**5.1 Exothermic and Endothermic Reactions**

**5.1.1** : If the reaction produces heat (increases the temperature of the surroundings) then it's exothermic. If it decreases the temp (ie absorbs heat) then it's endothermic. Also, the yield of an equilibrium reaction which is exothermic will be increased if it occurs at low temps, and so for endothermic reactions at high temps.

**5.1.2** : Exothermic -> a reaction which produces heat. Endothermic -> a reaction which absorbs heat. Enthalpy of reaction -> the change in internal energy (H) through a reaction is ∆H.

**5.1.3** : ∆H will be negative for exothermic reactions (because internal heat is being lost) and positive for endothermic reactions (because internal energy is being gained).

**5.1.4** : The most stable state is where all energy has been released...therefore when going to a more stable state, energy will be released, and when going to a less stable state, energy will be gained. On an enthalpy level diagrams, higher positions will be less stable (with more internal energy) therefore, if the product is lower, heat is released (more stable, ∆H is -ve) but if it is higher, heat is gained (less stable, ∆H is +ve).

**5.1.5** : Formation of bonds -> release of energy. Breaking of bonds -> gain / absorption of energy.

**5.2 Calculation of enthalpy changes**

**5.2.1** : change in energy = mass x specific heat capacity x change in temperature --> E = mc∆T

**5.2.2** : Enthalpy changes (∆H) are related to the number of moles in the reaction...if all the coefficients are doubled, then the value of ∆H will be doubled (attention must be paid to limiting reagents though).

**5.2.3** : When a reaction is carried out in water, the water will gain or lose heat from (or to) the reaction, with hopefully little escaping the water. Therefore, the change in energy, and so the ∆H value, can be calculated with E = mc∆T where E is equal to ∆H, m is the mass of water present, and c = 4.18 kJ Kg-1 K-1. This ∆H value can then be calculated back to find the enthalpy change for each mol of reactants.

**5.2.4** : The solution should be placed in a container as insulated as possible, to keep as much heat as possible from escaping. The temperature should be measured continuously , and the value used in the equation is the maximum change in temp from the initial position.

**5.2.5 :** The results will be a change in temperature. this can be converted into a change in heat (or energy) by using the above equation and a known mass of water. this can be used to calculate the ∆H for the amount of reactants present, and then this can be used to calculate for a given number of mols.

**5.3 Hess's Law**

**5.3.1** : Hess' Law states that the total enthalpy change between given reactants and products is that same regardless of any intermediate steps (or the reaction pathway). To calculate:

1) Reverse any reactions which are going the wrong way and invert the sign of their ∆H values (multiply by -1)

2) Divide or multiply the reactions until the intermediate products will cancel out when the reactions are vertically added (always multiply/divide the ∆H value by the same number)

3) Vertically add the reactions.

4) Divide or multiply the resulting reaction to the correct coefficients.

**5.4 Bond enthalpies**

**5.4.1** : Bond enthalpy (aka dissociation enthalpy) -- the enthalpy change when one mole of bonds are broken homolitically in the gas phase. ie X-Y(g) -> X(g) + Y(g) : ∆H(dissociation). Molecules such as CH4 have multiple C-H bonds to be broken, and so the bond enthalpy for C-H is actually an average value. These values can be used to calculate unknown enthalpy changes in reactions where only a few bonds are being formed/broken.

**5.4.2 :** If the reaction can be expressed in terms of the breaking and formation of bonds in a gaseous state, then by adding (or subtracting when bonds are formed) the ∆H values the total enthalpy of reaction can be found.

**5.5**

**5.5.1** An increased disorder (entropy) can be caused by mixing two different types of particles, increased movement of particles (including state changes), or increased number of particles. Increasing the number of particles in a gaseous state gives the largest change in entropy.

**5.5.2** Since ∆S = -∆H/T, if ∆H is -, then ∆s is positive, and visa versa.

 **5.6**

**5.6.1** Standard free energy of reaction is the free energy that a reaction takes or gives at standard values of temperature and pressure,.

**5.6.2** If ∆G is negative, the reaction is spontaneous. If ∆G is positive, the reaction is not spontaneous.

**5.6.3** If ∆sT > ∆H, the reaction is spontaneous. If temperature drops so that ∆sT < ∆H, then the reaction is non-spontaneous.