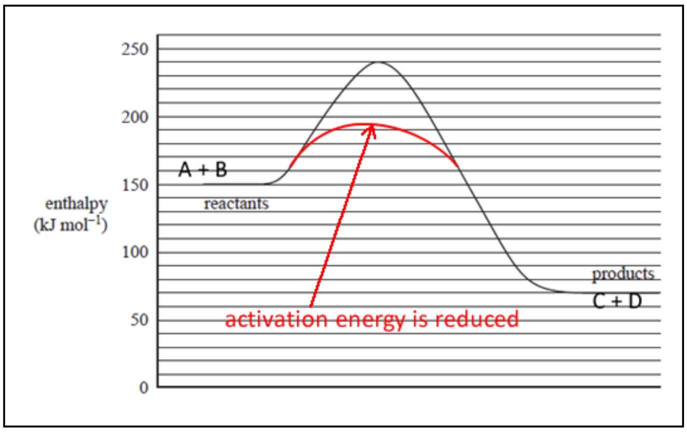
Ongoing Revision 7 – enthalpy, galvanic, redox reaction, isomers and atom economy.




1) The energy profile of the reaction $A(g) + B(g) \rightarrow C(g) + D(g)$ is shown on the right.

*a)* What is the ΔH for the reaction below?

$C(g) + D(g) \rightarrow A(g) + B(g)$ *ΔH =+80 kJ/mol*

b) What is the activation energy for the reaction $C(g) + D(g) \rightarrow A(g) + B(g)$? *170kJ/mol*

c) Draw on the energy profile above what changes take place when a catalyst is added.

d) Explain how a catalyst changes the forward and backward rates of reaction.

*At equilbrium a catalyst increases the rate of the backward and forward reactions equally. Prior to reaching equilibrium the catalyst causes a net increasese in the rate of the forward reaction causing the reaction to reach equilibrium faster.*

2) Consider the following thermochemical equations.

1) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H = -393.5$ kJ mol$^{-1}$ and

2) $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -571.6$ kJ mol$^{-1}$

a) Calulate the ΔH for the reaction below

$2C(s) + 4H_2O(l) \rightarrow 2CO_2(g) + 4H_2(g)$

*Step 1 reverse equation 2*

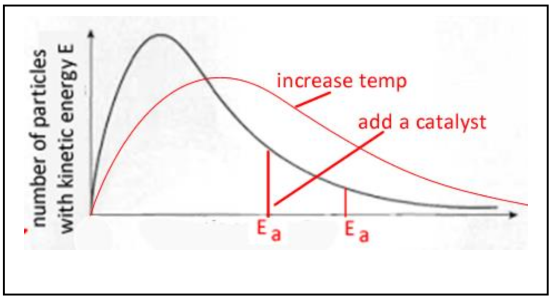
*3) $2H_2O(l) \rightarrow 2H_2(g) + O_2(g)$ $\Delta H = +571.6$ kJ mol$^{-1}$*

*Step 2 add equations 3 and 1*

*4) $C(s) + 2H_2O(l) \rightarrow CO_2(g) + 2H_2(g)$ $\Delta H = +178.1$ kJ mol$^{-1}$*

*Step 3 multiply equation 4 by 2.*

*$2C(s) + 4H_2O(l) \rightarrow 2CO_2(g) + 4H_2(g)$ $\Delta H = +356.2$ kJ mol$^{-1}$*




b) At a given temperature the distribution of the kinetic energy of reactant particles of the reaction $A(g) + B(g) \rightarrow C(g) + D(g)$ is shown on the right. $E_a$ is the activation energy for this reaction. At this temperature the number of successful collisions between reactant particles is very low. With the exception of increasing concentration and increasing surface area of reactant particles, discuss two other ways in which the rate of the reaction can be increased.

*Increase temperature or add a catalyst*

d) Draw on the energy distributaion curve, shown above, how each method mentioned in c) changes the curve and how this impacts on the rate of the reaction.

*Adding a catalyst reduces the activation energy of the reaction and hence more particles can undergo fruitfull collisions.*

*Increasing the temperature increases the number of particles with the required activation energy so more fruitfull collisions occur.*

3) Write the balanced chemical equation for the combustion of liquid propane and justify why this is or is not a redox reaction.

$C_3H_8(l) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

Carbon has an oxidation state of -2.67 in propane but an oxidatin state of +4 in $CO_2$. Carbon is oxidised while oxygen goes from an oxidation state of 0 in $O_2$ to -2 in $CO_2$ and $H_2O$, it is reduced. *This is a redox reaction.*

4) Consider the following cells.

Fuel cell, primary galvanic cell, secondary galvanic cell, electrolytic cell.

a) Which one of the classes of electrochemical cells, above, involves **only** a non-spontaneous redox reaction? *Electrolytic cell*

b) What is the difference between a secondary and primary cell?

*Secondary cells can be recharged whereas primary cells cannot be recharged.*

c) A student was heard to say " A fuel cell is a galvanic cell that can be recharged continuously" Is this comment true or false? Justify your answer.

*To recharge a battery is to provide energy to reverse the oxidation and reduction reactions that have taken place at each electrode. A fuel cell does not store the products of each reaction and hence can not be recharged. In a fuel cell reactants are constantly supplied and products continuously removed.*

d) Methanol is suitable for use in a micro fuel cell that is used to power laptop computers and similar small electrical items. The methanol is oxidised to carbon dioxide and water. The overall equation for the reaction is

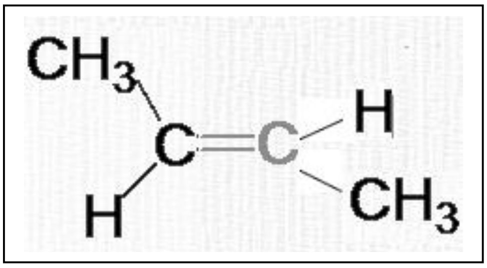
$CH_3OH(l) + 2O_2(l) \rightarrow 2H_2O(l) + CO_2(g)$

Write the balanced half equation for the

i. anode reaction $CH_3OH(l) + H_2O(l) \rightarrow CO_2(g) + 6H^+(aq) + 6e$

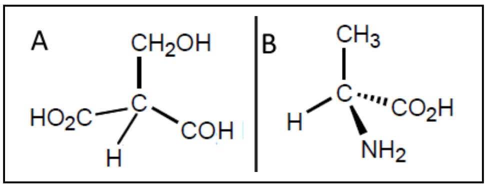
ii. cathode reaction $4H^+(aq) + 4e + O_2(l) \rightarrow 2H_2O(l)$

5) Name the molecule shown on the right.



*trans-but-2-ene*

6) Consider the two molecules shown on the right.



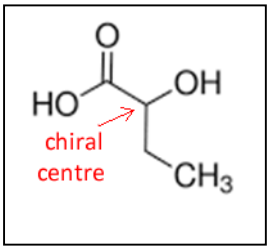
a) Which of the molecules shown are chiral?

*Both A and B*

b) How many optical isomers are possible for each?

*Each molecule has one chiral centre, hence two optical isomers each.*

c) i. Draw the structural formula for 2-hydroxybutanoic acid.




ii. Is it a chiral molecule and if so how many optical isomers are possible?

*It has one chiral centre and hence has two optical isomers.*

d) i. Write a balanced chemical equation for the formation of pentyl butanoate form pentanol and butanoic acid. (States not required)

$CH_3CH_2CH_2CH_2CH_2OH + CH_3CH_2CH_2COOH \rightarrow CH_3CH_2CH_2CH_2CH_2OOCCH_2CH_2CH_3 + H_2O$

ii. Calculate the percentage atom economy for this reaction.

*Formula mass of pentanol = 88.1*

*Formula mass of butanoic acid = 88.1*

*Formula mass of pentyl butanoate = 158.2*

⇨ *(158.2 / (2 X 88.1)) X 100 = 89.8%*

iii. Calculate the percentage yield of the reaction if 1.76 grams of butanoic acid was placed in excess pentanol to produce 2.56 grams of pentyl butanoate.

*Step 1 calculate the theoretical yield of pentyl butanoate*

*=> mol of butanoic acid = 1.76 / 88.1 = 0.0200*

*=> 0.0200 mol of pentyl butanoate should form*

*Step 2 theoretical mass of pentyl butanoate*

*=> 0.0200 X 158.2 = 3.16 g*

*Step 3 percentage yield*

*=> (2.56/3.16) X 100 = 81.0% yield.*