

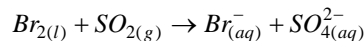
Chapter 20 – Electrochemistry

Kahoot!

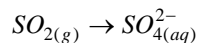
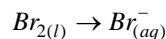
- The oxidation state of N in NH_4^+ is ____ . +1, 0, -1, **-3**
- The oxidation state of Mn in MnO_4^- is ____ . -1, +2, +4, **+7**
- The reducing agent in $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ is ____ . **Zn**, Cu^{2+} , Zn^{2+} , Cu
- The oxidizing agent in $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$ is ____ . Zn, **Cu^{2+}** , Zn^{2+} , Cu
- To balance $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$, ____ electrons need to be added to ____ side. 5; product, 2; product, 2; reactant, **5; reactant**
- The purpose of the salt bridge is to ____ . Provide H^+ ions needed to balance charges, **maintain neutrality by allowing the flow of ions**, serve as the site for oxidation to occur, serve as the site for reduction to occur.
- Reduction occurs at the ____ . anode, **cathode**, electrode, electrolytic cell
- Oxidation occurs at the ____ . **anode**, cathode, electrode, electrolytic cell
- Which is the strongest oxidizing agent? **F_2** , Cl_2 , Br_2 , I_2
- Which is the strongest reducing agent? Zn, Al, Na, **Li**
- If $\text{Cl}_2 \rightarrow 2\text{Cl}^-$ has $E = +1.36\text{V}$, and $\text{I}_2 \rightarrow 2\text{I}^-$ has $E = +0.54\text{V}$ then ____ . Cl_2 will reduce I^- to I_2 , **Cl_2 will oxidize I^- to I_2** , I_2 will reduce Cl^- to Cl_2 , I_2 will oxidize Cl^- to Cl_2
- If the value of E_{cell} is positive then the reaction is ____ . at equilibrium, **spontaneous**, nonspontaneous, very fast
- A ____ cell uses an external energy source to produce a redox reaction. galvanic, voltaic, **electrolytic**, anionic
- The Nernst equation is most useful when ____ are nonstandard. Oxidizing agents, reducing agents, **ion concentrations**, temperatures

Whiteboard Examples:

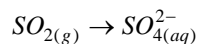
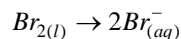
Balancing Redox Example I: Balance the following redox reaction given the solution is acidic:



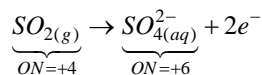
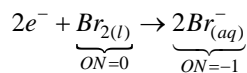
-- step 1 - separate the eqn into the ox'n and red'n half-rxns



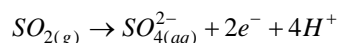
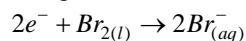
-- step 2 - balance the non-O and H atoms



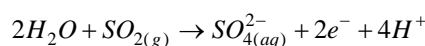
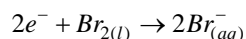
-- step 3 - determine the ON of the non-O and H atoms and balance charge with e-



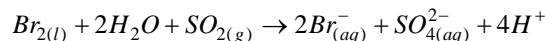
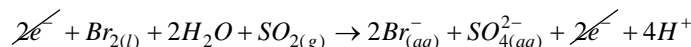
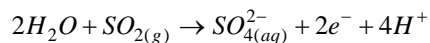
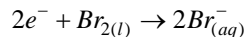
-- step 4 - balance charge for each half-rxn using H^+



-- step 5 - mass balance O and H with water

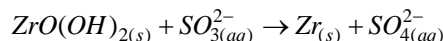


-- step 6 - use a constant to make sure electrons cancel and add the two equations

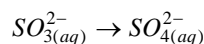
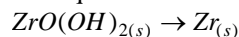


-- you can check and make sure your equation is correct by looking to make sure both sides of the equation are balanced both by charge and number of atoms

Balancing Redox Example 2 - Balance the following reaction in basic solution:

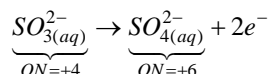
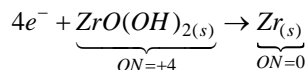


-- step 1 - separate the eqn into the ox'n and red'n half-rxns

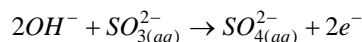
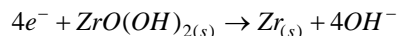


-- step 2 - balance the non-O and H atoms - NA

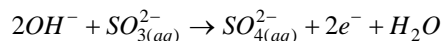
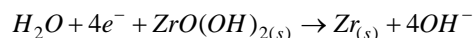
-- step 3 - determine the ON of the non-O and H atoms and balance charge with e-



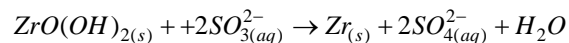
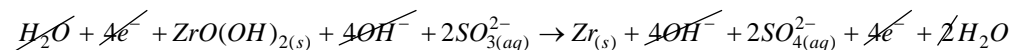
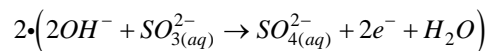
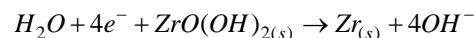
-- step 4 - balance charge for each half-rxn using OH⁻



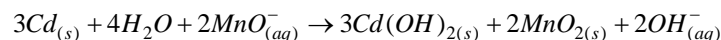
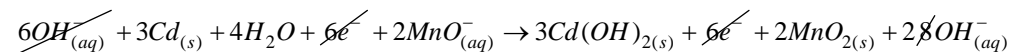
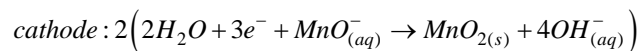
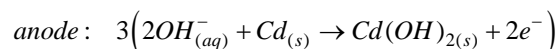
-- step 5 - mass balance O and H with water



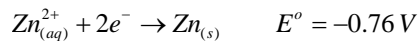
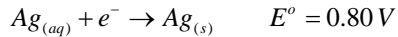
-- step 6 - use a constant to make sure electrons cancel and add the two equations



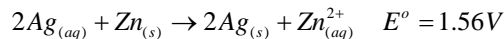
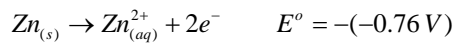
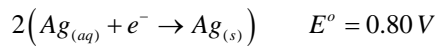
Voltaic Example: A voltaic cell with a basic electrolyte is based on the oxidation of $Cd_{(s)}$ to $Cd(OH)_{2(s)}$ and the reduction of MnO_4^- to $MnO_{2(s)}$. Write the half-rxns, the balanced cell reaction and draw a diagram of the cell.



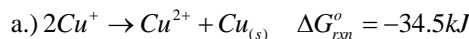
E_{cell} Example: Given the information below determine the voltage and write the balanced equation of the following voltaic cell.



Answer: If the cell is voltaic, it must be spontaneous so $E > 0$. Therefore, we must be oxidizing Zn – so we need to reverse the equation to oxidation for zinc and reverse the sign of the emf.

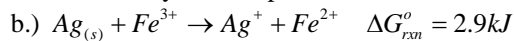


Gibbs Example: Calculate the E_{cell} for the following reaction and determine if the reaction is spontaneous.



$$E_{\text{cell}}^\circ = -\frac{\Delta G_{\text{rxn}}^\circ}{nF} = -\frac{-34.5 \text{ kJ} \times \frac{1000 \text{ J}}{\text{kJ}} \times \frac{1 \text{ C} \cdot \text{V}}{1 \text{ J}}}{1 \text{ mol } e^- \times 9.65 \times 10^4 \text{ C/mol}} = \boxed{0.358 \text{ V}}$$

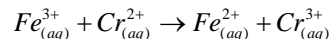
this system is spontaneous



$$E_{\text{cell}}^\circ = -\frac{\Delta G_{\text{rxn}}^\circ}{nF} = -\frac{2.9 \text{ kJ} \times \frac{1000 \text{ J}}{\text{kJ}} \times \frac{1 \text{ C} \cdot \text{V}}{1 \text{ J}}}{1 \text{ mol } e^- \times 9.65 \times 10^4 \text{ C/mol}} = \boxed{-0.0301 \text{ V}}$$

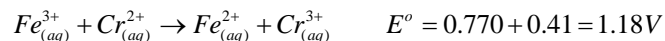
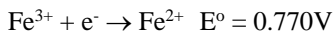
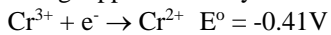
this system is nonspontaneous

Nernst Example: Calculate the E_{cell} at 298K for the cell based on the following:



$$[Fe^{3+}] = [Cr^{2+}] = 1.50 \times 10^{-3} \text{ M}, [Fe^{2+}] = [Cr^{3+}] = 2.5 \times 10^{-4} \text{ M}$$

Using Appendix E in your book



$$E_{\text{cell}} = 1.18 \text{ V} - \frac{0.0592}{1} \log \frac{[2.5 \times 10^{-4}]^2}{[1.50 \times 10^{-3}]^2} = 0.938 \text{ V}$$