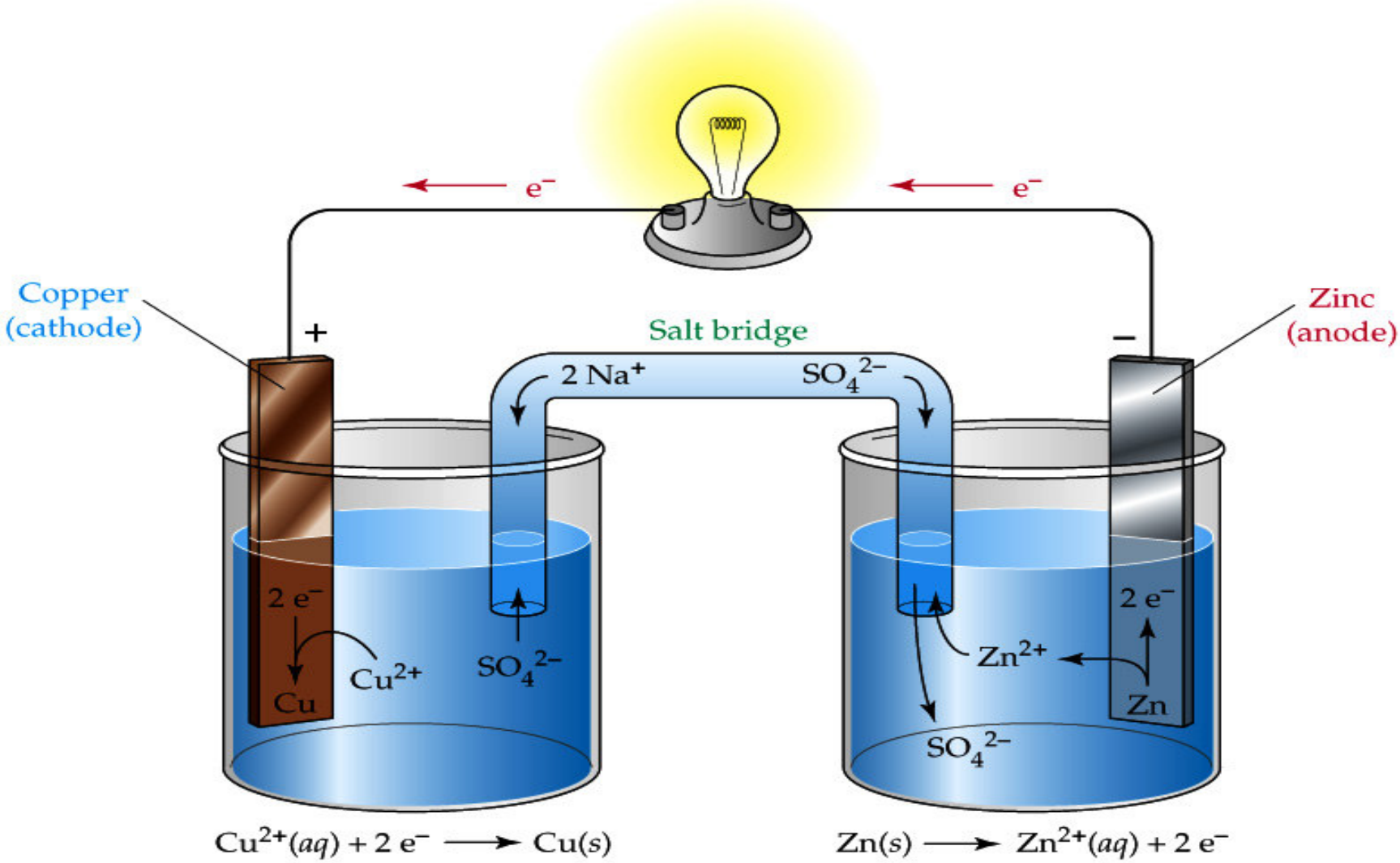
$Zn(s) \mid Zn^{2+}(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$

Anode is placed on **left** by convention

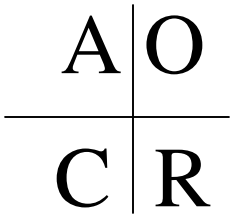
$\mid$ Used to indicate a change of phase

$\parallel$ Used to indicate a salt bridge

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



Electrochemical cell



Electrons “flow” from anode to cathode

Therefore Anode is Negative Electrode

Anions “flow” from cathode to anode

Cations “flow” from anode to cathode

Electrochemical Cells

Given the following cell notation,

$Pt(s) \mid Sn^{2+}(aq), Sn^{4+}(aq) \parallel Ag^{+}(aq) \mid Ag(s)$

write the anode and cathode reactions

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2\ e^{-}$$

$$2\ e^{-} + Ag^{1+}(aq) \rightarrow Ag(s)$$

Write oxidation reduction reactions for

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

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

Cathode $2e + Sn^{2+}(aq) \rightarrow Sn(s)$ reduction

Write the cell shorthand notation

$Fe(s) \mid Fe^{2+}(aq) \parallel Sn^{2+}(aq) \mid Sn(s)$

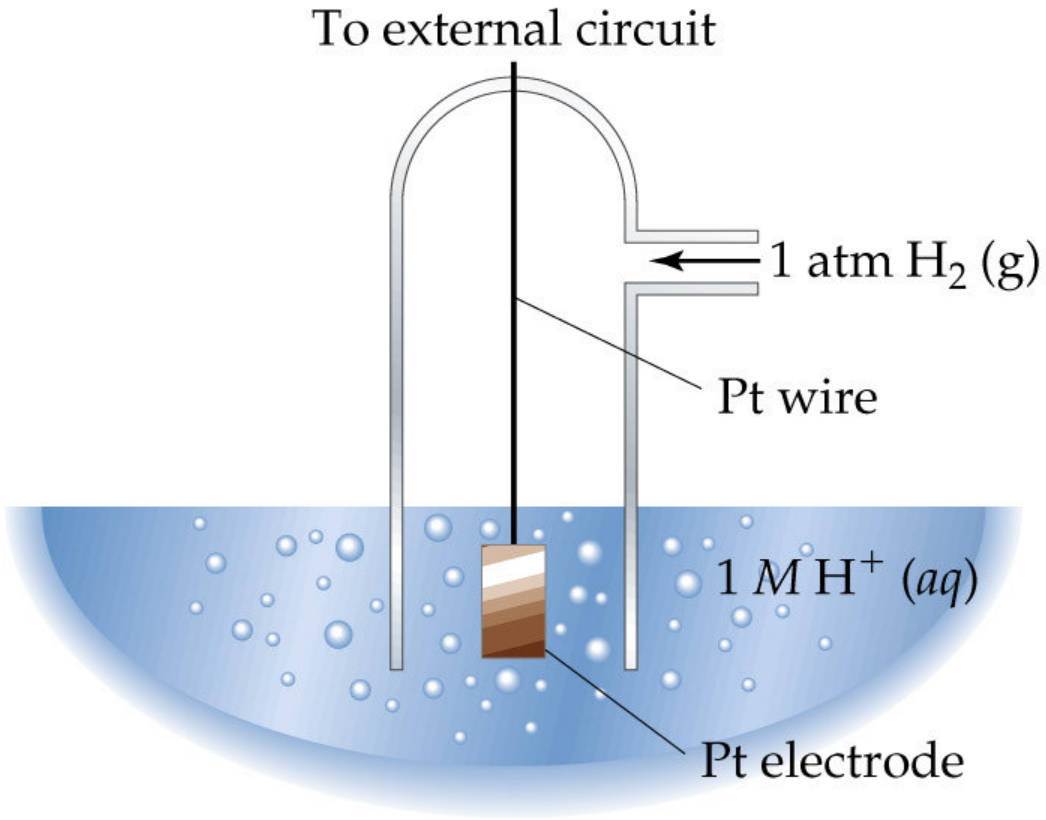
Anode **saltbridge** Cathode

$$E^\circ_{Cell} = E^\circ_{oxidation} + E^\circ_{reduction}$$

The standard potential of any galvanic cell is the **sum** of the standard half–cell potentials for the oxidation and reduction half–cells.

Standard half–cell potentials are always quoted as a reduction process. The sign must be changed for the oxidation process.

Standard Hydrogen Electrode (SHE)



standard half–cell potentials

- The reference point is called the *standard hydrogen electrode (S.H.E.)* and consists of a platinum electrode in contact with H_2 gas (1 atm) and aqueous H^+ ions (1M).
- The *standard hydrogen electrode* is assigned an arbitrary value of exactly 0.00V.

TABLE 20.1 Standard Reduction Potentials in Water at 25°C

Potential (V)	Reduction Half-Reaction
+2.87	$F_2(g) + 2e^- \longrightarrow 2F^-(aq)$
+1.51	$MnO_4^-(aq) + 8H^+(aq) + 5e^- \longrightarrow Mn^{2+}(aq) + 4H_2O(l)$
+1.36	$Cl_2(g) + 2e^- \longrightarrow 2Cl^-(aq)$
+1.33	$Cr_2O_7^{2-}(aq) + 14H^+(aq) + 6e^- \longrightarrow 2Cr^{3+}(aq) + 7H_2O(l)$
+1.23	$O_2(g) + 4H^+(aq) + 4e^- \longrightarrow 2H_2O(l)$
+1.06	$Br_2(l) + 2e^- \longrightarrow 2Br^-(aq)$
+0.96	$NO_3^-(aq) + 4H^+(aq) + 3e^- \longrightarrow NO(g) + 2H_2O(l)$
+0.80	$Ag^+(aq) + e^- \longrightarrow Ag(s)$
+0.77	$Fe^{3+}(aq) + e^- \longrightarrow Fe^{2+}(aq)$
+0.68	$O_2(g) + 2H^+(aq) + 2e^- \longrightarrow H_2O_2(aq)$
+0.59	$MnO_4^-(aq) + 2H_2O(l) + 3e^- \longrightarrow MnO_2(s) + 4OH^-(aq)$
+0.54	$I_2(s) + 2e^- \longrightarrow 2I^-(aq)$
+0.40	$O_2(g) + 2H_2O(l) + 4e^- \longrightarrow 4OH^-(aq)$
+0.34	$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$
0	$2H^+(aq) + 2e^- \longrightarrow H_2(g)$
−0.28	$Ni^{2+}(aq) + 2e^- \longrightarrow Ni(s)$
−0.44	$Fe^{2+}(aq) + 2e^- \longrightarrow Fe(s)$
−0.76	$Zn^{2+}(aq) + 2e^- \longrightarrow Zn(s)$
−0.83	$2H_2O(l) + 2e^- \longrightarrow H_2(g) + 2OH^-(aq)$
−1.66	$Al^{3+}(aq) + 3e^- \longrightarrow Al(s)$
−2.71	$Na^+(aq) + e^- \longrightarrow Na(s)$
−3.05	$Li^+(aq) + e^- \longrightarrow Li(s)$

Standard Hydrogen Electrode (SHE)

The SHE consists of a Pt electrode in a tube placed in $H^+(aq)$. H_2 is bubbled over electrode

$$E^\circ = \text{ zero} \quad \text{by definition}$$

$$2H^+(aq) + 2e^- \rightarrow H_2(gas) \qquad E^\circ = \text{ zero}$$

$$H_2(gas) \rightarrow 2H^+(aq) + 2e^- \qquad E^\circ = \text{ zero}$$

reduction potentials

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$, $E°_{red} = -0.76$ V.

Changing the stoichiometric coefficient does not affect $E°$

$2\ Zn^{2+}(aq) + 4e^- \rightarrow 2\ Zn(s)$, $E°_{red} = -0.76$ V

Reactions with $E < 0$ are nonspontaneous

oxidation potentials

Only the sign is changed !

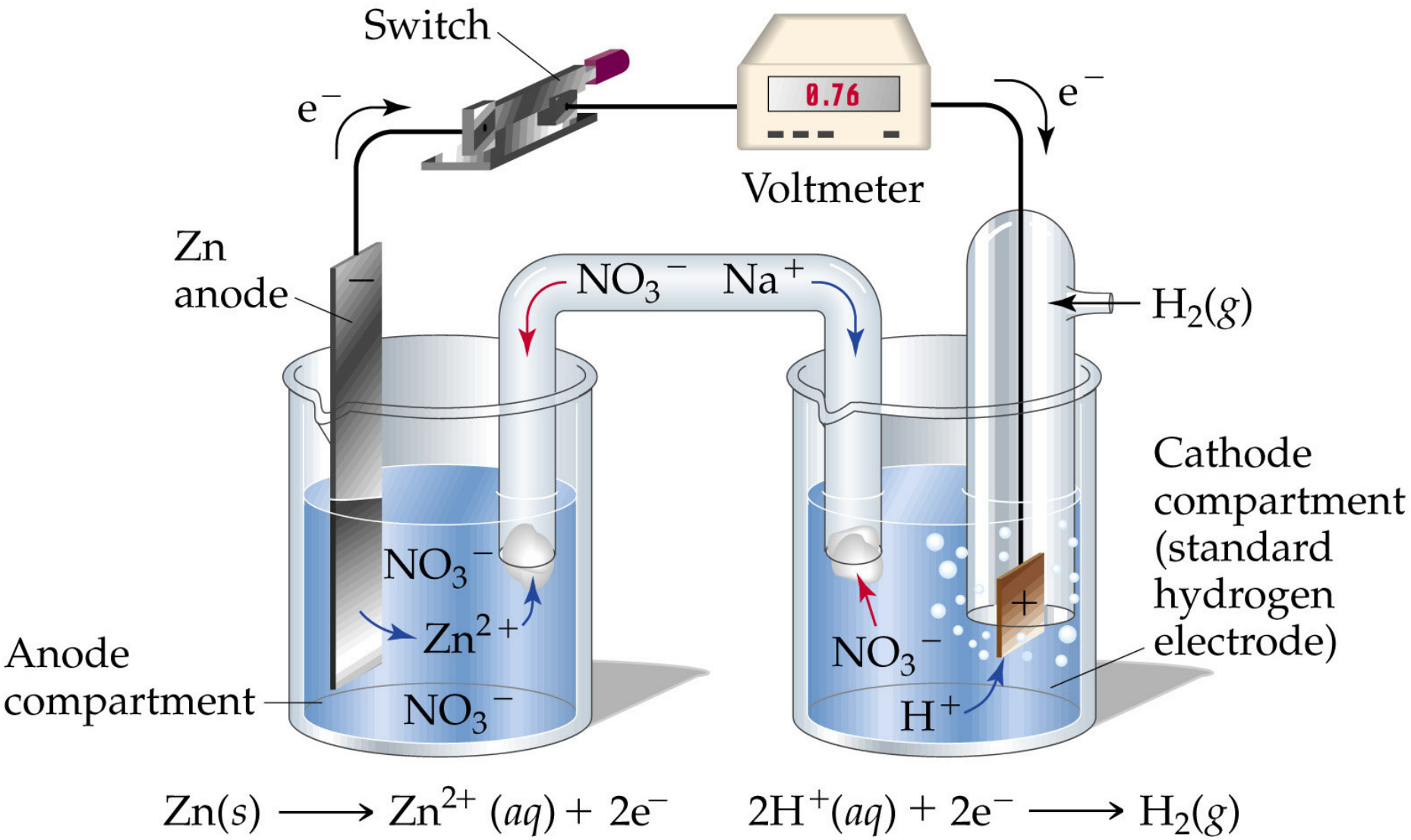
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-}$ $E°_{red} = +0.76$ V.

Changing the stoichiometric coefficient

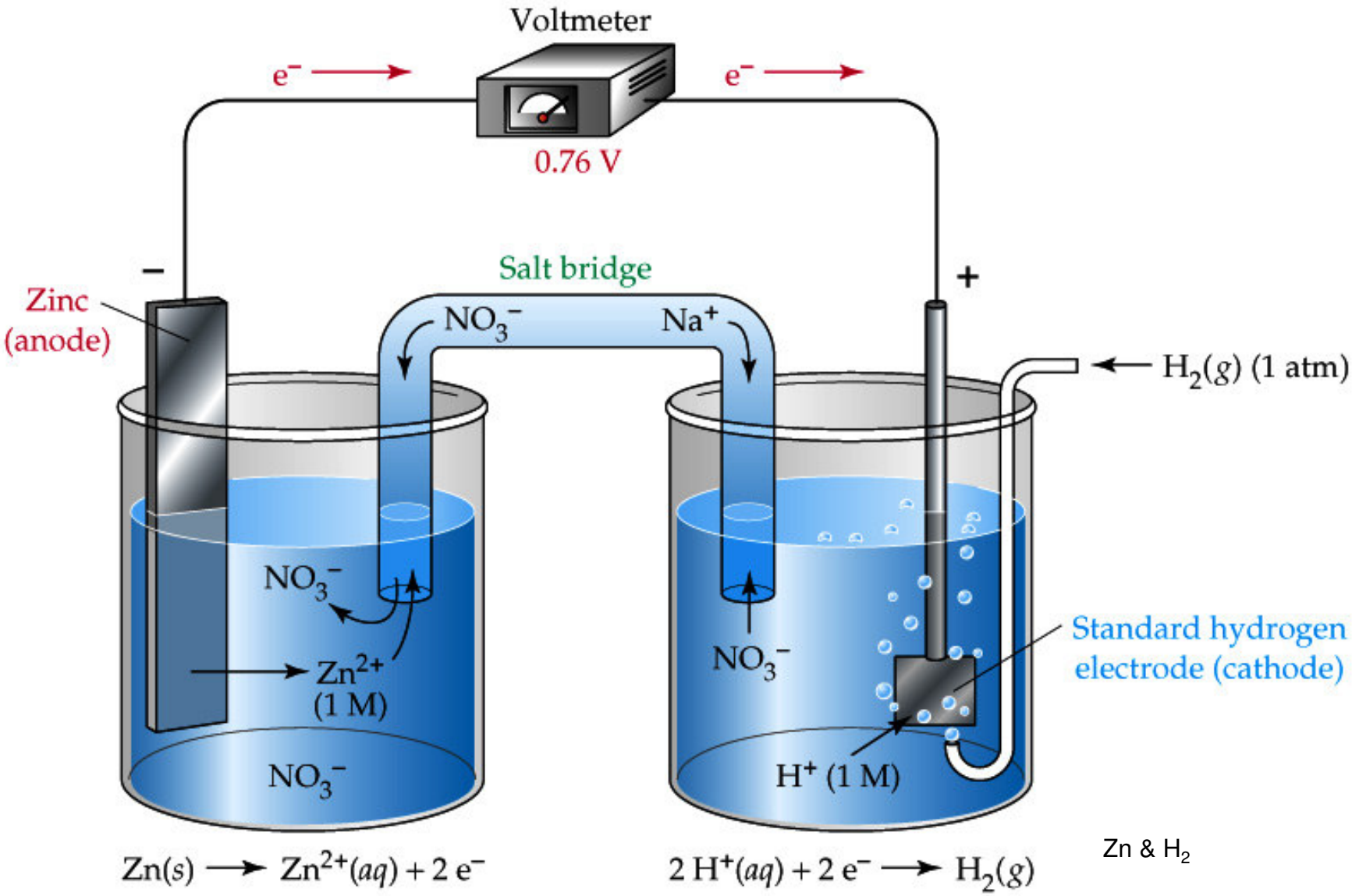
does not affect $E°$

$2Zn(s) \rightarrow 2Zn^{2+}(aq) + 4e^{-}$ $E°_{red} = +0.76$ V.

Reactions with $E > 0$ are spontaneous


Switch
e⁻
0.76
e⁻
Voltmeter
Zn
anode
−
NO3⁻ Na⁺
H2(g)
Cathode
compartment
(standard
hydrogen
electrode)
NO3⁻
Zn2+
Anode
compartment
NO3⁻
+
NO3⁻
H+
Zn(s) ⟶ Zn2+ (aq) + 2e⁻
2H+(aq) + 2e⁻ ⟶ H2(g)

Cell Voltage



Zn & H_2

$$Zn(s) \mid Zn^{2+}(aq) \parallel H^{+}(aq) \mid H_2(gas)$$

E_{cell} measured relative to the SHE

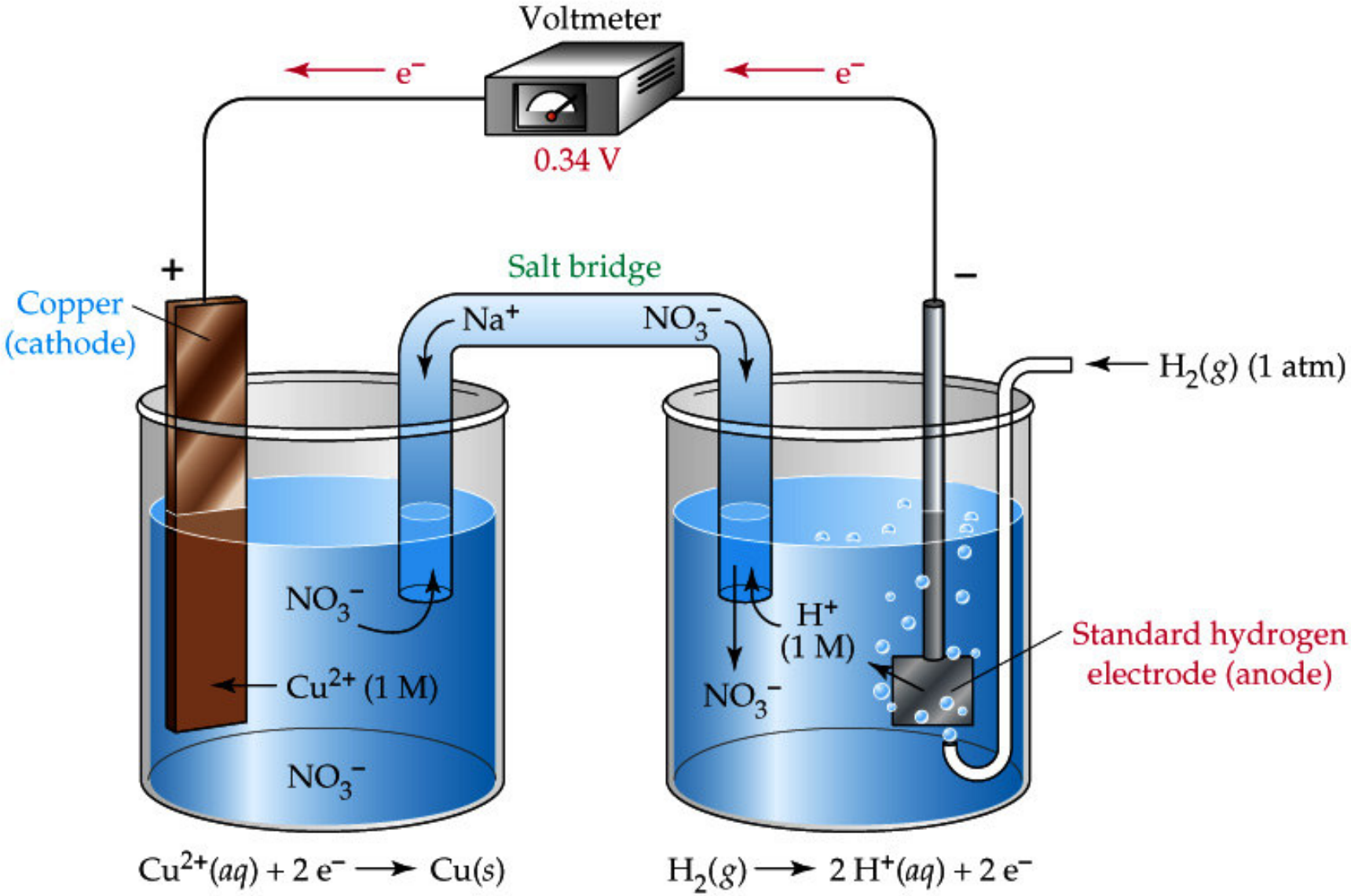
$$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode}) = +0.76$$

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

$$\underline{2H^{+}(aq) + 2e^- \rightarrow H_2(gas)} \qquad E° = +0.00\ V$$

$$Zn(s) + HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

For Copper


Voltmeter
e⁻
e⁻
0.34 V
+
Salt bridge
−
Copper
(cathode)
Na⁺
NO₃⁻
H₂(g) (1 atm)
NO₃⁻
Cu²⁺ (1 M)
NO₃⁻
H⁺
(1 M)
NO₃⁻
Standard hydrogen
electrode (anode)
Cu²⁺(aq) + 2 e⁻ → Cu(s)
H₂(g) → 2 H⁺(aq) + 2 e⁻

$$H_2(gas) \mid H^+(aq) \parallel Cu^{2+}(aq) \mid Cu(s)$$

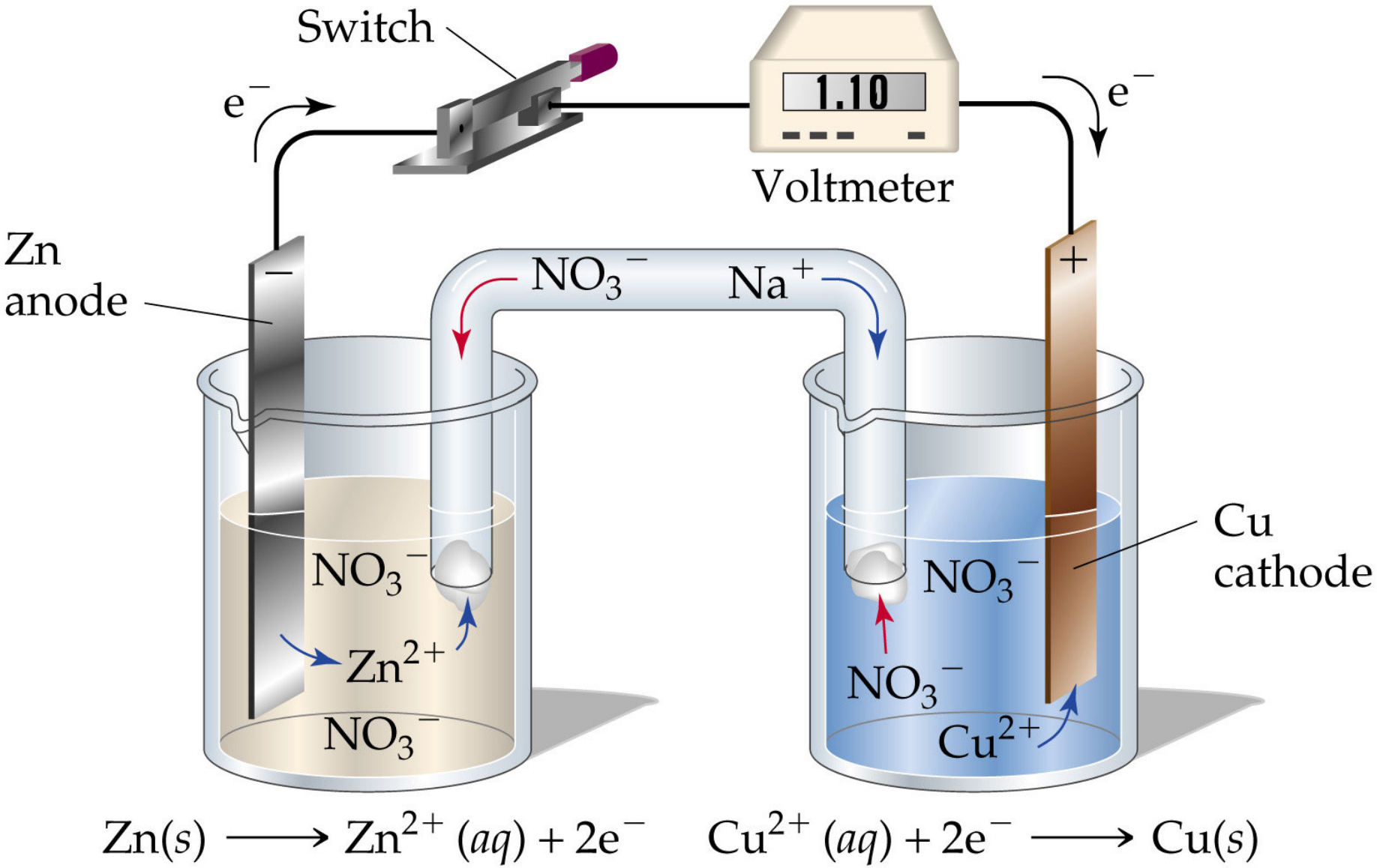
E_{cell} measured relative to the SHE

$$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode}) = +0.34$$

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

$$\underline{H_2(gas) \rightarrow 2H^+(aq) + 2e^-} \qquad E° = +0.00 \text{ V}$$

$$H_2(g) + CuCl_2(aq) \rightarrow Cu(s) + HCl(aq)$$


Switch
e⁻
1.10
Voltmeter
e⁻
Zn
anode
−
+
NO3⁻
Na⁺
NO3⁻
Zn2+
NO3⁻
NO3⁻
NO3⁻
Cu2+
Cu
cathode
Zn(s) ⟶ Zn2+ (aq) + 2e⁻
Cu2+ (aq) + 2e⁻ ⟶ Cu(s)
Movement of cations
Movement of anions

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

$$E°_{cell} \text{ measured} = +1.10 \text{ V}$$

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

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

$$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s) \qquad E°_{cell} = 1.10$$

Reactions with $E > 0$ are spontaneous

Reactions with $E < 0$ are nonspontaneous

When selecting two half cell reactions the more negative value will form the oxidation half–cell.

Consider the reaction between zinc and silver:

$$Ag^{+}(aq) + e^{-} \rightarrow Ag(s) \qquad E^{\circ} = 0.80V$$

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

- Therefore zinc forms the oxidation half–cell:

$$\mathbf{Zn(s) \rightarrow Zn^{2+}(aq) + 2e^{-} \quad E^{\circ} = -(-0.76V)}$$

Electrochemical Cells

Can $Fe^{2+}(aq)$ oxidize Al(s) ?

(a) yes (b) no (c) I do not know

Write Reaction

$Fe^{2+}(aq) + Al(s) \rightarrow Fe(s) + Al^{+3}(aq)$

$2e + Fe^{2+}(aq) \rightarrow Fe(s) \quad E_{red} = -0.44$

$Al(s) \rightarrow Al^{+3}(aq) + 3e \quad E_{oxid} = +1.66$

$E_{cell} = E_{oxid} + E_{red} = +$ therefore YES!

Electrochemical Cells

- Can $Pb^{2+}(aq)$ oxidize $Cu(s)$?
- Calculate E°_{cell} formed with lead and copper

Batteries

They produce **POSITIVE** voltage

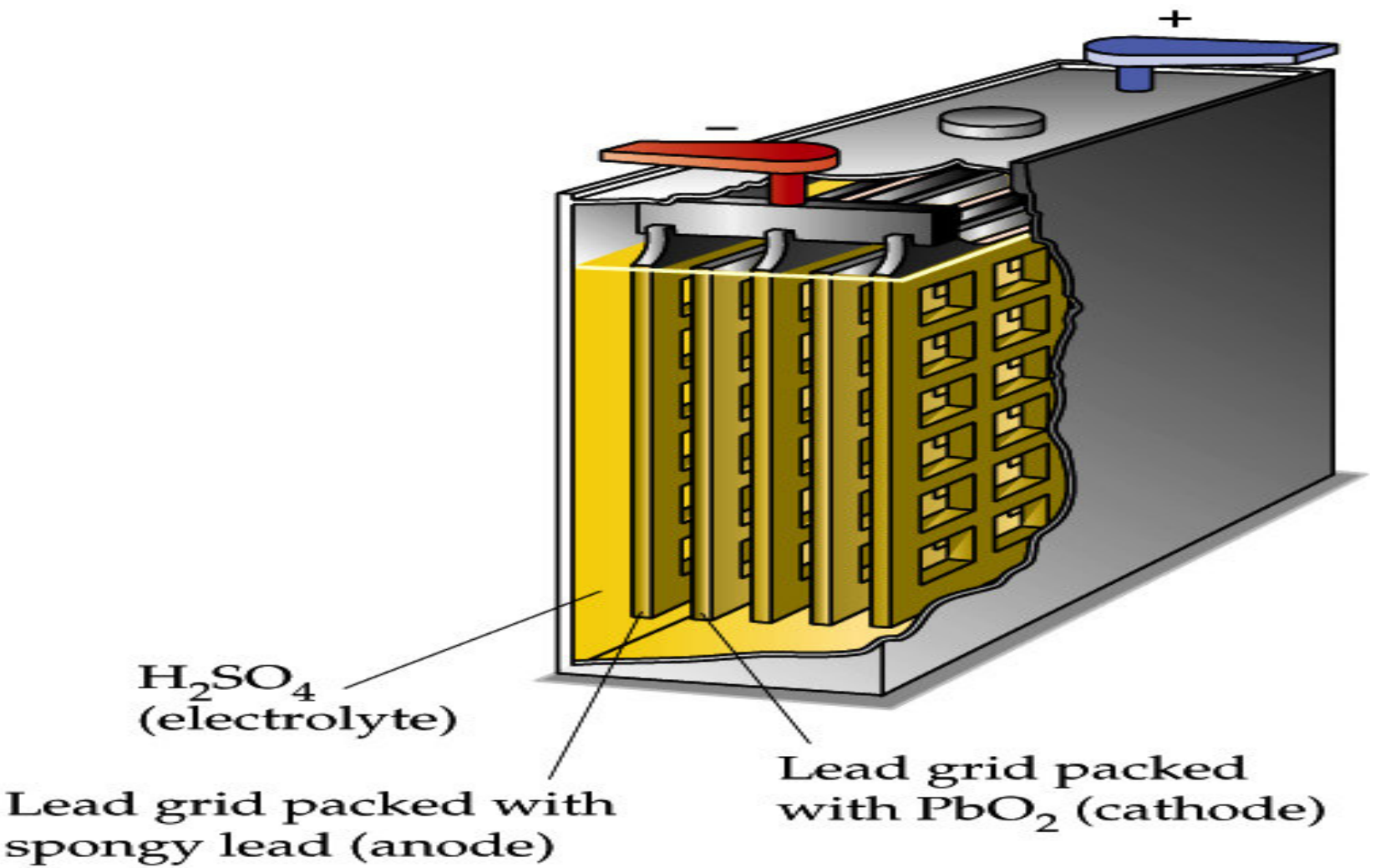
Batteries 101

Batteries are the most important practical application of galvanic cells.

Single cell batteries consist of one galvanic cell.

Multicell batteries consist of several galvanic cells linked in series to obtain the desired voltage

Lead Storage Battery:



Lead-Acid Battery

The overall electrochemical reaction is

$$PbO_2(s) + Pb(s) + 2SO_4^{2-}(aq) + 4H^+(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

for which

$$E°_{cell} = E°_{red}(\text{cathode}) - E°_{red}(\text{anode})$$

$$= (+1.685\ \text{V}) - (-0.356\ \text{V})$$

$$= +2.041\ \text{V}.$$

Wood or glass-fiber spacers are used to prevent the electrodes form touching

Lead Storage Battery:

A typical “12 volt” battery consists of six individual cells connected in series.

Cell Potential: 1.924V {each cell

Electrolyte: 38% by mass Sulfuric Acid.

Anode: Lead grid packed with spongy lead.

Cathode: Lead grid packed with lead IV oxide.

Lead Storage Battery:

Anode: lead

$$Pb(s) + HSO_4^-(aq) \rightarrow PbSO_4(s) + H^+(aq) + 2e^-$$

Cathode: lead IV oxide

$$PbO_2(s) + 3H^+(aq) + HSO_4^-(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O(l)$$

Lead Storage Battery:

- Pb is oxidized to $PbSO_4$ at the anode, and
- PbO_2 is reduced to $PbSO_4$ at the cathode
- $PbSO_4$ adheres to the electrodes causing the battery to "run–down."
- Recharging the battery reverses this reaction to reform Pb and PbO_2.

Fuel Cell:

Uses externally fed CH_4 or H_2, which react to form water. Most common is H_2.

Anode: Porous carbon containing metallic catalysts

$$2H_2(s) + 4OH^-(aq) \rightarrow 4H_2O(l) + 4e^-$$

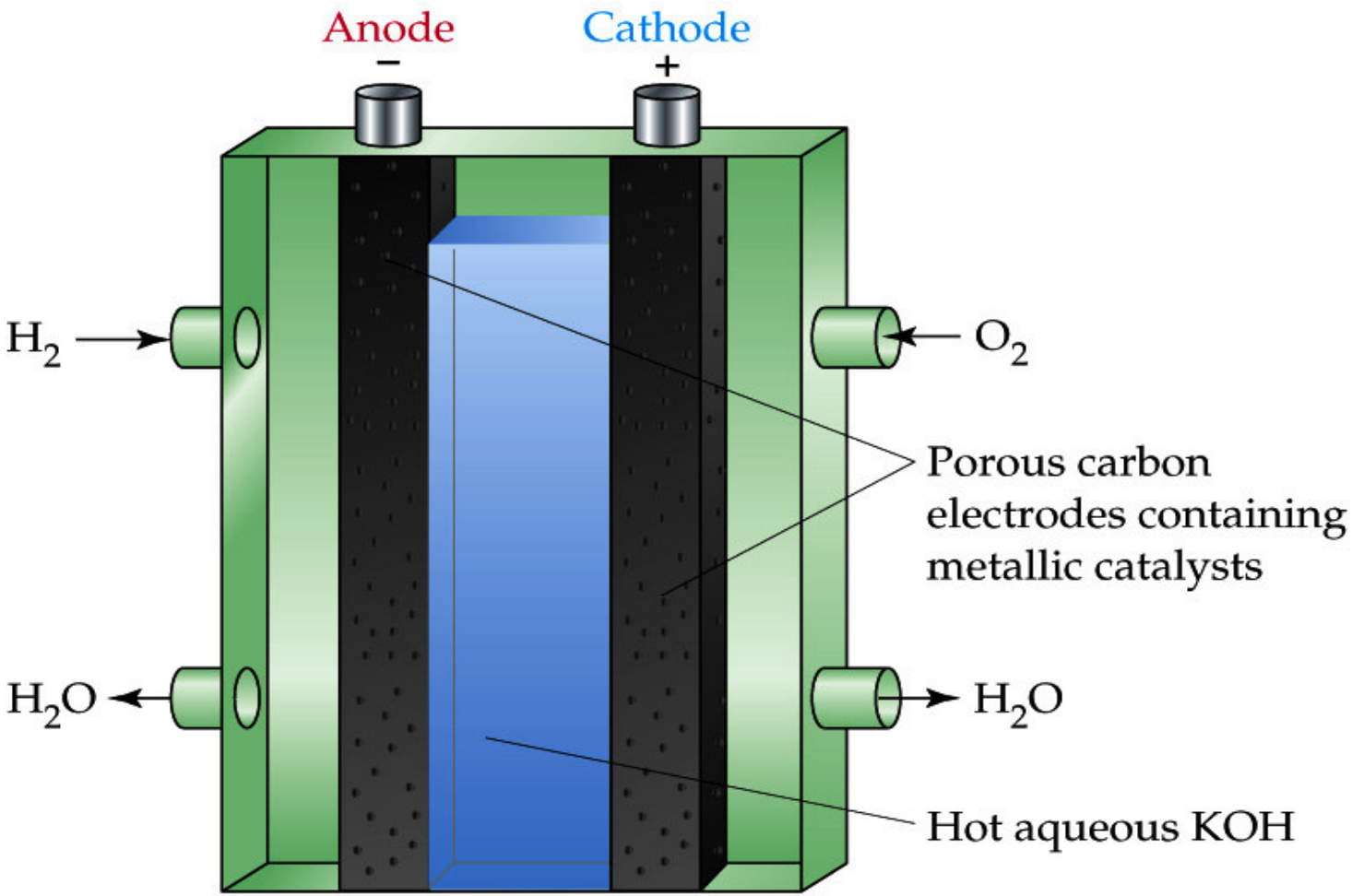
Cathode: Porous carbon containing metallic catalysts

$$O_2(s) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$$

Electrolyte: Hot aqueous KOH solution.

Cell Potential: 1.23V

Fuel Cell:


Anode
−
Cathode
+
H_2
O_2
Porous carbon electrodes containing metallic catalysts
H_2O
H_2O
Hot aqueous KOH

Fuel Cell:

- Fuel cells are **not batteries** because they are not self contained.
- Fuel cells are typically have about 40% conversion to electricity, the remainder is lost as heat.
- Excess heat can be used to drive turbine generators.

Electrolysis

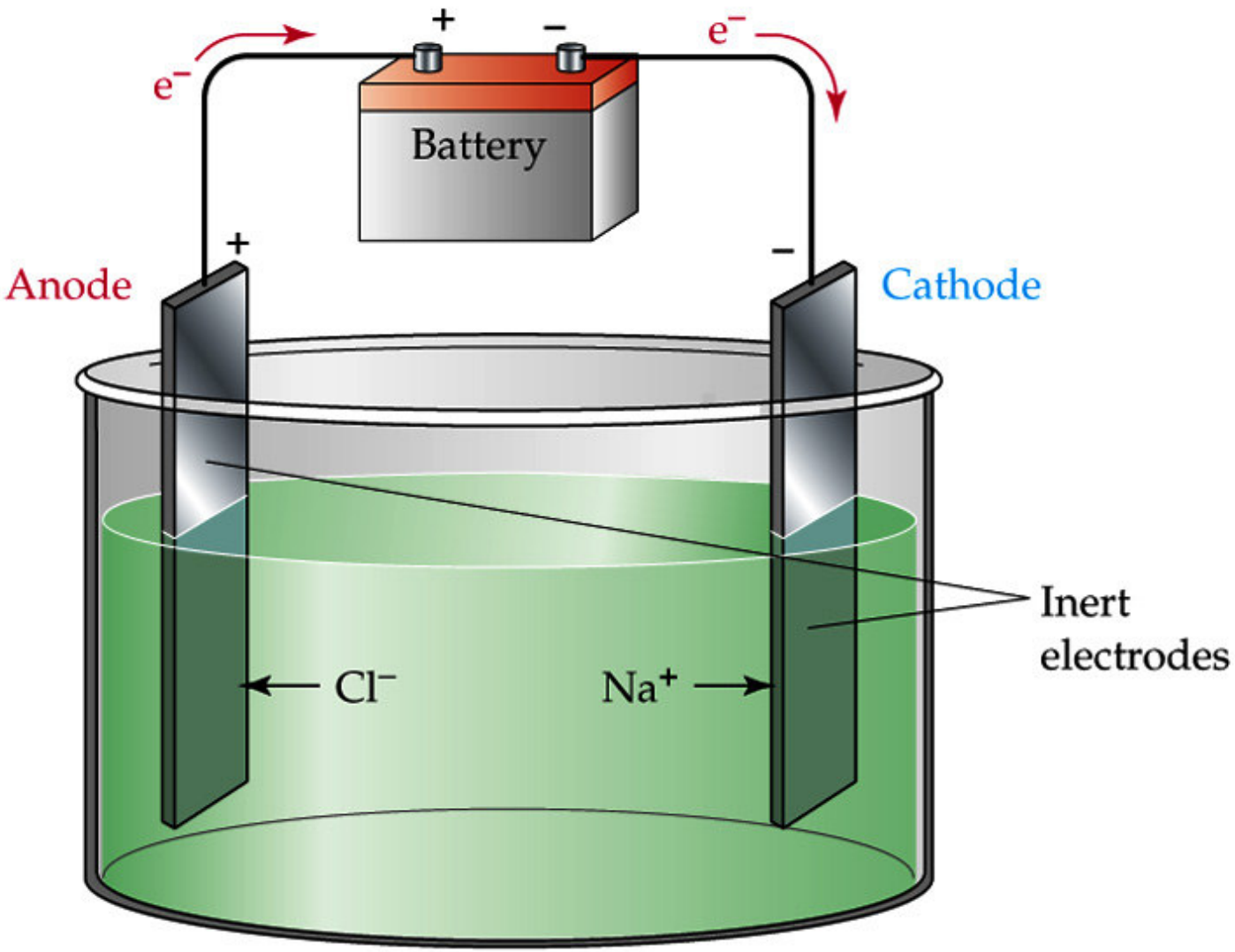
It is the OPPOSITE of batteries

It does not occur UNLESS you **make it happen**

Electrolysis 101

- **Electrolysis:** is the process in which electrical energy is used to drive a ***nonspontaneous*** chemical reaction.
- Processes in an electrolytic cell are the ***reverse*** of those in a galvanic cell.

Electrolysis of Molten Sodium Chloride:



$2\,Cl^-(l) \longrightarrow Cl_2(g) + 2\,e^-$

$2\,Na^+(l) + 2\,e^- \longrightarrow 2\,Na(l)$

Electrolysis of Water

- **Anode: Water is oxidized to oxygen gas.**

  $2\ H_2O(l) \rightarrow O_2(g) + 4\ H^+(aq) + 4e^-$

- **Cathode: Water is reduced to hydrogen gas.**

  $4\ H_2O(l) + 4e^- \rightarrow 2\ H_2(g) + 4\ OH^-(aq)$

Quantitative Electrolysis:

The amount of substance produced at an electrode by electrolysis depends on the quantity of charge passed through the cell.

Reduction of **1 mol of sodium ions** requires **1 mol of electrons** to pass through the system.

The charge on 1 mol of electrons is 96,500 Coulombs.

To determine the moles of electrons passed measure current and time

Charge (C) = Current (A) x Time (s)

- Because the charge on 1 mol of e^- is 96,500C, the number of moles of e^- passed through the cell is:

$$\text{Moles of } e^- = \text{Charge (C)} \times \frac{1 \text{mole } e^-}{95{,}000 \text{ C}}$$

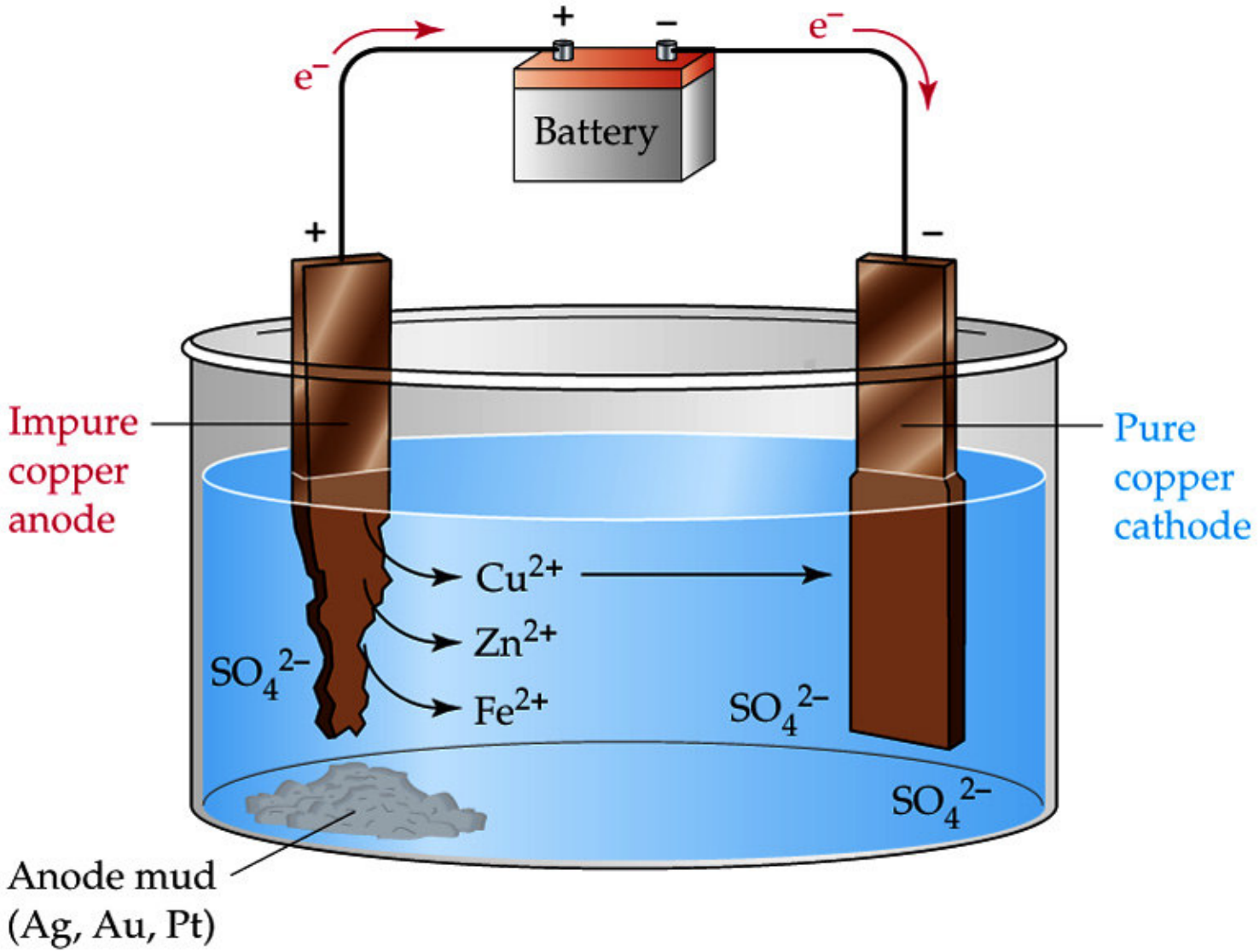
Electrolysis

A constant current of 30.0A is passed through an aqueous solution of NaCl for a time of 1.00 hour. How many grams of NaOH and how many liters of Cl_2 gas at STP are produced?

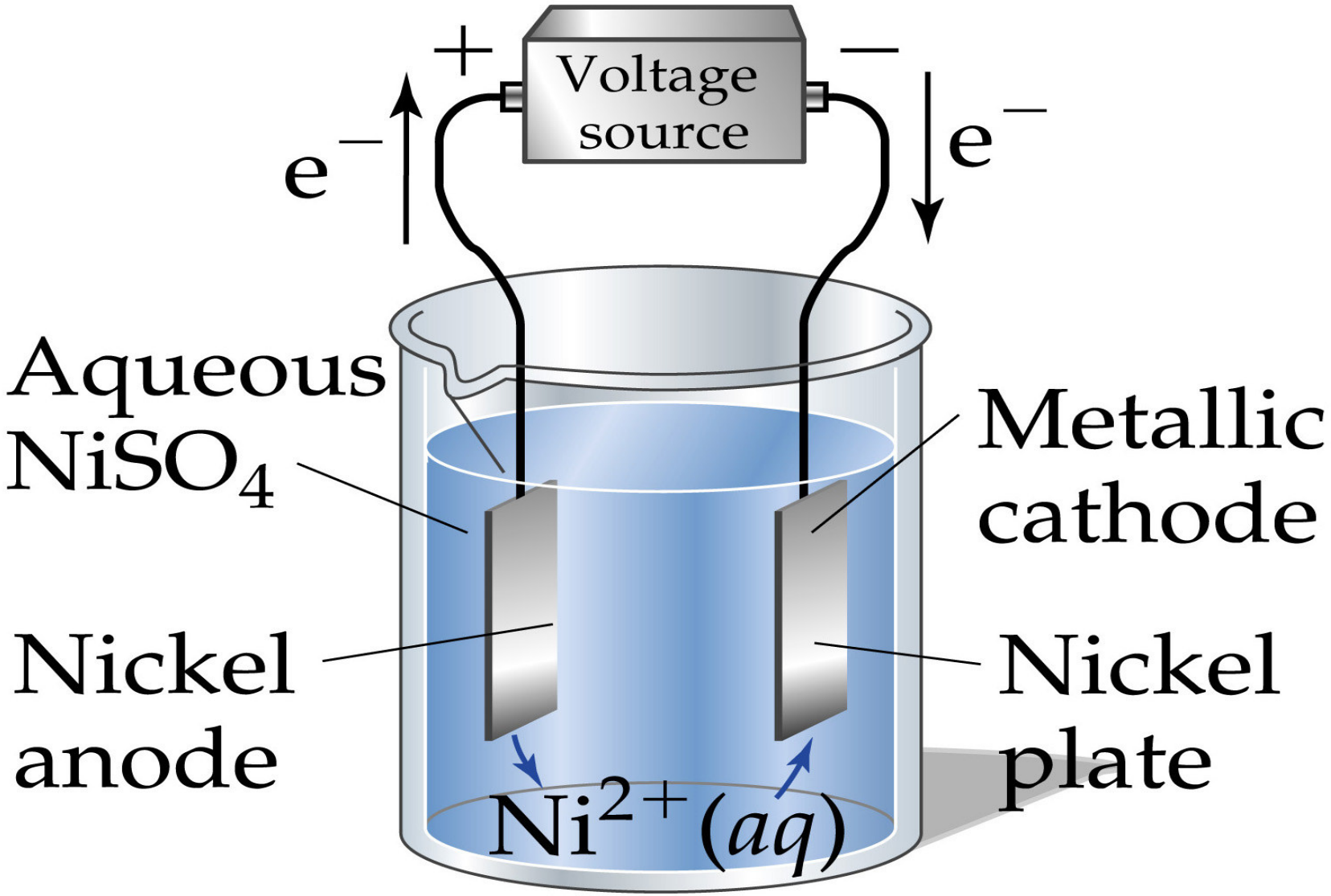
Electrolysis

A constant current is passed through a cell containing molten $MgCl_2$ for 1 hour. If 71.0 g of Cl_2 is obtained, How many grams of Mg is obtained and what is the current in amperes?

Electrolysis Applications: **Electroplating**



External source to drive reaction



Electrolysis Electroplating

Ni electrode and another metallic electrode in a solution of Ni^{2+} (aq)

- Anode: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$
- Cathode: $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s)$
- Ni plates on the inert electrode

Balancing Oxidation-Reduction Reactions

Balancing Equations by the Method of Half Reactions

Balancing Oxidation-Reduction Reactions

1. Write down the two half reactions.
2. Balance each half reaction:
   a. First with elements other than H and O
   b. Then balance O by adding water.
   c. Then balance H by adding H^+.
   d. Finish by balancing charge by adding electrons

Balancing Oxidation-Reduction Reactions

3. Multiply each half reaction to make the number of electrons equal.
4. Add the reactions and simplify.
5. Check!

Balancing Oxidation-Reduction Reactions

- Reaction of an acidic solution of $Na_2C_2O_4$ (sodium oxalate, colorless) with $KMnO_4$ (deep purple).

- MnO_4^- is reduced to Mn^{2+} (pale pink) while the $C_2O_4^{2-}$ is oxidized to CO_2.

Balancing Oxidation-Reduction Reactions

For $KMnO_4 + Na_2C_2O_4$

1. The two incomplete half reactions are

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

$$C_2O_4^{2-}(aq) \rightarrow 2CO_2(g)$$

For $KMnO_4 + Na_2C_2O_4 \rightarrow$

1. For the MnO_4^- half reaction

$$MnO_4^-(aq) \rightarrow Mn^{2+}(aq)$$

2. Adding water and H^+ yields

$$8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O$$

There is a 7+ charge on the left and 2+ on the right.

So 5 electrons need to be added to the left:

$$5e^- + 8H^+ + MnO_4^-(aq) \rightarrow Mn^{2+}(aq) + 4H_2O$$

Balancing Reactions

- In the oxalate reaction, there is a 2- charge on the left and a 0 charge on the right, so we need to add two electrons:

$$C_2O_4^{2-}(aq) \rightarrow 2CO_2(g) + 2e^-$$

3. To balance the 5 electrons for permanganate and 2 electrons for oxalate, we need 10 electrons for both. Multiplying gives:

$$10e^- + 16H^+ + 2MnO_4^-(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O$$

$$5C_2O_4^{2-}(aq) \rightarrow 10CO_2(g) + 10e^-$$

Balancing Reactions

4. Adding gives:

$$16H^+(aq) + 2MnO_4^-(aq) + 5C_2O_4^{2-}(aq) \rightarrow 2Mn^{2+}(aq) + 8H_2O(l) + 10CO_2(g)$$

5. Which is balanced!

Balancing Oxidation-Reduction Reactions

Balancing Equations for Reactions Occurring in Basic Solution

- We use OH^- and H_2O rather than H^+ and H_2O.
- The same method as above is used, but OH^- is added to "neutralize" the H^+ used.