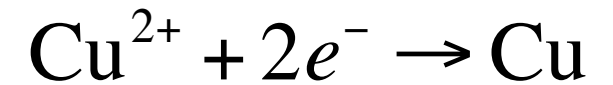
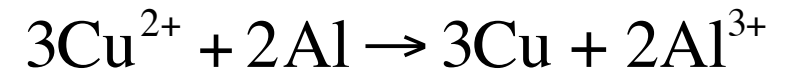


CH 302 – Unit 4 Review 1

INTRODUCTION TO ELECTROCHEMISTRY, THE STANDARD CELL

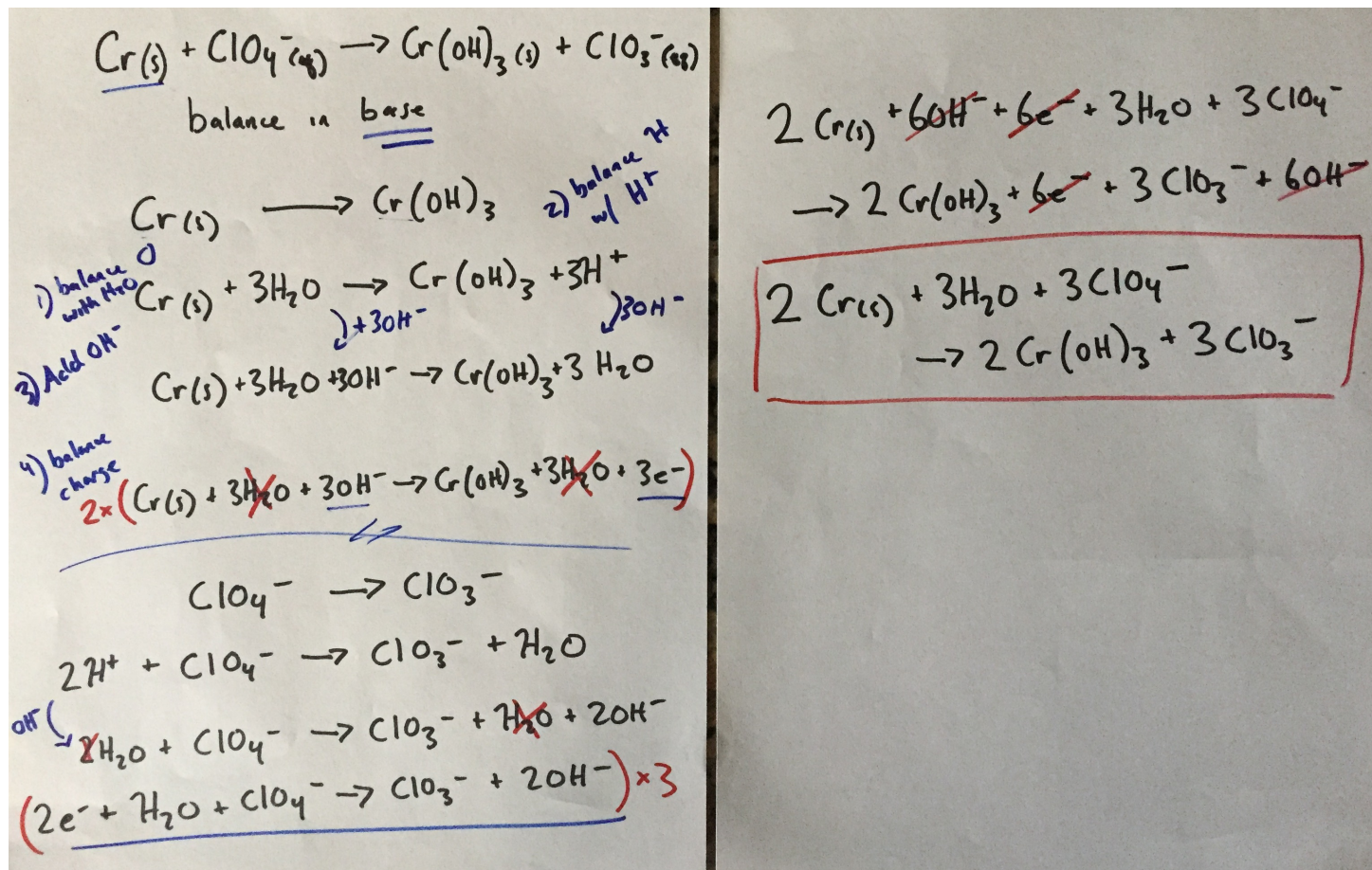
Electrochemistry Definitions

- 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 involves one species undergoing reduction and another undergoing oxidation.
- 2. Reduction:** a species gains electrons, resulting in a lower oxidation state
- 3. Oxidation:** a species loses electrons, 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
- 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



LEO says GER
OIL RIG

Balanced Reaction in Base



Electrochemical Potential

- **Electrical Cell Potential ($\epsilon^\circ_{\text{cell}}$):** the voltage associated with the redox reaction occurring in an electrochemical cell

$$\epsilon^\circ_{\text{cell}} = \epsilon^\circ_{\text{cathode}} - \epsilon^\circ_{\text{anode}}$$

in this equation, both ϵ° values are reduction potentials read from a table

- This can also be written in the following way:

$$\epsilon^\circ_{\text{cell}} = \epsilon^\circ_{\text{reduction}} + \epsilon^\circ_{\text{oxidation}}$$

**However, this equation uses the reduction potential plus the oxidation potential.

This means that you will not be able to directly use the data from a reduction potential chart.**

Question

Which of the following species will reduce Cu^{2+} but not Pb^{2+} ?

- a. Fe^{2+}
- b. H_2
- c. Zn
- d. Fe^{3+}
- e. H^+

Cl_2	+	$2e^-$	\rightleftharpoons	2Cl^-	+1.36 V
Ag^+	+	$1e^-$	\rightleftharpoons	Ag	+0.80 V
Fe^{3+}	+	$1e^-$	\rightleftharpoons	Fe^{2+}	+0.77 V
Cu^{2+}	+	$2e^-$	\rightleftharpoons	Cu	+0.34 V
2H^+	+	$2e^-$	\rightleftharpoons	H_2	0.00 V
Fe^{3+}	+	$3e^-$	\rightleftharpoons	Fe	-0.04 V
Pb^{2+}	+	$2e^-$	\rightleftharpoons	Pb	-0.13 V
Fe^{2+}	+	$2e^-$	\rightleftharpoons	Fe	-0.44 V
Zn^{2+}	+	$2e^-$	\rightleftharpoons	Zn	-0.76 V
Al^{3+}	+	$3e^-$	\rightleftharpoons	Al	-1.66 V

increasing strength as an reducing agent

Question

Which of the following species will reduce Cu^{2+} but not Pb^{2+} ?

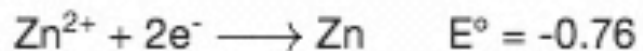
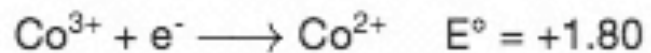
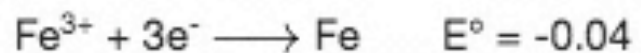
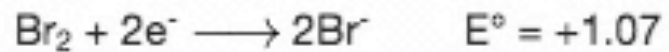
- a. Fe^{2+}
- b. H_2
- c. Zn
- d. Fe^{3+}
- e. H^+

Cl_2	+	2e^-	\rightleftharpoons	2Cl^-	+1.36 V
Ag^+	+	1e^-	\rightleftharpoons	Ag	+0.80 V
Fe^{3+}	+	1e^-	\rightleftharpoons	Fe^{2+}	+0.77 V
Cu^{2+}	+	2e^-	\rightleftharpoons	Cu	+0.34 V
2H^+	+	2e^-	\rightleftharpoons	H_2	0.00 V
Fe^{3+}	+	3e^-	\rightleftharpoons	Fe	-0.04 V
Pb^{2+}	+	2e^-	\rightleftharpoons	Pb	-0.13 V
Fe^{2+}	+	2e^-	\rightleftharpoons	Fe	-0.44 V
Zn^{2+}	+	2e^-	\rightleftharpoons	Zn	-0.76 V
Al^{3+}	+	3e^-	\rightleftharpoons	Al	-1.66 V

increasing strength as an reducing agent

Question

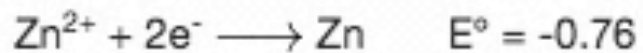
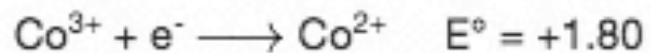
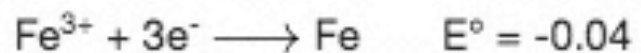
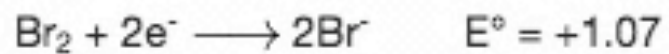
What is the standard cell potential of the strongest battery that could be made using these half-reactions?



Hint: look for the strongest reducing agent and the strongest oxidizing agent

Question

What is the standard cell potential of the strongest battery that could be made using these half-reactions?



Answer:



Electrochemical Cell Definitions

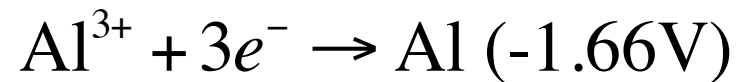
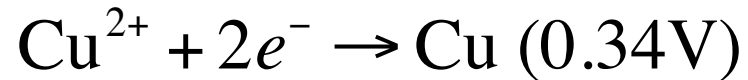
1. **Cathode**: the site of reduction (Red Cat) ; takes in electrons from the wire
2. **Anode**: the site of oxidation (An Ox) ; gives electrons to the wire
3. **Voltaic Cell (Galvanic Cell)**: an electrochemical cell with a positive standard cell potential ; the redox reaction of the cell occurs without an external power source (spontaneous)
4. **Electrolytic Cell**: an electrochemical cell with a negative standard cell potential ; the redox reaction of the cell relies on an external power source (non-spontaneous)
5. Shorthand Notation:

anode | anodic solution || cathodic solution | cathode

inert metal | anode reactant | anode product || cathode reactant | cathode product | inert metal

The Electrochemical Cell

Using the following half-reactions, show the voltaic and electrolytic cells that can be created:



$$\epsilon^{\circ}_{\text{cell}} = \epsilon^{\circ}_{\text{cathode}} - \epsilon^{\circ}_{\text{anode}}$$

Galvanic Voltage: **2.00V** = (0.34V) - (-1.66V) ; **Al | Al³⁺ || Cu²⁺ | Cu**

Electrolytic Voltage: **-2.00V** = (-1.66V) - (0.34V) ; **Cu | Cu²⁺ || Al³⁺ | Al**

The Electrochemical Cell Summary

	voltaic cells	electrolytic cells
free energy, ΔG	negative (-)	positive (+)
potential, E	positive (+)	negative (-)
push/pull of electrons	from the chemical reactions of the two half-reactions	from an external electrical power source
anode	negative (-)	positive (+)
cathode	positive (+)	negative (-)

In all electrochemical cells, the electrons travel from the site of oxidation (anode) to the site of reduction (cathode). The main difference is that voltaic cells are spontaneous cells, **where the redox reaction drives the current**. In an electrolytic cell, the redox reaction is non-spontaneous. **Therefore, the push/pull of current is driven by an external power source.**