This week's schedule

Group 1: Lab Synthesis

Group 2: Tutorial Stoichiometry 1 & 2

Redox Example (in acid solution):

Oxidation

$\overset{+7\ -2}{MnO_4^-}(aq) + \overset{+1}{H_2}\overset{+3}{C_2}\overset{-2}{O_4}(aq) \rightarrow \overset{+2}{Mn^{2+}}(aq) + \overset{+4\ -2}{CO_2}(g)$

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

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

- 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

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

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

$$
\begin{array}{l}
18OH^- + 6Fe^{2+} \rightarrow 3Fe_2O_3 + 9H_2O + 6e^- \\
+ \\
6e^- + 4H_2O + 2MnO_4^- \rightarrow 2MnO_2 + 8OH^- \\
\hline
\end{array}
$$

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

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

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

$$\overset{0}{P_4} + \overset{0}{P_4} \rightarrow \overset{-3}{P}\overset{+1}{H_3} + \overset{+1}{H_2}\overset{+1}{P}\overset{-4}{O_2^-}$$

Oxidation

Reduction

Half reactions:

$$12e^- + 12H_2O + P_4 \rightarrow 4PH_3 + 12OH^-$$

$$8OH^- + \cancel{8H_2O} + P_4 \rightarrow 4\,H_2PO_2^- + \cancel{8H_2O} + 4e^- \quad (X\ 3)$$

$$\cancel{12e^-} + 12H_2O + 4P_4 + \overset{12}{\cancel{24}}OH^- \rightarrow 4PH_3 + 12H_2PO_2^- + \cancel{12OH^-} + \cancel{12e^-}$$

$$= 12H_2O + 4P_4 + 12OH^- \rightarrow 4PH_3 + 12H_2PO_2^-$$

Divide equation by 4

$$\boxed{3H_2O + P_4 + 3OH^- \rightarrow PH_3 + 3H_2PO_2^-}$$

Mass check: 9H, 6O and 4 P on both sides of the equation

$Charge_{LHS} = -3$ $\quad$ $Charge_{RHS} = -3$ $\quad$ **Balanced**