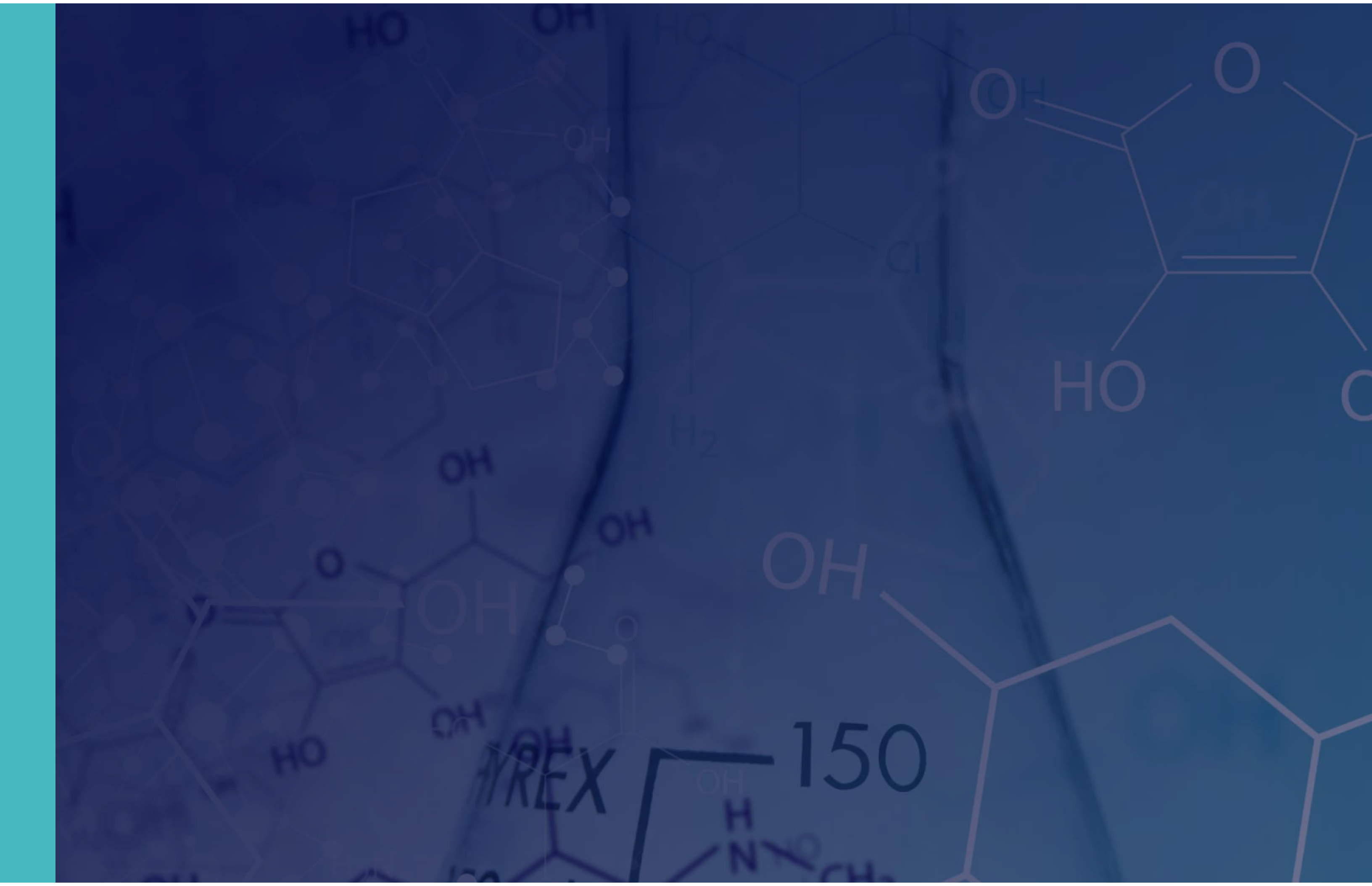

UWO

CHEM 1302

Winter 2024, Chapter 14 Notes

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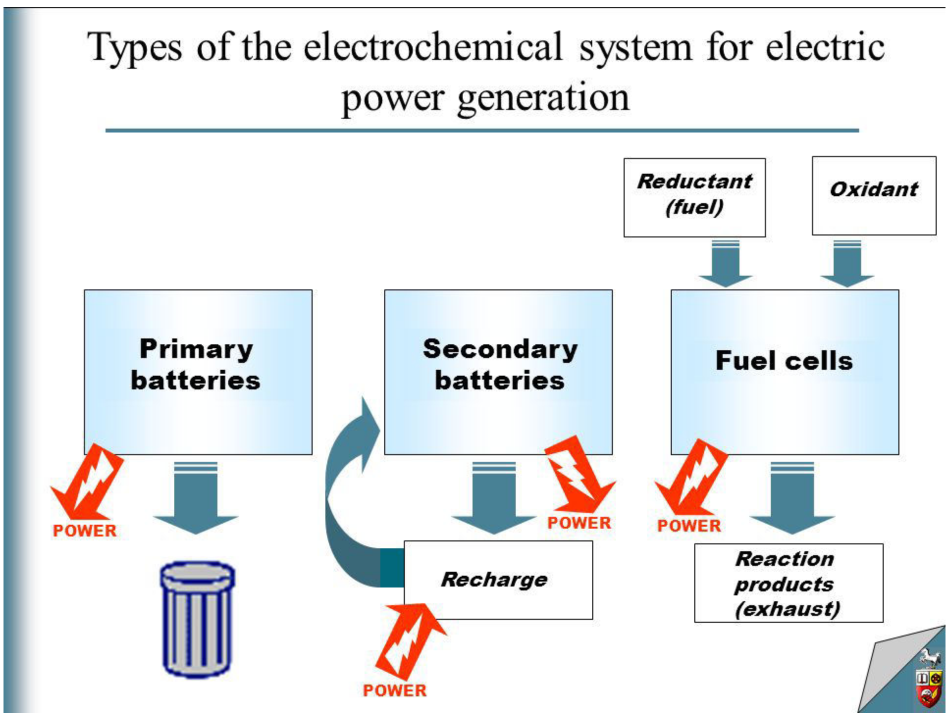
14. Batteries

14.1

14.1 Batteries and Other Applications

14.1.1

There are 3 main types of batteries:

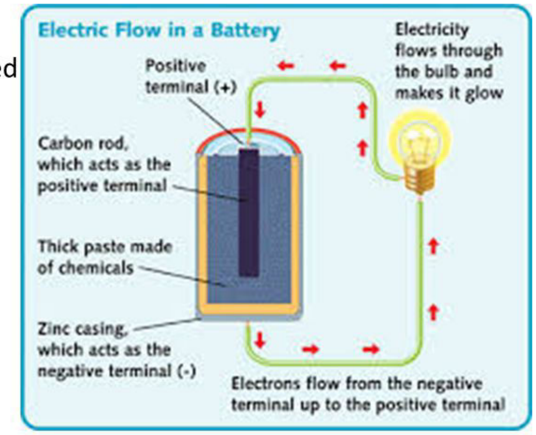


1) Primary batteries

- non-rechargeable (discarded when the cell reaction reaches equilibrium)
- one way (irreversible)
- self-contained series of voltaic cells
- ex. modern alkaline battery, silver button battery (like the one we see in watches), lithium battery,
- alkaline cells is about 1.5V

Primary Cell

- One use (non-rechargeable/disposable)
- Chemical reaction used, can not be reversed
- Used when long periods of storage are required
- Lower discharge rate than secondary batteries
- Use:
  smoke detectors, flashlights, remote controls



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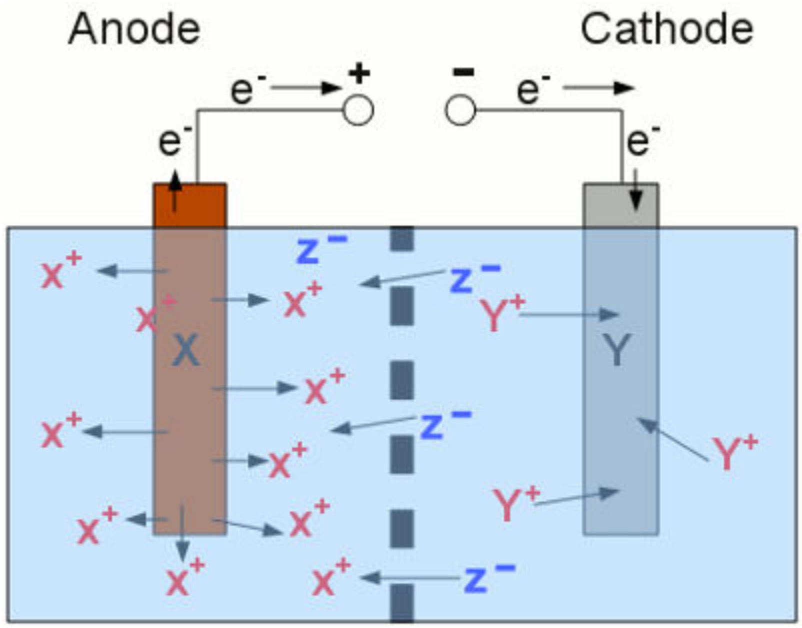
2) Secondary Batteries

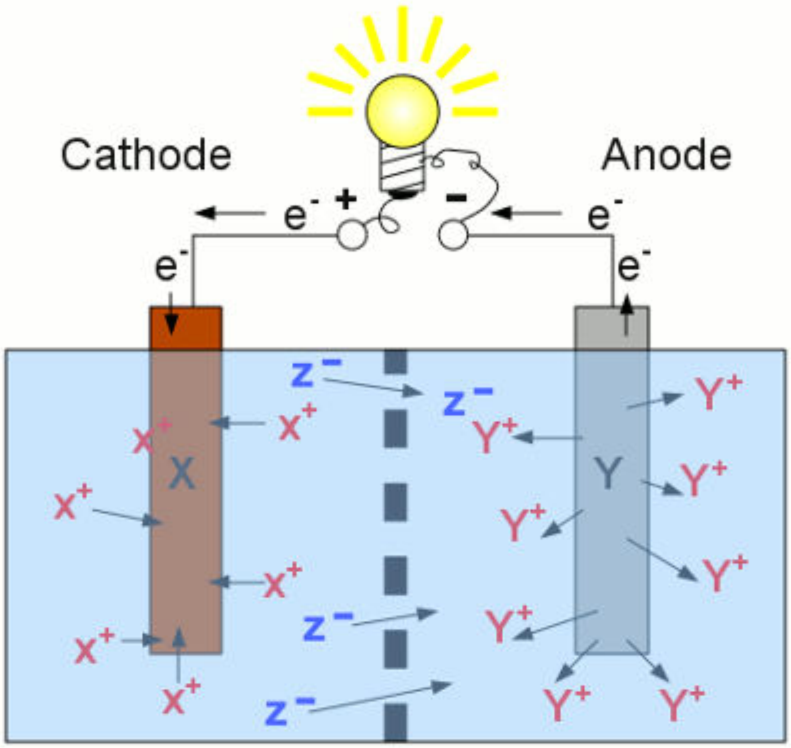
- rechargeable
- self-contained series of voltaic cells
- electrical current supplied to reverse the cell reaction
- ex: lead-acid batteries (car batteries), nickel-metal hydride batteries (power tools), lithium ion battery (laptops, cell phone)

a) Charging

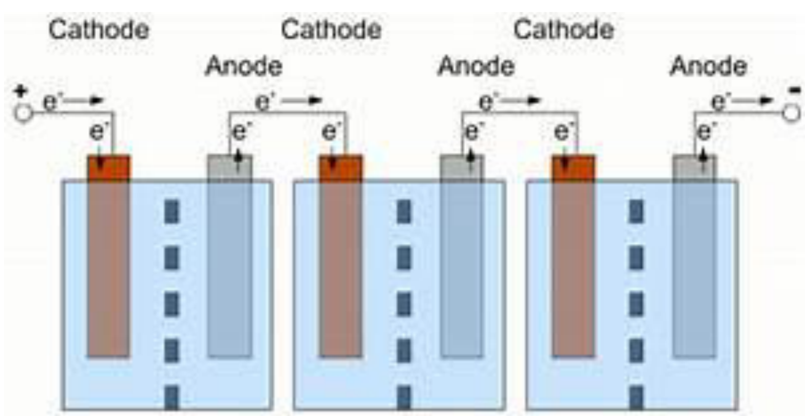


- electrolysis (not a voltaic cell)
- need a voltage source
- reaction stops when the cathodic solution runs out of cations

b) Discharging



- this acts like a galvanic cell (spontaneous electron flow)
- reverse process is occurring, metal on right is now the anode, metal on left is the cathode



3) Fuel Cells

- non self-contained voltaic cells
- controlled combustion (O2 and H2 enter cells, H20 leaves cells)
- can flow chemicals through system to generate electricity

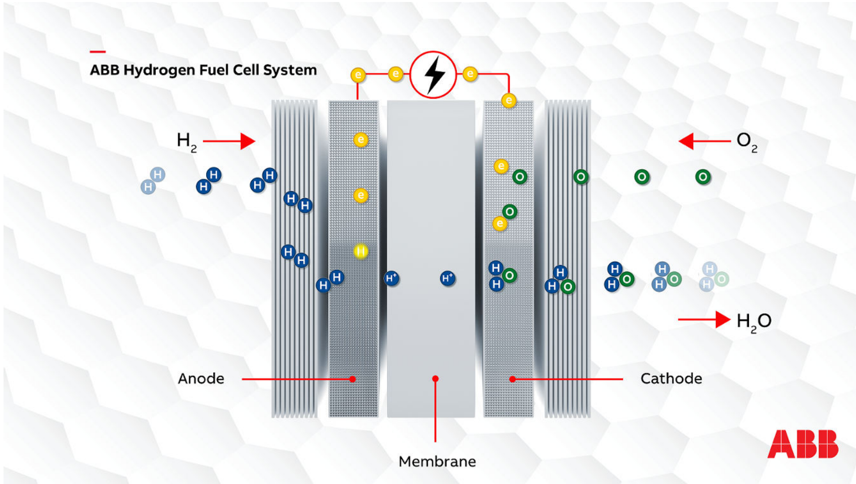


ABB Hydrogen Fuel Cell System
H2
O2
H2O
Anode
Cathode
Membrane
ABB

14.1.2

Rechargeable Batteries

- We will soon see that rechargeable batteries can act as both galvanic and electrolytic cells!
- Think of our phone batteries!
  - If we unplug it from our charger in the morning, then from that point onwards the battery is **discharging** (we will soon see that this is a **spontaneous** process)

  - Then at night, when we go to plug it into the charger, the battery is **charging** (this part requires a battery, which is an external source of energy so this part is **non-spontaneous**)

- Based on this info do you think the discharging process is more like a galvanic or electrolytic cell? ____________

- What about for the charging process? ____________

Let's consider a rechargeable lead-acid battery...

The overall equation for this battery is:

$$Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \text{ eqm } 2PbSO_4(s) + 2H_2O(s)$$

1) Draw in the oxidation states for Pb in the above equation

2) Now let's consider the discharge process

We have these 2 reactions:

a) $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$ E^o ____________ $=0.13V$

b) $Pb^{4+}(aq) + 2e^- \rightarrow Pb^{2+}(aq)$ E^o ____________ $=1.83V$

$E^o_{redox}=$ __________

Is this process spontaneous or not?

We can now draw out the cell for this: ____________ cell

Make one electrode Pb(s) and the other PbO_2(s) (which is Pb^{4+})

- Is a battery needed for this reaction? _______
- label the anode and the cathode
- Show the direction in which electrons flow
- Label the charges of both the anode and the cathode

3) Now let's consider the charging process

To charge the battery, we have the same 2 reactions but want them to run in reverse:

a) Pb^{2+}(aq) + 2e- -->Pb(s) E^o______________=__________V

b) Pb^{2+}(aq) --> 2e- + Pb^{4+}(aq) E^o______________=__________V

E^oredox=__________V

Is this process spontaneous or not?

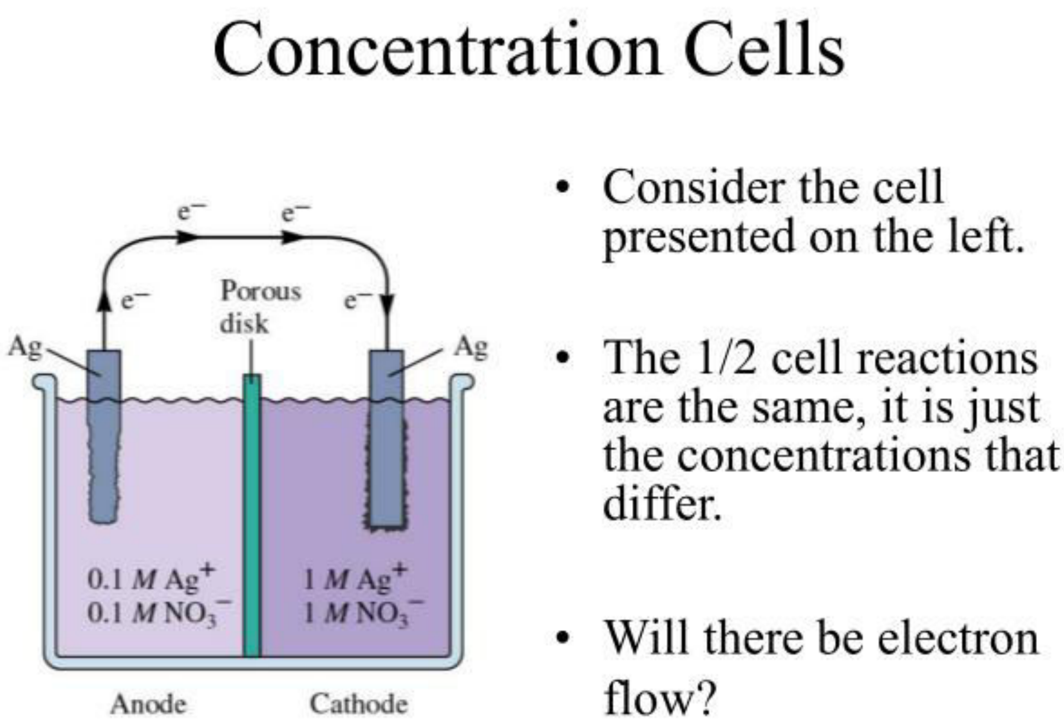
We can now draw out the cell for this: ________________ cell

Make one electrode $PbSO_4$ (which is also Pb^{2+}) and the other electrode the same.

- Is a battery needed for this reaction? ________
- Label the anode and the cathode
- Show the direction in which electrons flow
- Label the charges of both the anode and the cathode
- What are the charges on the battery (if there is one?)

14.1.3

Concentration Cells



Are these standard or non-standard conditions? ____________________

If there is electron flow, when would the flow of electrons stop?

Problem:

There are two half-cells connected together. Both half-cells have a zinc electrode

and $ZnCl_2$ electrolyte. The concentration of Zn^{2+} in one cell is 0.20 M and in the other cell is more dilute but unknown. The voltage across the electrodes is 0.0289 V. Find the unknown

[Zn^{2+}].

We will need to use the Nernst equation since the conditions are ________________!
Ecell=Eocell - (RT/nF)(lnQ)