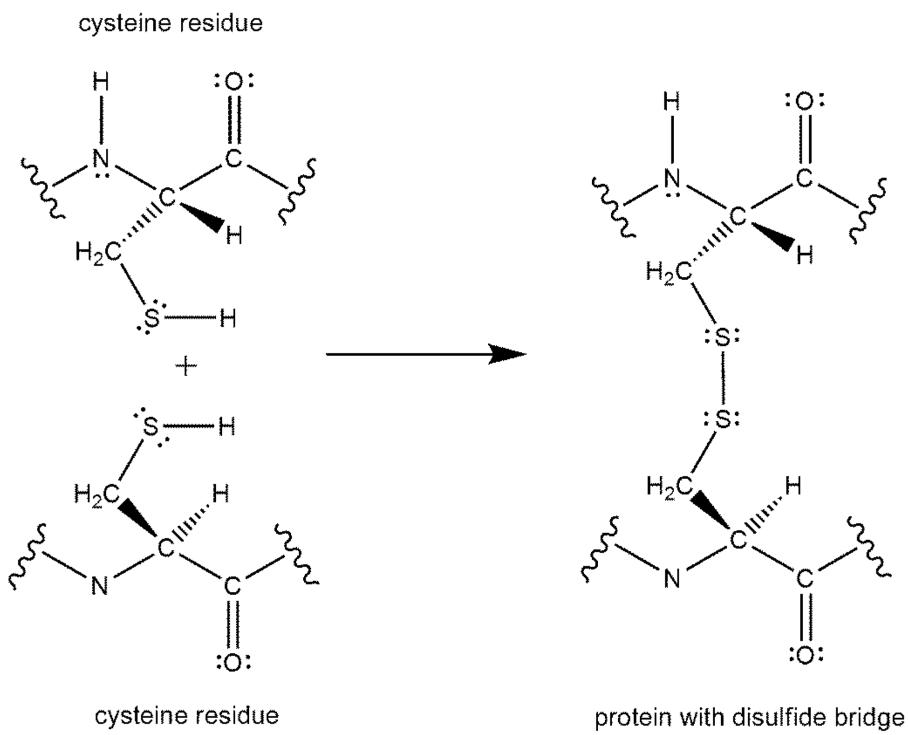
Practice Test Questions 9
Redox Reactions and Electrochemistry

1. Proteins are long chains of amino acids that fold up to make three-dimensional structures. One of the ways in which the chains are held in those three-dimensional shapes is through "disulfide bridges" between cysteine amino acid residues:



(a) What is the oxidation state of the sulfur atom in a cysteine residue?

(b) What is the oxidation state of the sulfur atom that is part of a disulfide bridge?

(c) Circle the statement which best describes the reaction above:

i. The sulfur atom in the cysteine residue is reduced. For this to occur, it must react with a reducing agent.

ii. The sulfur atom in the cysteine residue is reduced. For this to occur, it must react with an oxidizing agent.

iii. The sulfur atom in the cysteine residue is oxidized. For this to occur, it must react with a reducing agent.

iv. The sulfur atom in the cysteine residue is oxidized. For this to occur, it must react with an oxidizing agent.

v. The sulfur atom in the cysteine residue is neither oxidized nor reduced. As such, no oxidizing or reducing agent is necessary.

2. Balance the following redox reaction that occurs in acidic solution. Label each half reaction as oxidation or reduction.

$$Cr_2O_{7(aq)}^{2-} + Ag_{(s)} \rightarrow Cr_{(aq)}^{3+} + Ag_{(aq)}^{+}$$

3. Carbon dioxide reacts with calcium metal in basic solution to produce the oxalate ion ($C_2O_4^{2-}$) and calcium ions.

(a) Write a balanced chemical equation for this reaction. *Show both half-reactions.*

(b) Identify the reducing agent in this reaction.

4. Consider the following redox reaction:

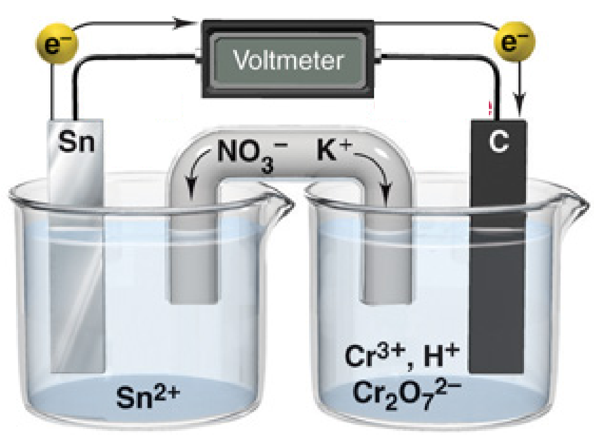
$$H_5IO_{6(aq)} + Rh_{(s)} \rightarrow IO_{3(aq)}^{-} + Rh_2O_{3(s)} \qquad E° = +0.73 \text{ V}$$

(a) Balance this equation under acidic conditions. *You must include balanced half reactions.*

(b) Calculate the standard free energy change for this reaction.

5. An initially pH-neutral solution of cysteine ($C_3H_7NO_2S$) reacts with solid HgO to form cystine ($C_6H_{12}N_2O_4S_2$) and metallic mercury. Write a balanced chemical equation for this reaction, labeling the oxidation and reduction half-reactions.

6. Consider the following electrochemical cell, shown below:



(a) Describe this cell using abbreviated standard notation (include phases).

(b) Provide the balanced oxidation and reduction half reactions.

(c) Provide the balanced overall redox reaction (*in acidic solution*).

7. Consider the electrochemical cell shown below under standard conditions:

$$Pt_{(s)}\left|NO_{(g)}\right|OH^-_{(aq)}\left|NO^-_{3\,(aq)}\right\|Cl_{2(g)}\left|Cl^-_{(aq)}\right|Pt_{(s)}$$

(a) Provide the balanced half reactions and overall balanced reaction for this cell.

(b) Calculate the potential for this cell under *standard* conditions.

(c) Is this cell thermodynamically allowed (spontaneous) as written?

(d) Will this electrochemical cell be thermodynamically allowed under the following conditions:

- pressure of $NO_{(g)}$ is 0.2 bar
- pressure of $Cl_{2\,(g)}$ is 0.05 bar
- concentration of $NO^-_{3\,(aq)}$ is 0.01 M
- concentration of $Cl^-_{(aq)}$ is 0.8 M
- concentration of $OH^-_{(aq)}$ is 0.03 M

8. Consider the following cell:

$$Pt_{(s)}\left|H_{2(g)}\right|H^+_{(aq)}\left\|Br^-_{(aq)}\right|AgBr_{(s)},Ag_{(s)}$$

(a) Sketch this cell, labeling all essential components.

(b) Write the overall reaction for this cell.

9. A fuel cell operates on propane supplied at a pressure of 1 bar at 25 ℃. Air is used as the source of oxygen, so the partial pressure of O_2 in the fuel cell is 0.2 bar. While the fuel cell is operating, the partial pressure of carbon dioxide in the cell is approximately 0.05 bar and the water produced is liquid. The two electrodes are separated by a concentrated sodium hydroxide solution.

(a) Write balanced half-reactions for the two electrodes. Label which half-reaction corresponds to the anode and which corresponds to the cathode.

(b) Calculate the potential of this fuel cell.

10. The hydrazine fuel cell is based on the reaction:

$$N_2H_{4(aq)} + O_{2(g)} \rightarrow N_{2(g)} + 2\,H_2O_{(l)}$$

For this cell, $E°_{cell} = +1.559\,V$.

(a) Is this reaction *spontaneous* in the thermodynamic sense? Explain.

(b) Calculate the standard free energy of formation for $N_2H_{4(aq)}$.

11. Use the half-reaction method to balance the following redox equation in 1 M acid solution.

$$C_3H_7CH_2OH_{(aq)} + Cr_2O^{2-}_{7(aq)} \rightarrow C_3H_7CO_2H_{(aq)} + Cr^{3+}_{(aq)}$$

12. The following voltaic cell operates at 1 bar pressure and 298.15 K:

$$Sn_{(s)}|Sn^{2+}_{(aq)}(0.075\ M)\|Pb^{2+}_{(aq)}(0.600\ M)|Pb_{(s)}$$

(a) Write a balanced redox reaction

(b) What is E_{cell} initially?

(c) If the cell is allowed to operate spontaneously, will E_{cell} increase, decrease, or remain constant with time? Explain.

13. In a fuel cell, the hydrogen reacts with oxygen as described by the following chemical equation:

$$2\ H_{2(g)}\ +\ O_{2(g)}\ \rightarrow\ 2\ H_2O_{(l)}$$

(a) Calculate the cell potential for this reaction under standard conditions.
Hint: H_2 is oxidized to H^+.

(b) A fuel cell would not actually operate under standard conditions. Calculate the cell potential for a fuel cell in which the partial pressure of the hydrogen gas is 75 kPa and the partial pressure of the oxygen gas is 15 kPa? The temperature is still 25 °C.

14. Gold is a highly prized metal which is so chemically inert that it won't react, even with the strongest concentrated acids. But it can be dissolved in *aqua regia* (recall: you made aqua regia solutions in qualitative analysis labs this semester).
The unbalanced equation for the reaction of gold with aqua regia is given below:

$$\_\_Au_{(s)} + \_\_HNO_{3\,(aq)} + \_\_HCl_{(aq)} \rightarrow \_\_HAuCl_{4\,(aq)} + \_\_NO_{2\,(g)} + \_\_H_2O_{(l)}$$

(a) Draw the Lewis structure of HNO_3 including any possible resonance structures. Assign the correct oxidation states for N, O, and H.

(b) What element is being oxidized?

(c) What element is being reduced?

(d) Provide the two balanced half-reactions for the two electrodes. Label which half-reaction corresponds to the anode and which corresponds to the cathode.

(e) Balance the reaction of gold with aqua regia by adding the proper stoichiometric coefficients.

15. An electrochemical cell is described below:

$$Pt_{(s)}|Mn^{2+}_{(aq)}\ (0.0320\ M), MnO_{4\,(aq)}^{-}\ (0.010\ M), H^{+}_{(aq)}(pH = 2.1)\|F^{-}_{(aq)}\ (0.0070\ M), F_{2\,(g)}\ (0.110\ bar)|Pt_{(s)}$$

(a) Write the overall balanced equation for this reaction.

(b) Calculate the potential for this cell at 25.00 °C.

16. Balance the following redox reaction that takes place in acidic solution.

$$Sn^{2+}_{(aq)}\ +\ IO^{-}_{4(aq)}\ \rightarrow\ Sn^{4+}_{(aq)}\ +\ I^{-}_{(aq)}$$

17. The cell described below develops a potential of -0.97 V at 25 °C.

Assume exact temperature so that temperature has infinite sig. fig.

$$Pt_{(s)}\left|H_{2(g)}(1\ bar)\right|H^{+}_{(aq)}(pH = 5.0)\|V^{2+}_{(aq)}(0.0010\ M)\left|V_{(s)}\right.$$

(a) What is the standard reduction potential of the vanadium(II) ion?

(b) What is the standard free energy of formation of the aqueous vanadium(II) ion?

18. Calculate the equilibrium constant for the reaction below at 25 °C *(exact temperature).*

$$Hg^{2+}_{2(aq)} \rightleftharpoons Hg_{(l)} + Hg^{2+}_{(aq)}$$

Which of the two aqueous ions will be more abundant at equilibrium?

Metallic mercury is insoluble in water, so it precipitates out if formed. If necessary, assume that the aqueous layer is in direct contact with metallic mercury initially.

19. Mercury(II) sulfide is only sparingly soluble in water. The sulfide ion is a stronger base than hydroxide, so the solubility equilibrium is:

$$HgS_{(s)} + H_2O_{(l)} \rightleftharpoons Hg^{2+}_{(aq)} + HS^{-}_{(aq)} + OH^{-}_{(aq)}$$

The equilibrium constant for this reaction is 2×10^{-53} at 25 °C *(exact temperature).*

(a) If we put some solid HgS in 1.0 L of water, how many Hg^{2+} ions are present in solution at equilibrium? Express your answer as an actual number of ions, i.e. not in moles.

Hint: Very little of the HgS will dissolve, so there will be a negligible effect on the pH.

(b) Can you imagine electrochemical measurements that would allow us to calculate this equilibrium constant? *"Yes" and "no" are not acceptable answers to this question! ☺ Please elaborate…*

20. Consider the following electrochemical cell:

$$Cd_{(s)}, Cd(OH)_{2(s)}\left|OH^{-}_{(aq)}(0.023\ M)\right\|Zn^{2+}_{(aq)}(0.11\ M)\left|Zn_{(s)}\right.$$

(a) Write a balanced chemical equation for the reaction in this electrochemical cell.

(b) Calculate the potential generated by this cell at 25 °C.

(c) Is the reaction, as written, thermodynamically allowed? Why or why not?

21. For the following cell reaction:

$$2\ Fe^{3+}_{(aq)} + 2\ I^{-}_{(aq)} \rightarrow 2\ Fe^{2+}_{(aq)} + I_{2(s)}$$

(a) Express the cell using standard cell notation

(b) Calculate the half cell reduction potential of I_2.