This week's schedule

Group 1: Lab Synthesis

Group 2: Tutorial Stoichiometry 1 & 2

Redox Example (in acid solution):

Oxidation

+7 -2 +1 +3 -2 +2 +4 -2

$MnO_4^-(aq) + H_2C_2O_4(aq) \rightarrow Mn^{2+}(aq) + CO_2(g)$

Reduction

$\therefore$ Half reactions are:

a) $MnO_4^- \rightarrow Mn^{2+}$

b) $H_2C_2O_4(aq) \rightarrow CO_2\ (g)$

a) $5e^- + 8H^+ + MnO_4^- \rightarrow Mn^{2+} + 4H_2O$

$Charge_{LHS} = 8 - 1 = +7$ $Charge_{RHS} = +2$ $\therefore$ Add $5e^-$ to reactant side **(Reduction as expected)**

b) $H_2C_2O_4\ (aq) \rightarrow 2\ CO_2\ (g) + 2H^+ + 2e^-$

$Charge_{LHS} = 0$ $Charge_{RHS} = +2$

- Products gain 2 e^-;
- **(Oxidization as expected)**

c) Multiply reduction reaction by x2 and the oxidation reaction by x 5 so that both reactions involve 10 e^-

$$10\ e^- + 16H^+ + 2MnO_4^- \rightarrow 2Mn^{2+} + 8H_2O$$

+

$$5H_2C_2O_4 \rightarrow 10CO_2 + 10H^+ + 10e^-$$

$$6\ \cancel{16}H^+ + 5H_2C_2O_4 + 2\ MnO_4^- \rightarrow 2Mn^{2+} + 10CO_2 + \cancel{10H^+} + 8H_2O$$

$$\boxed{6H^+ + 5H_2C_2O_4 + 2\ MnO_4^- \rightarrow 2Mn^{2+} + 10CO_2 + 8H_2O}$$

Mass check: 2Mn, 28O, 10C and 16H on both sides

$Charge_{LHS} = +6 - 2 = +4$ $Charge_{RHS} = +4$

Example (in basic solution):

Reduction

$$\overset{+2\ \ -1}{FeCl_2} + \overset{+1\ +7\ -2}{KMnO_4} \rightarrow \overset{+3\ \ -2}{Fe_2O_3} + \overset{+4\ \ -2}{MnO_2} + \overset{+1\ \ -1}{KCl}$$

Oxidation

Drop the counter ions whose oxidation state doesn't change: K^+, Cl^-.
Write the two half reactions:

a) $Fe^{+2} \rightarrow Fe_2O_3$

b) $MnO_4^- \rightarrow MnO_2$

a) $6OH^- + \cancel{3H_2O} + 2\ Fe^{2+} \rightarrow Fe_2O_3 + \overset{3}{\cancel{6}}H_2O + 2e^-$

$6OH^- + 2Fe^{2+} \rightarrow Fe_2O_3 + 3H_2O + 2e^-$ **Oxidation as expected**

b) $3e^- + \overset{2}{\cancel{4}}H_2O + MnO_4^- \rightarrow MnO_2 + \cancel{2H_2O} + 4OH^-$

$3e^- + 2H_2O + MnO_4^- \rightarrow MnO_2 + 4OH^-$ **Reduction as expected**

Multiply a) by x3 and b) by x2 to balance the number of electrons

$$18OH^- + 6Fe^{2+} \rightarrow 3Fe_2O_3 + 9H_2O + 6e^-$$

+

$$6e^- + 4H_2O + 2MnO_4^- \rightarrow 2MnO_2 + 8OH^-$$

$$\cancel{6e^-} + \overset{10}{\cancel{18}}OH^- + \cancel{4H_2O} + 6Fe^{2+} + 2MnO_4^- \rightarrow 3Fe_2O_3 + 2MnO_2 + \overset{5}{\cancel{9}}H_2O + \cancel{8OH^-} + \cancel{6e^-}$$

$$10OH^- + 6Fe^{2+} + 2MnO_4^- \rightarrow 3Fe_2O_3 + 2MnO_2 + 5H_2O$$

Mass balance: √ Charge balance: √

Disproportionation Reactions

For elements with 3 or more oxidation states, an element in an intermediate oxidation state can be both oxidized and reduced (a redox reaction know as disproportionation)

Example:

$$2\overset{+1}{Cu}\overset{-1}{Cl} \rightarrow \overset{0}{Cu} + \overset{+2}{Cu}\overset{2x-1}{Cl_2}$$

Here oxidation numbers help establish if redox reaction is also a disproportionation reaction

Such reactions can sometimes be difficult to balance

Example: Balance $P_4 \rightarrow PH_3 + H_2PO_2^-$ in basic solution

Hint: Rewrite the equation with the disproportionating species written as reacting with itself

See next lecture for how to balance this.

Electronic Structure and Periodic Properties

C020

Electron Orbitals

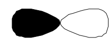
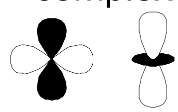
- The properties of elements result from the **electronic structure** of the atom
- Each electron in an atom is in an **orbital** of fixed energy; that is the orbitals are said to be **quantized**
- Each electron is described by **3 quantum numbers**

i) Principle quantum number n:

n = 1,2,3, …. This quantum number determines the energy of the orbital

ii) **Angular momentum quantum number, ℓ:**

ℓ = 0, 1, 2, …., n-1 This quantum number determines the shape of various orbitals having the energy of quantum number n

ℓ	**orbital name**	**shape**
0	s	spherical
1	p	dumbbell 
2	d	complex 
3	f	complex