Chapter 9: Introduction of electrochemistry

You will be able to:

1 - Determine oxidation number
2 - Identify reduction and oxidation reaction
3 - Predict spontaneous reaction

Introduction of electro-chemistry

electrochemistry = branch of chemistry concerned with the conversion of chemical energy to electrical energy (& vice versa)

- electrochemical reactions involve the **transfer of electrons** from one substance to another
- consider the reaction:

$$2Al(s) + 3CuCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Cu(s)$$

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

- electrons are transferred from Al to Cu^{2+}

- we can re-write this equation as two separate half-reactions

oxidation half-reaction: $Al \rightarrow Al^{3+} + 3e^-$

reduction half-reaction: $Cu^{2+} + 2e^- \rightarrow Cu$

loss of electrons = OXIDATION
gain of electrons = REDUCTION

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DGRO

L E O the lion says G E R!

Lose Electrons Oxidation

Gain Electrons Reduction

Oxidation Is Losing e^- Reduction Is Gaining e^-

= OIL RIG

- when a substance becomes **oxidized** it **becomes more positively charged** because it is losing electrons (which are negatively charged)

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

$$U^{3+} \rightarrow U^{4+} + e-$$

$$2Cl^{-} \rightarrow Cl_2 + 2e-$$

- when a substance becomes **reduced** it **becomes more negatively charged** because it is gaining electrons

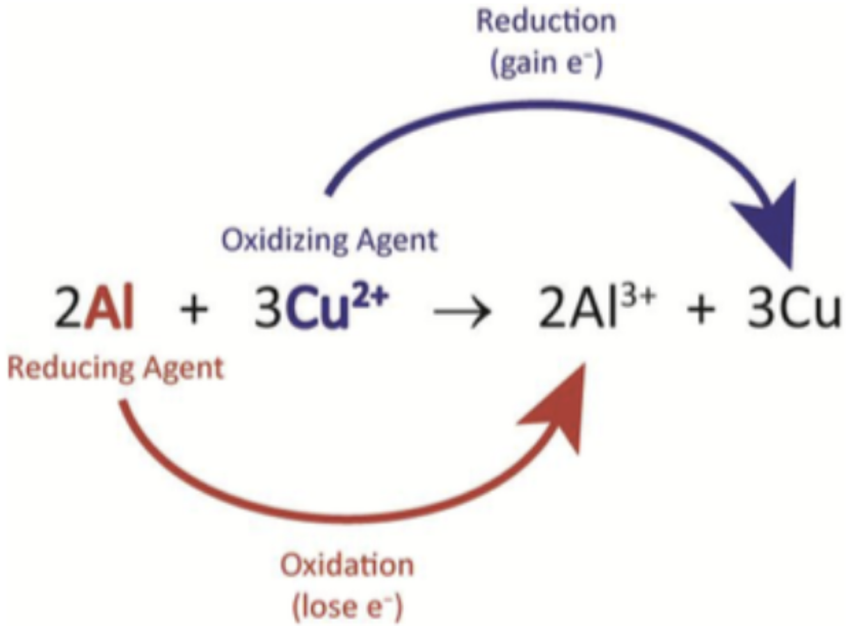
$$Cu^{2+}\ 2e- \rightarrow Cu$$

$$V^{3+} + e- \rightarrow V^{2+}$$

$$F_2 + 2e- \rightarrow 2F^{-}$$

- every reduction reaction must be accompanied by an oxidation reaction since the electrons must be transferred somewhere

  > these are called **REDOX reactions**



The **oxidizing agent** is the reactant reduced (gains e-) during a reaction.

The **reducing agent** is the reactant oxidized (loses e-) during a reaction.

> Al is the **reducing agent** because it causes Cu^{2+} to become reduced
> Cu^{2+} is the **oxidizing agent** because it causes Al to become oxidized

Oxidation Number

- oxidation numbers can be used to determine whether an atom has been oxidized or reduced

oxidation number = real or apparent charge an atom or ion has when all of the bonds are assumed to be ionic

Determining oxidation number:

1. atoms in elemental form = **0**
2. simple ions = **the charge on ion**
   1. Li^+, Na^+, K^+ and all other **group 1 ions** have an oxidation number of **1+**
   2. Ca^{2+}, Ba^{2+}, Mg^{2+} and all other **group 2 ions** have an oxidation number of **2+**
   3. F^-, Cl^-, Br^-, I^- (**halogens**) are usually **1-** but there are many exceptions, especially in covalent compounds
3. **hydrogen** = **+1** (except in metallic hydrides such as NaH or BaH_2 where it is 1-)
4. **oxygen** = **2-** (in peroxides, H_2O_2, it is 1-)
5. oxidation numbers of other atoms are assigned so that the **sum of the oxidation numbers** (positive & negative) **equals the net charge on the molecule or ion**

Determine the oxidation numbers of each atom for the following:

Cl_2 Cl: 0

CCl_4 C: 4+ ; Cl: 1− (4−)

K_2CrO_4 K: 1+ (2+) ; Cr: 6+ ; O: 2− (8−) ; 2−

N_2O_4 N: 4+ (8+) ; O: 2− (8−)

PO_4^{3-} P: 5+ ; O: 2− (8−)

C_3H_8 C: −8/3 ; H: 1+ (8+)

When an atom's oxidation number increases, it has become oxidized, and when an atom's oxidation number decreases, it has become reduced.

$\overset{5+}{Cl}\overset{2-(6-)}{O_3^-} \rightarrow \overset{7+}{Cl}\overset{2-(8-)}{O_4^-}$ Cl: 5+ → 7+ OX

$H_2\overset{1-}{O_2} \rightarrow H_2\overset{2-}{O}$ O: 1− → 2− red

$Cr^{3+} \rightarrow \overset{6+}{Cr}\overset{2-(8-)}{O_4^{2-}}$ Cr: 3+ → 6+ OX

$\overset{4+}{N}\overset{2-(4-)}{O_2} \rightarrow \overset{3+}{N_2}\overset{2-(6-)}{O_3}$ 4+ → 3+ Red

LEO

oxidation = loss of electrons = increase in oxidation number

reduction = gain of electrons = decrease in oxidation number

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Predicting Spontaneous Reaction

INCREASING TENDENCY TO REDUCE

INCREASING STRENGTH AS AN OXIDIZING AGENT

$$F_2 + 2e^- \rightleftarrows 2F^-$$

$$Ag^+ + e^- \rightleftarrows Ag$$

$$Cu^{2+} + 2e^- \rightleftarrows Cu$$

$$Zn^{2+} + 2e^- \rightleftarrows Zn$$

$$Li^+ + e^- \rightleftarrows Li$$

INCREASING TENDENCY TO OXIDIZE

INCREASING STRENGTH AS A REDUCING AGENT

oxidizing agents gain electrons & tend to be **cations** (+) or **non-metals**

reducing agents lose electrons & tend to be **anions** (-) or **metals**

- **stronger oxidizing agents** are located on the **upper left** and have a greater tendency to gain electrons (reduce)

- **stronger reducing agents** are located on the **lower right** and have a greater tendency to lose electrons (oxidize)

Consider the half-reaction for Zn and Zn^{2+}:

$$Zn^{2+} + 2e^- \rightleftarrows Zn$$

Zn^{2+} is an oxidizing agent and will gain electrons:

$$Zn^{2+} + 2e^- \rightarrow Zn \quad \text{(reduction)}$$

Zn is a reducing agent and will lose electrons:

$$Zn \rightarrow Zn^{2+} + 2e^- \quad \text{(oxidation)}$$

Spontaneous reactions will occur when there is:

an oxidizing agent (reduction) and a reducing agent (oxidation)

and

the oxidizing agent must be above the reducing agent in the table

ex. Which of the following metals - Al, Pb, Cu, Fe and Ag - can be oxidized by Cr^{3+}?

ex. Predict whether a spontaneous reaction is expected and the products that would be formed.

a) Pb^{2+} and MnO_2

b) $Cr_2O_7^{2-}$ and Sn^{2+}

ex. Consider the following spontaneous redox reactions:

$$X + Y \rightarrow X^{-} + Y^{+}$$
$$Y^{+} + Z \rightarrow Y + Z^{+}$$
$$Z + X \rightarrow Z^{+} + X^{-}$$

What is the relative strengths of oxidizing agents (strongest to weakest)?

ex. A solution containing Pd^{2+} reacts spontaneously with Ga to produce Pd and Ga^{3+}. However, a solution containing Pd^{2+} does not react with Pt. What are the reducing agents in order of increasing strengths?

Homework

- Start Career Planning Final Assignment
- **Read Read Read** Chapter 9.1 and 9.3
- Do page 607. #3, 4, 6, 9

  page 623. #1 – 8

Final is coming!

You guys can do it!