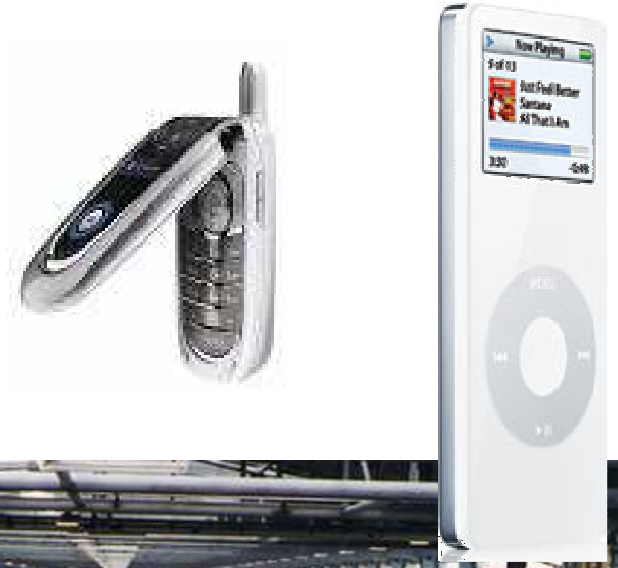
Lecture 28 – Chapter 19, Section 1
Redox Reactions – Oxidation Number

- Redox Reactions
- Oxidation Number



Redox Reactions – Nomenclature

- Oxidation-Reduction (Redox) reactions involve the transfer of electrons
- One reactant gains electrons – it is reduced
- The other loses electrons – it is oxidized

  LEO says GER (or other pneumonic)

- The thing that is oxidized leads to reduction of the other reactant – it is a reducing agent.
- The reactant that is reduced is an oxidizing agent.

- The oxidizing agent is reduced – gains electrons
- The reducing agent is oxidized – loses electrons

Example – Magnesium Combustion

- Remember last semester we used magnesium to generate H_2 gas? It would also have been fun to use pure oxygen gas…

$$2\ Mg_{(s)} + O_{2\ (g)} \rightarrow 2\ MgO\ (s)$$

- Mg goes from neutral Mg to Mg^{2+}
- O goes from neutral O to O^{2-}

- Mg loses electrons – it is oxidized. Mg is the reducing agent
- O gains electrons – it is reduced. O is the oxidizing agent

Recognizing Redox Rxns

Not always clear when electrons are transferred

$$FeCO_3\ (s) \rightarrow FeO(s) + CO_2(g)$$

$$FeO(s) + CO\ (g) \rightarrow Fe\ (l) + CO_2(g)$$

One of these is a redox reaction, the other is not.

How do we tell?

Oxidation Numbers

- Help to keep track of ‘where’ electrons are during a reaction
- We pretend that electrons are completely attached to one atom
  - Not really true, but helps us to keep track of which nuclei the electrons tend to be around
- In general the thing that is more electronegative will gain electrons

- The less electronegative species will loose electrons

Oxygen is very electronegative.

During a redox reaction oxygen is almost always …

25%	1.	Reduced
25%	2.	Oxidized
25%	3.	Neither
25%	4.	Both

1	2	3	4	5

Oxygen is very electronegative.

During a redox reaction oxygen almost always act as the …

25%	1.	Reducing Agent
25%	2.	Oxidizing Agent
25%	3.	Neither
25%	4.	Both

1	2	3	4	5

Oxidation Numbers – some guidelines

the electric charge an atom appears to have

1. Treat polyatomic ions separately.
2. Sum of all the oxidation numbers must equal the charge of the species.
3. Hydrogen usually has an oxidation number of +1.
4. The most electronegative atom in a species has a negative oxidation number equal to the number of electrons needed to complete its valence octet.

Hints on Oxidation Numbers

1. A pure element has an oxidation number of zero (Mg, O_2, etc...).
2. For monatomic ions, the oxidation number is equal to the charge on the ion (Ca^{2+} has a oxidation number of 2+).
3. Fluorine is always –1 in compounds with other elements.
4. Cl, Br and I are always –1 in the compound except when combined with oxygen or fluorine.
5. The oxidation number of H is +1 and O is –2 in most compounds.

   Exceptions:

   - When H is bonded to a metal, it becomes a –1. For example, in CaH_2, Ca (2+) and H (-1).
   - Oxygen can have a –1 charge when it is a peroxide

     any –O–O– (*e.g.* H_2O_2)

Some examples

- NH_4NO_3 : two polyatomic ions NH_4^+ and NO_3^-
  - NH_4^+: each H is +1, N must be –3
  - NO_3^-: each O is –2, N must be +5

- NaOCl: OCl^- is polyatomic ion
  - Na+: Na must be +1
  - OCl^-: O is –2, so Cl must be +1 (strange but true)

- MgO: ionic compound
  - O is –2, so Mg must be +2.

So what about those iron reactions?

Not always clear when electrons are transferred

$$FeCO_3\ (s) \rightarrow FeO(s) + CO_2(g)$$

$$FeO(s) + CO\ (g) \rightarrow Fe\ (l) + CO_2(g)$$

One of these is a redox reaction, the other is not.

We'll use our oxidation numbers to find out when electrons are moving around.

$$FeCO_{3\,(s)} \rightarrow FeO_{(s)} + CO_{2\,(g)}$$

- Let's look at the reactant
  - $FeCO_3$: CO_3^{2-} is polyatomic ion
    - CO_3^{2-} : O is –2, so C must be +4
    - Fe^{2+}: Fe must be +2
- Now let's look at the products
  - FeO:
    - O is –2, so Fe must be +2
  - CO_2:
    - O is –2, so C must be +4

- Everybody keeps the same oxidation number, so this is NOT a redox reaction

$$FeO_{(s)} + CO_{(g)} \rightarrow Fe_{(l)} + CO_{2\,(g)}$$

- FeO:
  - O is –2, so Fe must be +2
- CO:
  - O is –2, so C must be +2
- Fe:
  - Pure element so Fe must be 0
- CO_2:
  - O is –2, so C must be +4

- O is –2 throughout so it is neither oxidized nor reduced
- Fe goes from +2 to 0 → it gains electrons and is reduced
- C goes from +2 to +4 → looses electrons and is oxidized

Another example

- Let’s look at combustion

$$CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$$

- What is reduced and what is oxidized?

For the combustion of methane:

$CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$

What is the oxidation number of C in CH_4?

20%	1. –4
20%	2. –2
20%	3. 0
20%	4. +2
20%	5. +4

1	2	3	4	5

For the combustion of methane:

$CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$

H serves as the…

20%	1.	Oxidizing agent
20%	2.	Reducing agent
20%	3.	Neither
20%	4.	Both
20%	5.	This is not a redox reaction

1	2	3	4	5

For the combustion of methane:

$CH_{4\,(g)} + 2\,O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\,H_2O_{\,(g)}$

Oxygen is… ?

20%	1.	oxidized
20%	2.	reduced
20%	3.	neither
20%	4.	both
20%	5.	this is not a redox reaction

1	2	3	4	5

Today

- Pick up CAPA's
- Read Chapt 19
- Make sure you understand oxidation #
- **Go to Arkin seminar!**

Monday

- Start on CAPA #17 (due Thursday)
- Not too early to start reviewing last semester