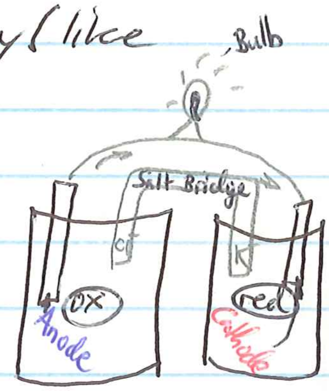
20.1 Electrochemistry

For redox rxns to occur, electrons have to move from an oxidized reactant to a reduced reactant.

* We can use electrochemistry techniques to allow us to move electrons between metals to convert chemical energy to electrical energy (like a battery) & back.



To make this work you need to make a voltaic cell - which houses the ox rxn in one compartment & red. rxn in another.
- known as half-cells

- wire leads are connected to the electrodes (metals) to allow electrons to move between them. ox → red
- a salt bridge is used to provide ions necessary to complete the circuit (but half-cells don't mix - only ions)

- Electrons move from ox. cell to red. cell
- Negative ions move to ox cell & positive ions to red. cell

• Name of oxidation electrode is anode
• Name of reduction electrode is Cathode

OaRCa

Differences in electrical poteritial energy allow current to flow via cell potential and generate voltage.

- the greater the difference = more voltage

Reduction Potential = tendancy of a substance to gain electrons

see Table 20.1 pg 712

- this can be determined using standard red. potentials

- The more negative, the less potential it has to absorb electrons
- the more positive, the more likely it is to absorb electrons.

Rules for determining Cell Potentials

1. Find the 2 half rxns on table 20.1
2. Compare the 2 half-cell potentials
   - The higher (or less negative) potential will proceed as the reduction & the other the oxidation

3. Write the half-reaction equations with the red & oxidation (often in reverse of table 20.1)

4. Balance the electrons in the half-cell eqns by multiplying each by a least common factor
- Add the equations

5. Plug the potentials into the following eqn

$$\boxed{E^\circ_{cell} = E^\circ_{Red} - E^\circ_{ox}}$$

E°_{cell} = cell potential (how much voltage made)

Once this is done, you will likely be asked to write the cell notation version of the voltaic cell rxn.

This uses a single vertical line to separate the reactants & products in the oxidation & reduction half cells. The add a double horizontal line to separate the half-cells. = salt bridge

salt bridge

ex) $\underbrace{Fe | Fe^{2+}}_{ox} || \underbrace{I_2 | I^-}_{red}$ ← doesn't need to be balanced.

* Spontaneous reactions are those whose cell potential is a positive voltage

* Non-spontaneous reactions are those whose cell potential is a negative voltage

ex) $Ag^{+}_{(aq)} + e^- \rightarrow Ag_{(s)}$ & $Ni^{2+}_{(aq)} + 2e^- \rightarrow Ni_{(s)}$

Table 20.1

1. half-cell potentials (Red. potentials)

$Ag^+ + e^- \rightarrow Ag_{(s)} = +0.7996$

$Ni^{2+} + 2e^- \rightarrow Ni_{(s)} = -0.257$

which has higher red. potential (is the reduction half-cell?)?

2. $Ag^+ \rightarrow Ag_{(s)}$ less negative $+0.7996$ = reduction

oxidation	Reduction
3. $Ni_{(s)} \rightarrow Ni^{2+} + 2e^-$	$Ag^+ + e^- \rightarrow Ag_{(s)}$
(reverse of red potential)	x 2

Balance the electrons

ox	an
4. $Ni_{(s)} \rightarrow Ni^{2+} + 2e$	$2Ag^+ + 2e^- \rightarrow 2Ag_{(s)}$

cell notation $Ni_{(s)} | Ni^{2+} || Ag^+ | Ag_{(s)}$

5. What is the cell potential?

$$\boxed{E^\circ_{cell} = E_{red} - E_{ox}}$$

$$E^\circ_{cell} = +0.7996 - (-0.257) = +1.0566\ V$$

Spontaneous rxn?

$Mg^{2+} + 2e^- \rightarrow Mg_{(s)}$ | $2H^+ + 2e^- \rightarrow H_2$

-2.372 0.000

ox red

$2H^+ + 2e^- \rightarrow H_2$ $Mg_{(s)} \rightarrow Mg^{2+} + 2e^-$

$E_0 = 0.00 - (-2.372) = 2.372$ ✓

$Mg \mid Mg^{2+} \parallel 2H^+ \mid H_2$ overall = $Mg_{(s)} + 2H^+ \rightarrow Mg^{2+} + H_2$

$Sn^{2+} + 2e^- \rightarrow Sn_{(s)}$ & $Fe^{3+} + 3e^- \rightarrow Fe_{(s)}$

-0.1375 -0.037

ox red

$2Fe^{3+} + 6e^- \rightarrow 2Fe_{(s)}$ & $3Sn_{(s)} \rightarrow 3Sn^{2+} + 6e^-$

-0.037 -0.1375

$E_0 = -0.037 - (-0.1375) = +0.1005$ ✓