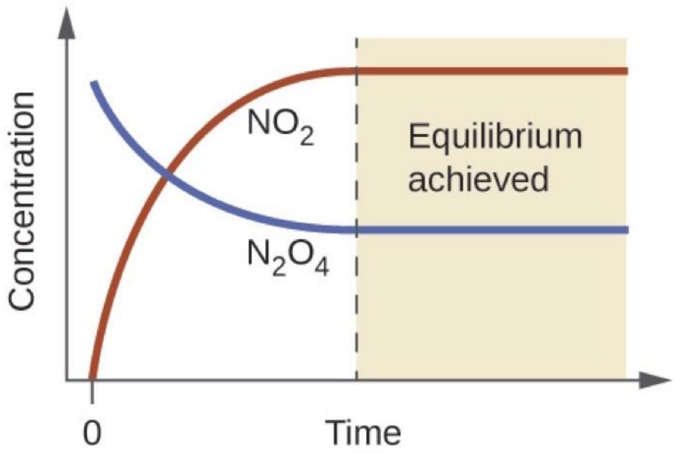

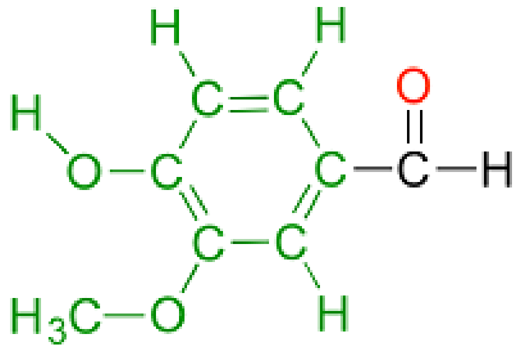

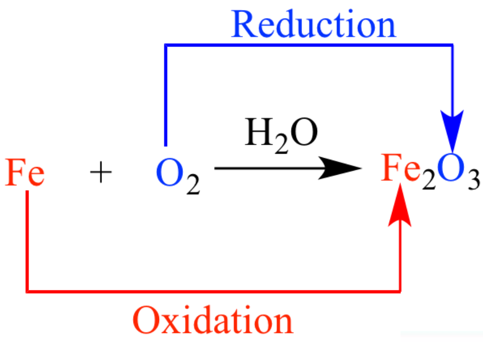

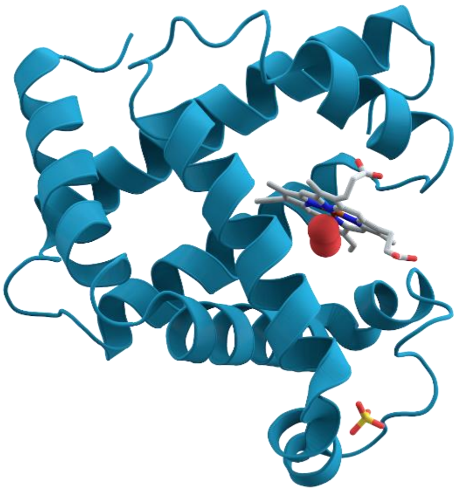

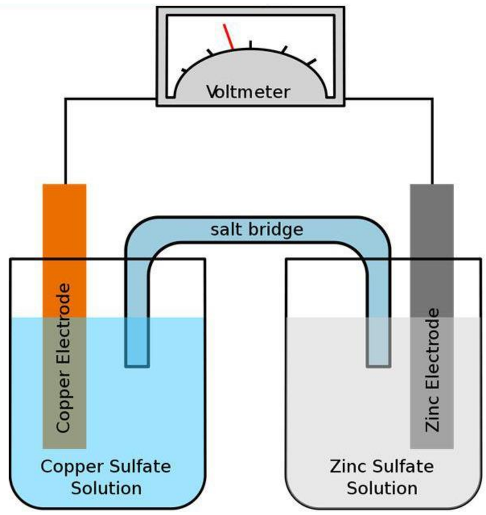

TOPIC NOTES FOR CHEM1202: CHEMISTRY FOR LIFE SCIENCE

Completed in 2016 with High Distinction

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WK1: Reactions in Aqueous Solution- (Dr. Jason Gascooke)

Solution: A homogenous mixture of a solute dissolved in a solvent

Electrolyte: substances that dissociate into ions in water (its 'strength' is a measure of how much of the solute dissolves into ions)

Solution State	Ability to conduct electricity
Non-electrolyte	Non-existent
Weak-electrolyte	Poorly
Strong electrolyte	Very effectively

Non-electrolyte: substances that dissolve in water but do not dissociate into ions

- → Ionic compounds form strong electrolytes (provided they dissolve), whereas molecular compounds can produce non-, weak or strong electrolytes.

Exchange reactions: reactions in which anions and cations exchange partners

- → $A^+B^- + C^+D^- \rightarrow A^+D^- + C^+B^-$
- → IF C^+B^- is soluble then there is no net reaction (all ions in solution)
- → But if the combination is not soluble then a net reaction has occurred
  - $C^+B^-_{(s)}$: precipitate
  - $C^+B^-_{(aq)}$: weak or nonelectrolyte
  - $C^+B^-_{(g)}$: gas

Acid-base reactions		
	An acid produce H^+ (aq) in water.	A base produces OH^-(aq) in water
100% dissociation	Strong acid	Strong Base
Small proportion dissociates	Weak acid	Weak base

WK1: Redox Reactions – (Dr. Jason Gascooke)

(reduction- oxidation) Redox reactions: reactions where electrons are transferred between species. These reactions create a flow of electrons which is utilised for electricity.

Oxidised: to loose electrons, the species that is oxidised is called the reducing agent or reductant

Reduced: to gain electrons, the species that is reduced is called the oxidising agent or oxidant,

Full equation: $Cu_{(s)}$ *(reductant)* + $2Ag^+_{(aq)}$ *(oxidant)* → $2Ag_{(s)}$ + $Cu^{2+}_{(aq)}$

Cu ½ reaction: $Cu_{(s)} \rightarrow Cu^{2+}_{(aq)} + 2e^-$ *Oxidation*

Ag ½ reaction: $2Ag^+_{(aq)} + 2e^- \rightarrow 2Ag_{(s)}$ *Reduction*

The electrons lost via oxidation are all gained via reduction.

The <u>oxidation numbers</u> (or oxidation states) represent the total charge the atom would have if all the bonds were ionic. It can be utilized to determine the movement of electrons so to determine which reactants are oxidised or reduced.

WK1: Redox Rules and Oxidation Numbers – (Dr. Jason Gascooke)

The rules of redox reactions are:

1. Any atom in its elemental state has an oxidation number of zero. E.g. O_2, C, N_2
2. For <u>monoatomic ions</u> the oxidation number is equal to the net charge on the species e.g. Fe^{3+}(+3), O^{2-}(-2)
3. Fluorine atoms always have an oxidation number of -1, e.g. NaF, HF
4. Chlorine, Bromine and Iodine atoms always have oxidation numbers of -1 except in compounds with oxygen or fluorine (ClO_4^-) e.g. HCl, KBr
5. Hydrogen atoms are usually +1 in most compounds, except metal hydrides (NaH, CaH_2) where the oxidation number is -1.
   Oxygens atoms are usually -2 in most compounds, with the exception of peroxides (H_2O_2) e.g. H_2O
6. The sum of the oxidation numbers for all atoms in a polyatomic species are equal to the species' charge. E.g. for NH_4^+ which has an overall charge of +1, the sum of the oxidation numbers = N(-3) + 4H(+1) = +1

Oxidation = an increase in oxidation number

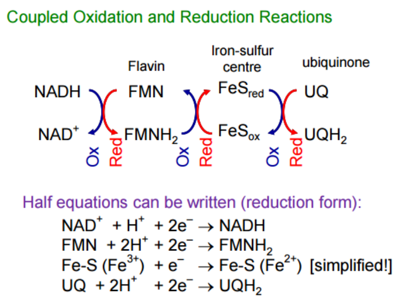
Reduction = a decrease in oxidation number

Common Redox reactions:

<u>Rusting</u>: $2Fe_{(s)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 2Fe(OH)_{2(S)}$ $\quad$ $4Fe(OH)_{2(s)} + O_{2(g)} \rightarrow 2Fe_2O_3.H_2O_{(s)} + 2H_2O_{(l)}$

<u>Silver tarnishing</u>: $2Ag_{(s)} + H_2S_{(g)} \rightarrow Ag_2S_{(s)} + H_{2(g)}$

Redox reactions are essential for energy storage and conversion in biological organisms.



$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$

Many of the reduction and oxidation reactions in biochemical reactions are carried out by <u>NADH</u> or <u>NAD^+</u>. And the direction of the reaction can change depending on whether a species 'needs' to be oxidised or reduced.

i.e. $NAD^+_{(aq)} + H^+_{(aq)} + 2e^- \underset{reduction}{\overset{oxidation}{\leftrightharpoons}} NADH_{(aq)}$

WK2: Balancing Redox Reactions – (Dr. Jason Gascooke)

In every redox reaction there are two reactions, a reduction and oxidation reaction referred as the half reactions.

Redox reaction steps:

1. Work out whether it is a redox reaction via checking the oxidation numbers.
2. Write the half reactions
3. Balance the half reactions for each element
   i. Atoms other than H and O
   ii. Balance O by adding H_2O

iii. Balance H by adding H^+

4. Balance charges in half reactions with electrons.
5. Multiply half reactions by whole numbers so the number of electrons on both sides of the reaction are equal (electrons lost = electrons gained)
6. Add the balanced half reactions together to get the overall reaction
7. Simplify overall equation if necessary
8. Check that the overall reaction is balanced.

(acidic conditions)

$Cu + NO_3^- \rightarrow Cu^{2+} + NO_2$