CHEMISTRY

2.3.1 ENTHALPY CHANGES

ENTHALPY, H

- Heat content that is stored in a chemical system

CHEMICAL SYSTEM

- The reactants and products

SURROUNDINGS

- Outside the chemical system

CONSERVATION OF ENERGY

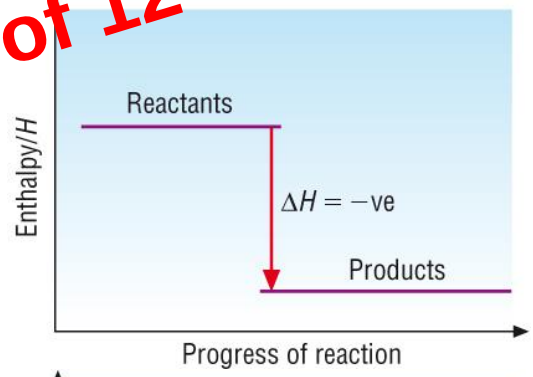
- No energy is lost
- Energy changes is measured as heat
- Heat loss in a chemical system = heat gain to surroundings
- Heat gain in a chemical system = heat loss from surroundings

ENTHALPY CHANGE, ΔH

- The heat exchange with the surroundings during a chemical reaction, at constant pressure
- $\Delta H = H_{products} - H_{reactants}$
- All chemical reactions either release or absorb heat

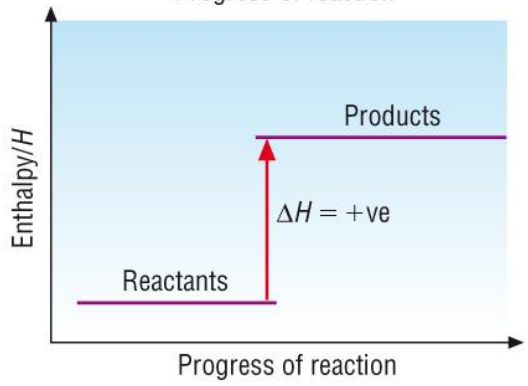
EXOTHERMIC REACTIONS

- ΔH, negative
- Enthalpy of products is smaller than the enthalpy of the reactants
- Heat loss to the surroundings
- E.g. chemicals reacting together in an inner chamber of self-heating cans
- $CaO\ (s) + H_2O\ (l) \rightarrow Ca(OH)_2\ (aq)$ ΔH = -ve



ENDOTHERMIC REACTIONS

- ΔH, positive
- Enthalpy of products is greater than the enthalpy of the reactants
- Heat gain from the surroundings
- E.g. evaporation of water to absorb heat from beer in self-cooling bear cans
- $H_2O\ (l) \rightarrow H_2O\ (g)$ ΔH = +ve



EXOTHERMIC REACTIONS

EXO OR ENDO?

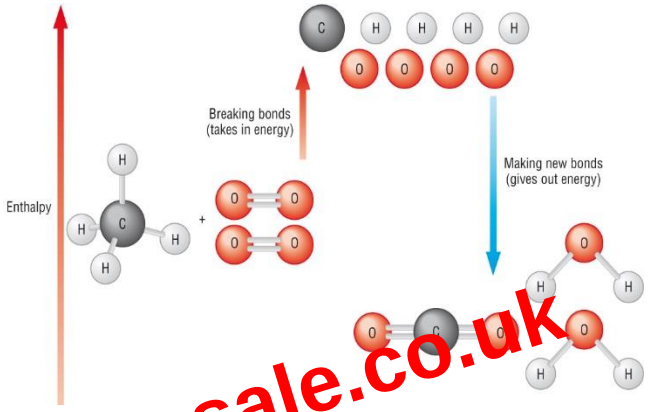
- Depends of bond enthalpies (relative strengths of bonds being broken or made)
- Exothermic reaction – bonds formed are stronger than the bonds that are broken
- Endothermic reaction – bonds broken are stronger than the bonds that are formed

BOND ENTHALPIES TO DETERMINE ENTHALPY CHANGES

- For reactions involving simple gaseous molecules
- $\Delta H = \Sigma$(bond enthalpies of bonds broken) – Σ(bond enthalpies of bonds made)
- Σ(bond enthalpies of bonds broken) = energy required to break bonds
- Σ(bond enthalpies of bonds made) = energy released when bonds are made

BOND ENTHALPIES TO DETERMINE ENTHALPY CHANGES (EXAMPLE)

- Average bond enthalpies is used to work out the enthalpy change of reaction for reactions involving gases
- CH_4 (g) + $2O_2$ (g) → CO_2 (g) + $2H_2O$ (g)
- Average bond enthalpies:
  - C–H: +413 kJ mol^{-1}; O=O: +497 kJ mol^{-1}; C=O: +805 kJ mol^{-1}; O–H: +463 kJ mol^{-1}
- Bonds broken = 4 (C–H) + 2 (O=O)
- Bonds made = 2 (C=O) + 4 (O–H)
- $\Delta H = \Sigma$(bond enthalpies of bonds broken) – Σ(bond enthalpies of bonds made)
- ΔH = [(4 x 413) + (2 x 497)] – [(2 x 805) + (4 x 463)]
- ΔH = -816 kJ mol^{-1}



HESS' LAW

- Hess' law states that, if a reaction can take place by more than one route and the initial and final conditions are the same, the total enthalpy change is the same for each route
- Method for finding an enthalpy change indirectly
- Route 1: A (reactants → products)
- Route 2: B + C (reactants → intermediate →product)
- Total enthalpy change is the same for each route (by Hess' law)
- A = B + C