Oxidation – Reduction Chemistry

REDOX

- Balance the following equation by inspection:

*(Note: this is a net ionic equation – **charge** and **number of atoms** must be balanced.)*

$\_\_\ H^+ + \_\_\ Cr_2O_7^{2-} + \_\_C_2H_5OH \rightarrow$

$\_\_\ Cr^{3+} + \_\_\ H_2O + \_\_\ CO_2$

Oxidation-Reduction Reaction

- A reaction in which electrons are transferred from one species to another.
- Many common chemical reactions are REDOX reactions:
  - "Rusting" or oxidation of metals
  - Combustion reactions
  - Single replacement reactions
  - Metabolism of sugars by our bodies

$2\ Na_{(s)} + Cl_{2(g)} \rightarrow 2\ NaCl_{(s)}$

- In this synthesis reaction, electrons are transferred from sodium to chlorine:
- Two Na atoms lose an electron (oxidation):

  $2\ Na \rightarrow 2\ Na^+ + 2\ e^-$
- Each Cl gains an electron (reduction):

  $Cl_2 + 2\ e^- \rightarrow 2\ Cl^-$
- The above reactions are called half-reactions, as each represents ½ of the actual reaction.

Oxidation States (Numbers)

- Oxidation Numbers are a convenient way to keep track of electrons in REDOX reactions.
- Oxidation numbers are SIMILAR to charge, but they are NOT the same thing.
- Oxidation numbers may be assigned to elements in compounds in which electrons are shared – not just to charged species.

Assigning Oxidation Numbers

1. The oxidation number of atoms in their elemental form is zero.
2. The oxidation state of monatomic ions is the same as the charge.
3. The oxidation state of fluorine is always -1 in its compounds.
4. The oxidation state of other halogens (Cl, Br, I) is -1 unless combined with O, F, or a more reactive halogen. In these cases assign using rule #7.

Assigning Oxidation Numbers (cont.)

5. Oxygen is almost always assigned an oxidation number of -2 in its compounds. (Exceptions: OF_2, peroxides and superoxides)

6. The oxidation state of Hydrogen in compounds is usually +1. (Exception: In binary compounds with metals, hydrogen acts as the hydride ion, H^{-1})

Assigning Oxidation Numbers (cont.)

7. All other oxidation numbers may be assigned by the following principles:
   - The sum of the oxidation states in a neutral compound must equal zero.
   - The sum of the oxidation states in a polyatomic ion must equal the charge on the ion.

$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$

- In the above reaction, assign oxidation numbers to all elements.
- *Carbon is losing electrons.*
- Carbon is OXIDIZED. Its oxidation number is being *increased*.
- *Oxygen is gaining electrons.*
- Oxygen is REDUCED. Its oxidation number is being *reduced*.

LEO the lion goes **GER**

- **LEO = Losing Electrons Oxidation**
- **GER = Gaining Electrons Reduction**

$SiCl_4 + 2\ Mg \rightarrow 2\ MgCl_2 + Si$

- The species (atom, ion, or molecule) that is oxidized is called the reducing agent.
- The species that is reduced is called the oxidizing agent.
- ***In the above reaction, identify the atoms that are oxidized and reduced.***
- ***Identify the oxidizing and reducing agents in the above reaction.***

$Al + MnO_4^- \rightarrow MnO_2 + Al(OH)_3$

Consider the above, UNBALANCED redox reaction.

- What is being oxidized?
- What is being reduced?
- What is the oxidizing agent?
- What is the reducing agent?

Balancing Redox Equations:
The Method of ½ - Reactions

- Balance the reaction:

  $MnO_4^- + Fe^{+2} \rightarrow Fe^{+3} + Mn^{+2}$

  in acidic (H+) aqueous solution (H_2O).

- ***Step 1***
  Identify as an oxidation-reduction reaction.
  Assign oxidation numbers (if needed).

Balancing Redox Equations (cont.)

- ***Step 2***
  Separate into ½-reactions: ***Oxidation & Reduction***

- ***Step 3***
  A) Balance the ½-reactions for mass (# of atoms).
    1) Balance all atoms except H or O (unless they are the oxidized or reduced species).
    2) Balance O with waters (H_2O).
    3) Balance H with acid (H^+ ions).

  B) Balance ½-reactions for charge using electrons.
  (***e***$^-$ is a product in oxidation reactions and a reactant in reduction reactions.)

Balancing Redox Equations (cont.)

- ***Step 4***
  Multiply ½-reactions by appropriate factors to equalize electrons.

- ***Step 5***
  Add ½-reactions together to get the overall reaction.

- ***Step 6***
  CHECK to see that reaction is balanced for charge and # of atoms.

Balance the following Reactions:

1. $\mathbf{Cu_{(s)} + NO_3^-{}_{(aq)} \rightarrow Cu^{2+}{}_{(aq)} + NO_{(g)}}$
   in acidic aqueous solution.

2. $\mathbf{Cr_2O_7^{2-}{}_{(aq)} + C_2H_5OH_{(aq)} \rightarrow Cr^{3+}{}_{(aq)} + CO_{2(g)}}$
   in acidic aqueous solution.

Disproportionation

- An oxidation-reduction reaction in which the reactant is both oxidized and reduced.

- Balance the following redox reaction in acidic aqueous solution:

$$\mathbf{Cl_{2(g)} \rightarrow Cl^-{}_{(aq)} + ClO_2^-{}_{(aq)}}$$

Balancing redox reactions in BASIC (OH^-) aqueous (H_2O) solution:

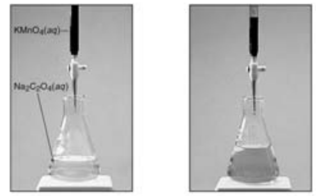
1) Balance as if in ACIDIC solution.
2) Add OH^- ions equal to the number of H^+ ions to BOTH SIDES of the equation.
3) Combine OH^- and H^+ ions on the same side into waters. Cancel waters.
4) Check balancing for charge and # of atoms.

- Balance the following reactions in BASIC solution.
  A. $\mathbf{NO_2^-{}_{(aq)} + Al_{(s)} \rightarrow NH_{3(g)} + AlO_2^-{}_{(aq)}}$
  B. $\mathbf{Ag_{(s)} + CN^-{}_{(aq)} + O_{2(g)} \rightarrow AgCN_{(s)} + H_2O_{(l)}}$

Redox titrations

- The previous titration examples we have considered were all acid-base titrations.
- However, we can measure the number of moles by titration for other types of reactions as well.
- If the substance to be measured will react with another compound in such a way that we can observe an endpoint (such as with an indicator), titration may be useful.

Redox Titration



Net ionic equation:

$$2\overset{+7}{Mn}O_4^-(aq) + 5\overset{+3}{C_2}O_4^{2-}(aq) + 16H^+(aq) \longrightarrow 2\overset{+2}{Mn}{}^{2+}(aq) + 10\overset{+4}{C}O_2(g) + 8H_2O(l)$$

What is the concentration of sodium oxalate ($Na_2C_2O_4$) if a 50.0 mL sample of solution requires 42.77 mL of 0.164 *M* $KMnO_4$ to fully titrate it?