


Particle-Level Electrochemistry!

In this activity, you will use animations and simulations to visualize redox reactions at the particle level and understand the setup and function of electrochemical cells.

Link to the simulations for this activity: <https://tcel-hu-berlin.github.io/silc/redox/>
(Google Chrome recommended)

 = hint available on the website

 = discussion point

BONUS = optional question

Preparation Exercises

1. Open the page for Animation A and read the overview.
 - a) View the first video, which shows the reaction between a strip of copper metal and a solution of silver nitrate (AgNO_3). Note your observations about what happens (i) to the solution and (ii) to the solid phase during the reaction.
 - b) View the animation below the video and describe the processes that are occurring at the submicroscopic level. Pay attention to the difference between ions and atoms. Use the animation to explain your qualitative observations in (a).
2. Define the terms (a) oxidation, (b) reduction, (c) anode, and (d) cathode.
3. Consider the species $\text{Cu}_{(s)}$, $\text{Cu}^{2+}_{(aq)}$, $\text{Ag}_{(s)}$, $\text{Ag}^{+}_{(aq)}$ and $\text{NO}_3^{-}_{(aq)}$ in the reaction (Animation A). Give the species that fulfills each of the conditions (a-h). Note that a species may fulfil more than one of the roles below.
 - a) is oxidized:
 - b) is reduced:
 - c) is an electron donor:
 - d) is an electron acceptor:
 - e) acts as oxidizing agent:
 - f) acts as reducing agent:
 - g) is part of a redox pair along with $\text{Ag}^{+}_{(aq)}$:
 - h) acts as a spectator ion and is not involved in the reaction:

4. Open the page for Simulation B and read the “crash course” on galvanic cells. In the simulation, set up a galvanic cell that has iron connected to the “-” side and silver connected to the “+” side of the voltmeter. Use the electrochemical series (see Help menu) and the steps (a-e) below to determine the net redox reaction and the voltage of this galvanic cell.
- a) Write out the reduction half-reaction for the “+” side (Ag) and note its corresponding standard reduction potential (see electrochemical series):
 - b) Write out the reduction half-reaction for the “-” side (Fe) and note its corresponding standard reduction potential (see electrochemical series):
 - c) Calculate the cell voltage $\Delta E_{\text{cell}}^o = E_+^o - E_-^o$. For this calculation, you should use only the reduction potentials, regardless of which side the reduction actually occurs on. Verify that your calculated cell voltage has the same sign as the voltage reading in Simulation B.
 - d) A positive cell voltage means that the reduction is occurring on the “+” side of the voltmeter. This would mean that the reaction in (b) occurs in the opposite direction (oxidation). A negative cell voltage means that the reduction occurs on the “-” side of the voltmeter. In this case, the reaction in (a) would occur in the opposite direction as it is written. Flip one of the reactions (a) or (b) accordingly.
 - e) If necessary, multiply one or both of the half-reactions by an integer in order to equalize the number of electrons. Then, combine the two reactions, ensuring charge balance.



5. Open Simulation C and read the “crash course” to familiarize yourself with the construction of an electrolytic cell and the controls of the simulation.
6. Review the rules for determining oxidation states in compounds and for balancing redox reactions. A guideline for balancing redox reactions can be found in the help menu.



Main Exercises

1. The electrochemical series lists the reduction reactions for a variety of chemical species in order of decreasing standard reduction potential. In thermodynamics, the standard potential of a reaction (E°) is related to the free energy (ΔG°) and the equilibrium constant (K) of the reaction:

$$\Delta G^\circ = -nFE^\circ = -RT\ln(K)$$

Here, n is the number of moles of electrons involved in the redox reaction, F (Faraday's constant) is the charge in Coulomb carried by one mole of electrons, T is the temperature in Kelvin, and R is the gas constant in $\text{J K}^{-1} \text{mol}^{-1}$. Use the electrochemical series (Help Menu) to answer the following questions):



- a) Compare the half-reactions for the reduction of Cu^{2+} to $\text{Cu}_{(s)}$ and the reduction of Fe^{2+} to $\text{Fe}_{(s)}$. Which reaction is more thermodynamically favorable (or product-favored)? Justify your answer.
- b) How is the reduction potential related to the tendency of a redox pair to undergo reduction?
- c) The result for $\Delta E^\circ_{\text{cell}}$ in Preparation Question (4c) is positive. What does this tell you about...
 - i. ...the ability of $\text{Ag}_{(s)}$ to reduce Fe^{2+} ?
 - ii. ...the ability of $\text{Fe}_{(s)}$ to reduce Ag^+ ?

Justify your answer by addressing the relationship between E° and ΔG° .

2. Use Simulation B to set up a galvanic cell with copper on the minus-side and silver on the plus-side of the voltmeter. Choose the correct answer for each of the questions below and verify each answer using the simulation:
- Which side acts as cathode in this configuration?
 Cu^{2+}/Cu Ag^+/Ag
 - In which direction are electrons flowing in the circuit?
 from Cu to Ag from Ag to Cu
 - In which direction are K^+ ions moving inside the salt bridge?
 from Cu to Ag from Ag to Cu
3. Answer the same questions for a configuration in which the connections (plus and minus) of the voltmeter are switched. Afterwards, check your answers using the simulation.
- Which side will act as cathode in the new configuration?
 Cu^{2+}/Cu Ag^+/Ag
 - In which direction will electrons flow in the circuit?
 from Cu to Ag from Ag to Cu
 - In which direction will K^+ ions move inside the salt bridge?
 from Cu to Ag from Ag to Cu
4. For the following redox pair combinations, make predictions about (i) which half-reactions should occur at the plus and minus sides and (ii) which cell potential $\Delta E^\circ_{\text{cell}}$ (with the correct sign!) should be expected. In addition, provide (iii) the net redox reaction that will occur spontaneously and (iv) the direction in which electrons will flow in the circuit. Afterwards, check your predictions with Simulation B.



Example: Redox pair at - side: Zn (s) and $\text{Zn(NO}_3)_2 \text{ (aq, 1 M)}$
 Redox pair at + side: Fe (s) and $\text{Fe(NO}_3)_2 \text{ (aq, 1 M)}$

- (i) Half-reaction at + side (with E°_{red}):

$$\text{Fe}^{2+}_{(\text{aq})} + 2\text{e}^- \rightarrow \text{Fe}_{(\text{s})} \quad (E^\circ_{\text{red},+} = -0.41 \text{ V})$$
- Half-reaction at - side (with E°_{red}):

$$\text{Zn}_{(\text{s})} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + 2\text{e}^- \quad (E^\circ_{\text{red},-} = -0.76 \text{ V})$$
- (ii) Cell potential:

$$\Delta E^\circ_{\text{cell}} = E^\circ_{\text{red},+} - E^\circ_{\text{red},-} = -0.440 - (-0.763) = + 0,35 \text{ V}$$
- (iii) Spontaneous net reaction and argumentation:

$$\Delta E^\circ_{\text{cell}} > 0, \Rightarrow + \text{ side: cathode, - side: anode}$$

$$\Rightarrow \text{Fe}^{2+}_{(\text{aq})} + \text{Zn}_{(\text{s})} \rightarrow \text{Zn}^{2+}_{(\text{aq})} + \text{Fe}_{(\text{s})}$$
- (iv) Direction of electron flow:
from anode to cathode; here, from - to +

- a) Redox pair at - side: Cu (s) and Cu(NO₃)₂ (aq, 1 M)
 Redox pair at + side: H₂ (g, 1 bar) and HCl (aq, 1 M) on a Pt electrode.

(i) Half-reaction at + side (with E^o_{red}):

Half-reaction at - side (with E^o_{red}):

(ii) Cell potential:

(iii) Spontaneous net reaction and argumentation:

(iv) Direction of electron flow:

- b) Redox pair at - side: O₂ (g, 1 atm) and HCl (aq, 1 M) on a Pt electrode
 Redox pair at + side: H₂ (g, 1 atm) and HCl (aq, 1 M) on a Pt electrode.

(i) Half-reaction at + side (with E^o_{red}):

Half-reaction at - side (with E^o_{red}):

(ii) Cell potential:

(iii) Spontaneous net reaction and argumentation:

(iv) Direction of electron flow:

BONUS

5. Consider the following combinations of metals and metal nitrate solutions.

- a) For each combination, determine whether a reaction will occur. If a reaction occurs, write down the balanced reaction and calculate the standard cell potential.

	Combination	React?	Reaction	$\Delta E^{\circ}_{\text{cell}}$
i	Zn _(s) and Fe(NO ₃) ₂	yes	$\text{Zn}_{(s)} + \text{Fe}^{2+}_{(aq)} \rightarrow \text{Fe}_{(s)} + \text{Zn}^{2+}_{(aq)}$	0.35 V

ii	$\text{Hg}_{(l)} \text{ and } \text{Ni}(\text{NO}_3)_2$			
iii	$\text{Ag}_{(s)} \text{ and } \text{NaNO}_3$			
iv	$\text{Na}_{(s)} \text{ and } \text{Pb}(\text{NO}_3)_2$			
v	$\text{CuNO}_3 \text{ and CuNO}_3$ (yes, you read correctly!)	yes		



b) What is the term for the type of reaction in (v)? What is the term for the opposite type of reaction?

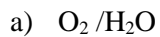
6. Determine the oxidation states of the highlighted elements in the compounds (a-g).



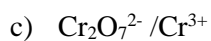
BONUS

BONUS

7. The Help Menu contains a guideline for balancing redox reactions. Use it to balance the half-reactions for the following redox pairs:



BONUS



8. Let's assume we want to "split" water into hydrogen and oxygen gas. This is a sought-after process in industry and research, largely due to the value of hydrogen gas as a clean fuel. However, as you may have seen in Main Exercise 4b, the water splitting reaction is not thermodynamically favorable.

a) Write out the reaction that should occur at the cathode during water splitting.



b) Write out the reaction that should occur at the anode during the same process.

c) Write out the net redox equation for the water splitting reaction.

d) The two poles of a 5-V battery are placed in an aqueous solution of H_2SO_4 to carry out a water splitting reaction. (H_2SO_4 acts as a source of H^+ ions and as an electrolyte.) On which pole of the battery (plus or minus) will H_2 gas be produced? Explain your answer.



9. Set up the electrolytic cell for water splitting in Simulation C.

a) Predict the minimum (threshold) voltage that should be applied to drive the electrolytic reaction forward.

b) Test your prediction by varying the applied voltage and observing the species consumed and produced (if at all) at both the macroscopic and the submicroscopic levels. Note your observations as you cross the threshold voltage.

10. Consider an electrolytic cell such as the above, but with the goal being to collect the H_2 and O_2 gases that are produced,



a) How much charge (in Coulomb) is required to produce 0.005 mol of $\text{H}_{2(g)}$?

b) How long would an electrolysis need to run, with an applied current of 2 A, to produce the above amount of $\text{H}_{2(g)}$?

c) What volume of $\text{H}_{2(g)}$ would be produced during this electrolysis under 1 atm and 298 K? You can treat hydrogen as an ideal gas. Check your calculation using Simulation C.

BONUS

BONUS

- d) What volumes of $\text{H}_{2(g)}$ and $\text{O}_{2(g)}$ will be produced if 18 mL of water is completely electrolyzed at 1 atm and 298 K?

**BONUS**

11. Alex would like to produce $\text{Na}_{(s)}$ and $\text{Cl}_{2(g)}$ from table salt, NaCl . He makes a 1 M aqueous solution of NaCl and places a platinum electrode and a graphite electrode in a beaker containing the solution. Looking at the standard reduction potentials for sodium ($E^\circ = -2.714 \text{ eV}$) and chlorine ($E^\circ = +1.359 \text{ V}$), Alex connects the electrodes to a 5-V battery, which should provide enough potential difference to drive the reaction forward. He expects chlorine gas to be formed at the anode and sodium metal to be produced at the cathode. Contrary to his expectations, when Alex connects the electrodes to the battery, he observes gas bubbles forming on both electrodes, while no signs of sodium formation can be observed at the cathode.



Construct Alex's setup in **Simulation C. Use** the magnifying glass to observe the reactions occurring at the two electrodes.

- Write out the chemical reaction that occurs at the anode and indicate the species responsible for the gas bubbles observed.
- Write out the chemical reaction that occurs at the cathode and indicate the species responsible for the gas bubbles observed.
- Discuss why the reaction occurring at the cathode is NOT what Alex (and maybe you!) originally expected. Consider the species present in solution and the tendency of each of those species to be reduced or oxidized.

BONUS

12. The superscript $^\circ$ denotes "standard conditions," i.e. the conditions at which the temperature is 298 K, the pressures of all gases are 1 bar, and the concentrations of all aqueous species are 1 M. Which mathematical equation does one use to accurately calculate the potential if concentrations are different than 1 M?