Oxidation-Reduction (Redox) Review

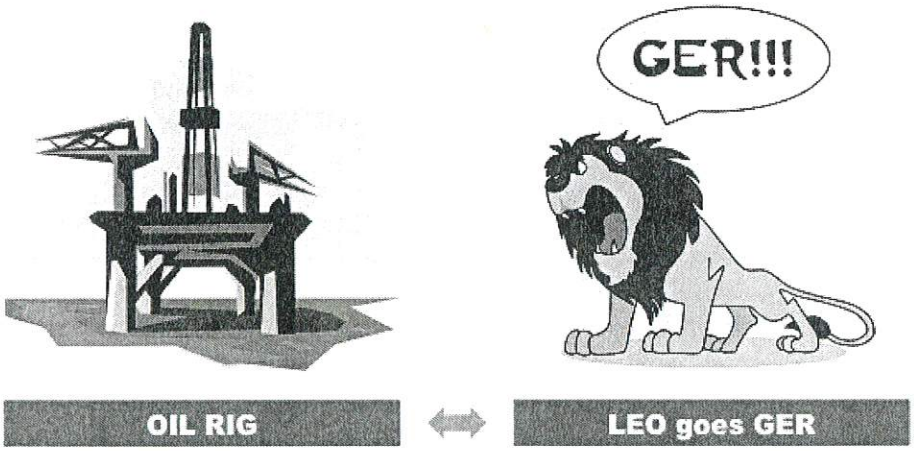
Oxidation–reduction (redox) reactions: where electrons are transferred from one atom to another.

- If a substance accepts an electron, it is reduced.
- If a substance loses an electron, it is oxidized.
- Electrons are always transferred from the species that is oxidized to the species that is reduced.

➔ Reduction (gain in electrons) + Oxidation (loss of electrons) = Redox!

Two great mnemonics!

1. OIL RIG: **O**xidation **I**s **L**oss (OIL) and **R**eduction **I**s **G**ain (RIG)
2. LEO goes GER: A species **l**oses **e**lectrons when **o**xidized, and **g**ains **e**lectrons when **r**educed.



***Almost* all reaction types (except double replacement) are redox.** However, that statement is not always true – the only way to be certain if a reaction is redox is to determine if any species has gained or lost electrons by looking for a change in oxidation state/number.

- If a chemical reaction does have a species which changes oxidation number: yes, **it's redox!**

Oxidation Numbers/ Oxidation States

Oxidation states are imaginary charges assigned based on a set of rules simply used to determine electron flow.

→ Even though they look like them, oxidation states are **NOT** ionic charges.

→ Oxidation numbers can be assigned to each atom in an element, ion, or compound...whether the compound is ionic or covalent!

How to Assign Oxidation States (aka Oxidation Numbers)

Free Elements or Monatomic Ions

1. Free elements = 0
2. Monoatomic ions = their charge

Examples	Free Elements			Monatomic Ions	
	$Fe(s)$	$Br_2(l)$	$O_3(g)$	Au^{3+}	S^{2-}
Oxidation #	0	0	0	+3	−2

Atoms in a Compound

3. All atoms in a neutral compound add up to 0.
4. All atoms in a polyatomic ion add up to the ion's charge.

	In order of priority	Oxidation State
Metals	Group 1A metals	+1
	Group 2A metals	+2
	Group 3A metals	+3
Non-metals	fluorine	−1
	hydrogen	+1
	oxygen	−2
	Group 7A	−1
	Group 6A	−2
	Group 5A	−3

Helpful hints:

- Group 4A (the carbon family) and transition metals are NOT listed – you will ALWAYS have to solve for them.
- Coefficients do NOT affect oxidation numbers.
- You can split up ionic compounds with a polyatomic and use the ion's overall charge to solve for the oxidation states of each element in the ion

***Easy practice*: Elements, Ions, and Simple Ionic Compounds**

1. $Na^+(aq)$

+1

2. $H_2(g)$

ϕ

3. $Na_2S(s)$

+1, −2

***Medium practice*: More compounds and polyatomic ions**

4. $C_2H_4(g)$

−2, +1

$2x + 4(+1) = \phi$

$2x + 4 = \phi$

$x = -2$

5. $Cr_2O_7^{2-}(aq)$

+6, −2

$2x + 7(-2) = -2$

$2x - 14 = -2$

$x = +6$

***Conflict-resolution practice*: What do you do if the rules don't agree?**

6. $SO_3^{2-}(aq)$

+4, −2

$x + 3(-2) = -2$

$x - 6 = -2$

$x = +4$

7. $H_2O_2(l)$

+1, −1

$2(+1) + 2x = \phi$

$2 + 2x = \phi$

$x = -1$

Fun practice! ☺

8. $Fe(ClO_4)_3(s)$

+3, +7, −2

Fe^{3+}, ClO_4^- ⇒ $x + 4(-2) = -1$

$x - 8 = -1$

9. $H_2SeO_3(aq)$

+1, +4, −2

$2(+1) + x + 3(-2) = \phi$

$2 + x - 6 = \phi$

$x = +4$

How to Identify What is Oxidized or Reduced in a Reaction

Once you have identified a redox reaction by the change in oxidation state, now you can tell what was oxidized or reduced!

a. A substance that has the element that has been oxidized (LOST electrons) will have an oxidation number that becomes more positive (or less negative).

b. A substance that has the element that has been reduced (GAINED electrons) will have an oxidation number that becomes more negative (or less positive).

Balancing Redox Reactions: We split redox reactions into two separate reactions

- The oxidation half-reaction has electrons as a product.
- The reduction half-reaction has electrons as a reactant.

Oxidation Half-Reaction	Reduction Half-Reaction
$Zn(s) \rightarrow Zn^{2+}(aq) + 2\,e^-$	$Fe^{2+}(aq) + 2\,e^- \rightarrow Fe(s)$

Practice!

1. Using oxidation numbers, identify what was oxidized and reduced in each reaction below.

a) $4\,Al(s) + 3\,O_2(g) \rightarrow 2\,Al_2O_3(s)$

$\phi \quad \phi \quad +3, -2$

oxidized: $Al(s)$ $(\phi \rightarrow +3)$

reduced: $O_2(g)$ $(\phi \rightarrow -2)$

b) $2\,H_2O(l) + 4\,MnO_4^-(aq) + 3\,ClO_2^-(aq) \rightarrow 4\,MnO_2(aq) + 3\,ClO_4^-(aq) + 4\,OH^-(aq)$

$+1, -2 \quad +7, -2 \quad +3, -2 \quad +4, -2 \quad +7, -2 \quad -2, +1$

oxidized: ClO_2^- $(Cl: +3 \rightarrow +7)$

reduced: MnO_4^- $(Mn: +7 \rightarrow +4)$

2. In the reaction below, a piece of solid nickel is added to a solution of potassium dichromate.

$$14\,H^+(aq) + Cr_2O_7^{2-}(aq) + 3\,Ni(s) \rightarrow 2\,Cr^{3+}(aq) + 3\,Ni^{2+}(aq) + 7\,H_2O(l)$$

$+1 \quad +6, -2 \quad \phi \quad +3 \quad +2 \quad +1, -2$

Which species is being oxidized and which is being reduced?

	Oxidized	Reduced
a.	$Cr_2O_7^{2-}(aq)$	$Ni(s)$
b.	$Cr^{3+}(aq)$	$Ni^{2+}(aq)$
(c.)	$Ni(s)$	$Cr_2O_7^{2-}(aq)$
d.	$Ni^{2+}(aq)$	$Cr^{3+}(aq)$