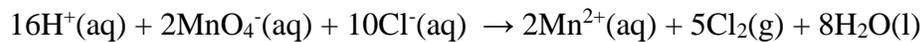


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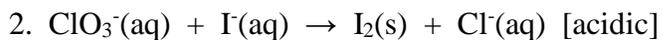
Redox worksheet part I

1. Consider the following balanced redox reaction:



- (a) Which species is being oxidized?
- (b) Which species is being reduced?
- (c) Which species is the oxidizing agent?
- (d) Which species is the reducing agent?
- (e) From which species to which is electron transfer occurring?

Balance the following redox reactions

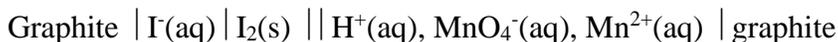


You should always show the following details in a cell diagram:

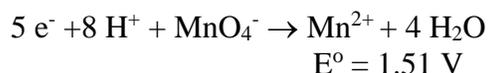
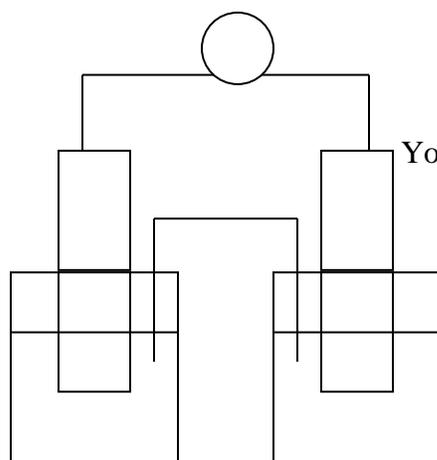
- Components of the half-cells: electrode materials, ions, and other substances involved in the reaction
- Electrode name (anode or cathode) and charge. By convention, the anode is on the left and the cathode is on the right
- Each half-reaction with its half-cell and the overall cell reaction
- Direction of electron flow in the circuit
- Nature of the ions and direction of ion flow in the salt bridge.

You should do the following if asked to write the shorthand notation for a cell:

- The components of the anode compartment (oxidation half-cell) are written to the left of the components of the cathode compartment (reduction half-cell)
- A double vertical line separates the half-cells and represents the wire and salt bridge.
- Within each half-cell, a single vertical line represents a phase boundary; for example, $\text{Zn(s)} \mid \text{Zn}^{2+}(\text{aq})$ indicates that the solid Zn is a different phase from the aqueous Zn^{2+} . A comma separates half-cell components in the same phase. For example, the notation for the voltaic cell shown to the right is



- Half-cell components appear in the same order as in the half-reaction, while electrode appear at the extreme left and right of the notation.

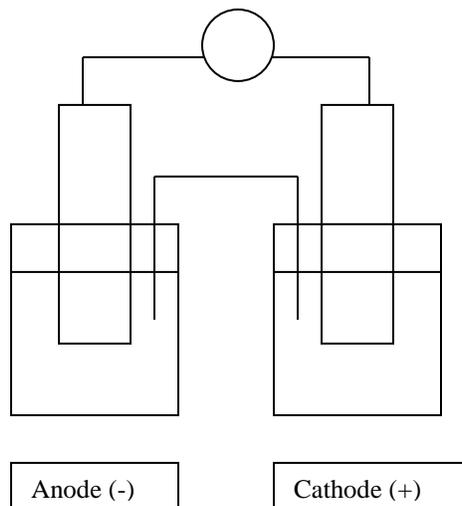


You will need this later on this worksheet

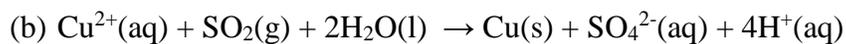
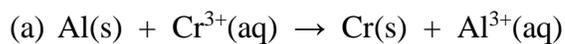
Anode (-)

Cathode (+)

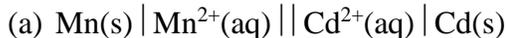
4. Diagram and write the notation for a voltaic cell that consists of one half-cell with a Cr bar in a $\text{Cr}(\text{NO}_3)_3$ solution, another half-cell with an Ag bar in an AgNO_3 solution, and a KNO_3 salt bridge. Measurements indicate that the Cr electrode is negative relative to the Ag electrode.



5. Write the cell notation for the voltaic cells that incorporate each of the following redox reactions:



6. Write a balanced equation for each of the following cell notations:



Determining cell potential and determining half-cell potentials are next. This can get tricky. You are usually given a STANDARD REDUCTION TABLE of half-reactions. But pay attention! Sometimes a STANDARD OXIDATION TABLE is presented just to see if you're paying attention. Here are some key points to a one of these tables:

- (a) All values are relative to the standard hydrogen electrode (0.00 V)
- (b) Since by convention, the half-reactions are written as reductions, all reactants are oxidizing agents and all products are reducing agents.
- (c) The E° value shown is for the half-reaction as written. The more positive the E° , the greater is the tendency for the half-reaction to occur.
- (d) Half-reactions are usually shown with equilibrium arrows to show that a half-reaction can occur as a reduction or as an oxidation (the sign changes when the reaction is reversed).
- (e) Changing the balancing coefficients of a half-reaction NEVER CHANGES THE VALUE OF E° !!
- (f) A spontaneous reaction will occur between an oxidizing agent (species on the left) and a reducing agent (species on the right) that lies below it on a standard reduction (or oxidation) table.

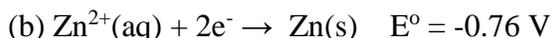
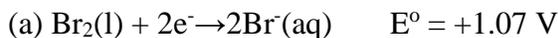
Some other things to consider from this point on:

(a) $E^\circ_{\text{cell}} = E^\circ_{\text{oxidation}} + E^\circ_{\text{reduction}}$

(b) $E^\circ_{\text{cell}} > 0$ means spontaneous ($\Delta G^\circ < 0$; $\Delta S^\circ_{\text{univ}} > 0$); this is required if the cell is to be a galvanic (voltaic) cell. You must always arrange the half-reactions accordingly to yield a spontaneous reaction.

Example 1

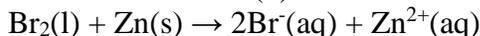
A voltaic cell incorporates the reaction between aqueous bromine and zinc metal. The standard reduction reactions are listed below:



Write the balanced equation for the redox reaction that occurs in the voltaic cell and give the value of the standard cell potential.

- The question specifically states that this cell is voltaic (galvanic) and therefore must be spontaneous ($E^\circ_{\text{cell}} > 0$).
- Both half-reactions are written as standard reductions. One, then must be reversed.
- When the two reactions (oxidation half-cell + reduction half-cell) are added, they must add to a (+) value because the cell is spontaneous.

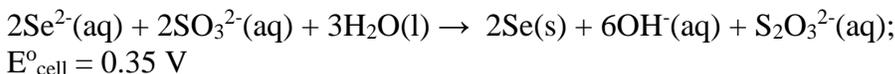
Solution → Reverse half-reaction (b). The result is



$$E^\circ_{\text{cell}} = +1.07 \text{ V} + (-0.76 \text{ V}) = 1.83 \text{ V}$$

Example 2

In basic solution, selenide and sulfite ions react spontaneously:



Write the balanced half-reactions for the process; if E°_{sulfite} is -0.57 V , calculate $E^\circ_{\text{selenium}}$. First, break the overall reaction into balanced half-reactions.



Now we need to find the half-cell potential for the oxidation half-reaction. We are given the cell potential and the half-cell potential for the reduction half-reaction.

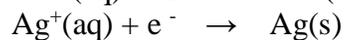
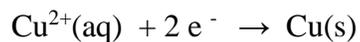
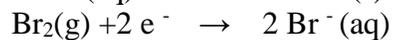
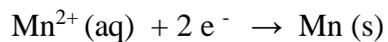
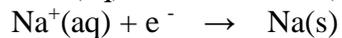
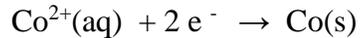
REMEMBER: we never multiply an E° by a coefficient!

$$E^\circ_{\text{cell}} = E^\circ_{\text{sulfite}} + E^\circ_{\text{selenium}}$$

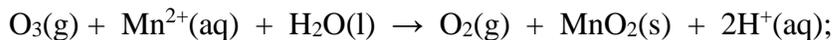
$$0.35 \text{ V} = (-0.57 \text{ V}) + E^\circ_{\text{selenium}}$$

$$E^\circ_{\text{selenium}} = 0.92 \text{ V}$$

7. Write the balanced thermodynamically favored reactions, and calculate the E°



8. In acidic solution, ozone and manganese(II) ion react spontaneously:



$$E^\circ_{\text{cell}} = 0.84 \text{ V}$$

(a) Write the balanced half-reactions.

(b) The half cell reduction potential for $\text{Mn}^{2+}/\text{MnO}_2$ is 1.21 V, determine the reduction potential for O_3/O_2

Under STANDARD equilibrium conditions, there are a couple of equations that you can use:

(a) $\Delta G^\circ = -nFE^\circ$ where "n" is moles of electrons transferred and "F" is Faraday's constant = 96,485 Coulombs/mol e⁻

(b) $\Delta G^\circ = -RT \ln K$ "R" is 8.314 J/mol rxn K

(c) $nFE^\circ = RT \ln K$ Which is just combining the two to relate E^o to K

Also, at equilibrium, $E_{\text{cell}} = 0$ and $\Delta G = 0$. The cell will no longer produce a current.

If conditions are NONSTANDARD then use the Nernst equation:

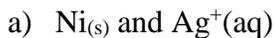
(d) $E_{\text{cell}} = E^\circ_{\text{cell}} - \frac{RT}{nF} \ln Q$ Nernst equation is off the test, but a qualitative understanding of the values is still on the test

Remember $Q = \frac{[\text{products}]}{[\text{reactants}]}$, therefore if there are a lot of reactants and a small amount of product (beginning of a reaction) Q will be less than 1. As the reaction proceeds to the right Q will become larger than 1 once there are more products than reactants. Using the following derivation of the Nernst equation

$E^\circ - E = \frac{RT}{nF} \ln Q$ If Q is less than 1, ln Q is negative and $E > E^\circ$
If Q is greater than 1, ln Q is positive and $E^\circ > E$

Remember batteries are only at standard state for a second while running. This means when Q is less than 1, like if an electrochemical reaction is just starting, then our E should be higher than E^o. As the reaction proceeds, the Q will increase and decrease the electric potential until it is eventually less than the standard state value when there are more products than reactants.

9. What is the value of the equilibrium constant at 25 °C for the reaction between:



The last topic is electrolysis. This process is nonspontaneous. We can, however, make the reaction work if an external power source is supplied. The main thing you need to be able to do here is the stoichiometry involved with electrolysis. You should be methodical, writing out every unit possible to get the desired result. The following are useful conversion factors that you should know:

- a) 1 Ampere = 1 Coulomb/ second
- b) 1 Volt = 1 Joule/Coulomb
- c) 1 F = 96485 Coulombs/mol e⁻

Some other helpful hints:

- a) You will need the # mol e⁻ transferred in the balanced redox reaction.
- b) You may be asked to solve for the molar mass of an unknown metal. Molar mass = gram / mole

11. How many grams of potassium can be deposited by the passage of 315 C through an electrolytic cell?

12. How long in seconds does it take to deposit 15.5 g Zn on a steel gate when 25.0 A are passed through a ZnSO₄ solution?

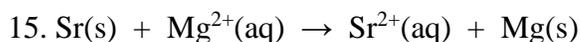
13. Electrolysis of molten metal (MCl₃) using a current of 6.50 A for 1397 seconds deposits 1.41 g of the metal at the cathode. What is the metal? (Hint: based on the chemical formula, the metal has a 3+ charge, inferring that 3 electrons were transferred.)

14. In an electrolytic cell, a current of 0.250 ampere is passed through a solution of a chloride of iron, producing Fe(s) and Cl₂(g).

- a) Write the equation for the reaction that occurs at the anode.

- b) When the cell operates for 2.00 hours, 0.521 grams of iron is deposited at one electrode. Determine the formula of the chloride of iron in the original solution.

c) Write the balanced equation for the overall reaction that occurs in the cell.



Consider the reaction represented above that occurs at 25 °C. All reactants and products are in their standard states. The value of the equilibrium constant, K_{eq} , for the reaction is 4.2×10^{17} at 25 °C.

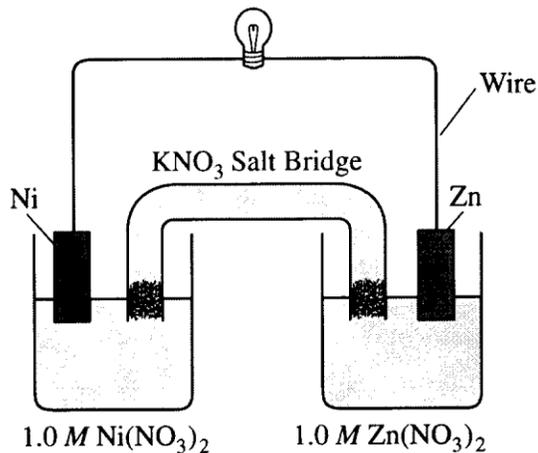
- (a) Predict the sign of the standard cell potential, E°_{cell} , for a cell based on the reaction. Explain your prediction. Do not look at a table to get your answer.

- (b) Identify the oxidizing agent for the spontaneous reaction.

- (c) How would the cell potential change if the reaction were carried out at 25 °C with a 1.0 M solution Mg(NO₃)₂ and 0.10 M solution of Sr(NO₃)₂? Explain.

- (d) When the cell reaction in (d) reaches equilibrium, what is the cell potential?

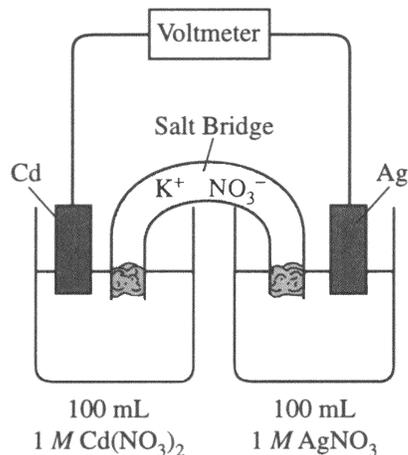
16.



Answer the following questions that refer to the galvanic cell shown in the diagram above.

- Identify the anode of the cell and write the half-reaction that occurs there.
- Write the net ionic equation for the overall reaction that occurs as the cell operates and calculate the value of the standard cell potential, E°_{cell} .
- Indicate how the value of E_{cell} would be affected if the concentration of $\text{Ni}(\text{NO}_3)_2(\text{aq})$ was changed from 1.0 M to 0.10 M and the concentration of $\text{Zn}(\text{NO}_3)_2(\text{aq})$ remained 1.0 M. Justify your answer.
- Specify whether the value of K_{eq} for the cell reaction is less than 1, greater than 1, or equal to 1. Justify your answer.

17.



Answer the following questions regarding the electrochemical cell shown above.

(a) Write the balanced net-ionic equation for the spontaneous reaction that occurs as the cell operates, and determine the cell voltage.

(b) In which direction do anions flow in the salt bridge as the cell operates. Justify your answer.

(c) If 10.0 mL of 3.0 M AgNO₃ solution is added to the half-cell on the right, what will happen to the cell voltage? Explain.

(d) If 1.0 grams of solid NaCl is added to each half-cell, what will happen to the cell voltage? Explain.

- (e) If 20.0 mL of distilled water is added to both half-cells, the cell voltage decreases. Explain.

18. Answer the following questions that relate to electrochemical reactions.

- (a) Under standard conditions at 25 °C, Zn(s) reacts with Co²⁺(aq) to produce Co(s).

(i) Write the balanced equation for the oxidation half-reaction.

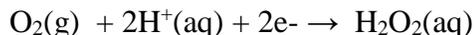
(ii) Write the balanced net-ionic equation for the overall reaction.

(iii) Calculate the standard cell potential, E° , for the reaction at 25 °C.

(b) At 25 °C, H₂O₂ decomposes according to the following equation.



- (i) Determine the value of the standard free energy change, ΔG° , for the reaction at 25 °C.
- (ii) Determine the value of the equilibrium constant, K_{eq} , for the reaction at 25 °C.
- (iii) The standard reduction potential, E° , for the half-reaction $\text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^- \rightarrow 2\text{H}_2\text{O}(\text{l})$ has the value of 1.23 V. Using the information in addition to the information given above, determine the value of the standard reduction potential, E° , in the half-reaction below.



(c) In an electrochemical cell, Cu(s) is produced by the electrolysis of CuSO₄ (aq). Calculate the maximum mass of Cu(s) that can be deposited by a direct current of 100. amperes passed through 5.00 L of 2.00 M CuSO₄(aq) for a period of 1.00 hour.