Unit 4 Review 1

INTRODUCTION TO REDOX REACTIONS AND STANDARD CELLS

Unit 4 Outline: Electrochemistry

1. Understand fundamental redox reactions
   - Assigning oxidation numbers
   - Balancing redox reactions
   - Identifying the roles of the different species in the reaction

} 20 -25% exam

2. Understand and apply the fundamentals of standard cells (voltaic and electrolytic)
3. Apply the concepts of electrochemical cells to non-standard conditions
   - Concentration Cells
   - Nernst Potential
4. Complete the storylines of thermodynamics and equilibrium by converting electrical potential into *K* and ΔG.
5. Common applications of batteries
   - Primary and secondary cells
   - Fuel Cells

OIL RIG, LEO GER

Electrochemistry Definitions

1. **Redox Reaction**: a chemical reaction that involves the transfer of electrons from one species to another, resulting in a change in oxidation state. A redox reaction balances both electron-transfer AND stoichiometric coefficients

2. **Reduction**: a species gains electrons in a half-reaction, resulting in a lower oxidation state

3. **Oxidation**: a species loses electrons in a half-reaction, resulting in a higher oxidation state

4. **Oxidizing Agent:** the species that drives the oxidation of another species in a redox reaction **; the oxidizing agent is always the species undergoing reduction as a reactant**

5. **Reducing Agent:** the species that drives the reduction of another species in a redox reaction **; the reducing agent is always the species undergoing oxidation as a reactant**

$(Cu^{2+} + 2e^- \rightarrow Cu) \times 3$

$(Al \rightarrow Al^{3+} + 3e^-) \times 2$

$3Cu^{2+} + 6e^- + 2Al \rightarrow 3Cu + 2Al^{3+} + 6e^-$

✓ $3Cu^{2+} + 2Al \rightarrow 3Cu + 2Al^{3+}$

Cu^{2+}: oxidizing agent

Al: reducing agent

Electrochemistry Definitions- Redox Reaction

$Fe^{3+} + 3e^- \rightarrow Fe$

$Pb^{2+} + 2e^- \rightarrow Pb$

What is the balanced redox reaction if Pb is the reducing agent & Fe^{3+} is the oxidizing agent?

$3\times$ $Pb \rightarrow Pb^{2+} + 2e^-$ $\rightarrow$ $3Pb \rightarrow 3Pb^{2+} + \cancel{6e^-}$

$2\times$ $Fe^{3+} + 3e^- \rightarrow Fe$ $\rightarrow$ $2Fe^{3+} + \cancel{6e^-} \rightarrow 2Fe$

$3Pb + 2Fe^{3+} \rightarrow 3Pb^{2+} + 2Fe$ ✓

Oxidation Numbers

1. Atoms in their standard state are neutral
   - Example: O_2, Na, Fe, Br_2
2. If the question directly assigns an oxidation number, that's its oxidation number
   - Example: iron(III), aluminum(II)

   Fe^{3+}
3. Group 1 is +1, Group 2 is +2…Group 7A is -1
   - Example: Na^+, K^+, F^-
4. Hydrogen is +1, Oxygen is -2 when bonded
   - Exceptions: hydrides are -1, peroxides are -1
5. Assign peripheral charges first

$0 = 2 + 2(6) - 14$

$2 \times (+1)$ $\quad$ $(-2) \times 7$

$K_2Cr_2O_7$

6

$0 = 3 + (-1) \times 3$

IF_3

Exam Question: Change in Oxidation

In the redox conversion of SO_3 to SO^-, S is ___?___ and its oxidation number goes from ___6___ to ___1___

$$SO_3 \rightarrow SO^-$$

$$\emptyset = 6 + -6 \qquad -1 = 1 + -2$$

reduction

Exam Question: Change in Oxidation

For the conversion of $Na_2Cr_2O_7$ to $Cr(OH)_3$, the Cr atom gets __?__ and the change in oxidation number is equal to __?__.

1. oxidized, +6
2. oxidized, +3
3. reduced, -6
4. reduced, +9
5. oxidized, -3
6. reduced, -3

$Na_2Cr_2O_7$ → 6

$6 \rightarrow 3$

$Cr^{+3}(OH)_3$

$(-1) \times 3$

Δ oxidation # = -3

reduction

Warm-Up Question

We will look in depth at balancing these three REDOX reactions. First, can you look at each reaction and determine:

a. What is being oxidized?
b. What is being reduced?
c. What is the oxidizing agent?
d. What is the reducing agent?

$$NbO_2 + W \rightarrow Nb + WO_4^{2-}$$

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$$

$$Pu + NpO_3^+ \rightarrow PuO_2 + NpO_2^{2+}$$

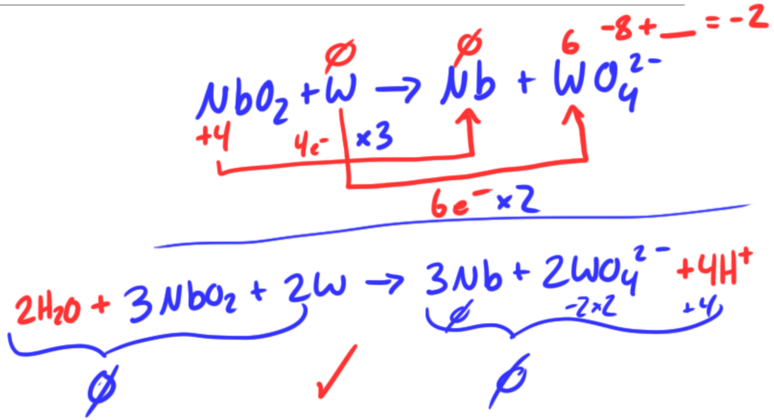
Balanced Reaction in Acid – Balance by Inspection

Balance the following redox reaction in acidic solution. You will have to provide the H_2O and the H^+ for the reaction. Make sure all the coefficients are whole numbers.

$$NbO_2 + W \rightarrow Nb + WO_4^{2-}$$

What is the coefficient for WO_4^{2-} in the balanced equation?

Bonus: Identify the reducing and oxidizing agents



Note: the full, cleaner solution is shown in the last slide

Balanced Reaction in Base

Using the set of smallest whole number coefficients to balance the redox equation

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$$

in basic solution, you get

Bonus: Identify the reducing and oxidizing agents

$MnO_4^- \rightarrow MnO_2$

1) H_2O

$MnO_4^- \rightarrow MnO_2 + 2H_2O$

2) H^+

$4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$

$4OH^-$ / $4OH^-$

$4H_2O + MnO_4^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

-1 / -4

$4H_2O + MnO_4^- + 3e^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

$NO_2^- \rightarrow NO_3^-$

$H_2O + NO_2^- \rightarrow NO_3^-$

$H_2O + NO_2^- \rightarrow NO_3^- + 2H^+$

$2OH^-$ / $2OH^-$

$2OH^- + H_2O + NO_2^- \rightarrow NO_3^- + 2H_2O$

-3 / -1

$2OH^- + H_2O + NO_2^- \rightarrow NO_3^- + 2H_2O + 2e^-$

4) e^-

$4H_2O + MnO_4^- + 3e^- \rightarrow MnO_2 + 2H_2O + 4OH^-$

$2OH^- + H_2O + NO_2^- \rightarrow NO_3^- + 2H_2O + 2e^-$

5) Merge

Balance a Reaction: You Choose the Method

Consider the following rather planetary and radioactive redox reaction:

$$Pu + NpO_3^{+} \rightarrow PuO_2 + NpO_2^{2+}$$

Use the smallest possible integer coefficients to balance the reaction in acidic aqueous solution. What is the coefficient for H_2O and which side of the equation is it on?

Bonus: Identify the reducing and oxidizing agents

$Pu \rightarrow PuO_2$

$NpO_3^{+} \rightarrow NpO_2^{2+}$

$2H_2O + Pu \rightarrow PuO_2 + 4H^+ + 4e^-$ (+4)

$(2H^+ + NpO_3^{+} + 1e^- \rightarrow NpO_2^{2+} + H_2O) \times 4$

$8H^+ + 4NpO_3^{+} + 4e^- \rightarrow 4NpO_2^{2+} + 4H_2O$

$2H_2O, 8H^+$

$4H_2O, 4H^+$

$2H_2O + Pu \rightarrow PuO_2 + 4H^+ + 4e^-$

$8H^+ + 4NpO_3^{+} + 4e^- \rightarrow 4NpO_2^{2+} + 4H_2O$

$4H^+ + Pu + 4NpO_3^{+} \rightarrow PuO_2 + 4NpO_2^{2+} + 2H_2O$ ✓

Using the set of smallest whole number coefficients to balance the redox equation

$$MnO_4^- + NO_2^- \rightarrow MnO_2 + NO_3^-$$

in basic solution, you get

$MnO_4^- \rightarrow MnO_2$ $\quad NO_2^- \rightarrow NO_3^-$

1) H_2O

$MnO_4^- \rightarrow MnO_2 + 2H_2O$ $\quad H_2O + NO_2^- \rightarrow NO_3^-$

2) H^+

$4H^+ + MnO_4^- \rightarrow MnO_2 + 2H_2O$ $\quad H_2O + NO_2^- \rightarrow NO_3^- + 2H^+$

3) OH^- ($4OH^-$; $2OH^-$)

×2 $(4H_2O + MnO_4^- + 3e^- \rightarrow MnO_2 + 2H_2O + 4OH^-)$ $\quad (H_2O + NO_2^- + 2OH^- \rightarrow NO_3^- + 2H_2O + 2e^-)$ ×3

4) e^-

5) balanced between

$8H_2O + 2MnO_4^- + 6e^- \rightarrow 2MnO_2 + 4H_2O + 8OH^-$

$3H_2O + 3NO_2^- + 6OH^- \rightarrow 3NO_3^- + 6H_2O + 6e^-$

$H_2O + 2MnO_4^- + 3NO_2^- \rightarrow 2MnO_2 + 3NO_3^- + 2OH^-$

-5 $\quad -5$ ✓