WJEC (Eduqas) Chemistry A-level

PI4 - Energy Changes

Definitions and Concepts

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PI4.1 - Enthalpy Changes for Solids and Solutions

Atomisation enthalpy ($\Delta_{at}H$): The energy required for the formation of one mole of gaseous atoms from an element in its standard state under standard conditions.

Born-Haber cycle: Calculates the lattice enthalpy by applying Hess's law and comparing the standard enthalpy change of formation of the ionic compound to values such as ionisation energy and electron affinity.

Endothermic reaction: A reaction which takes in energy from the surroundings so the temperature of the surroundings decreases. In an endothermic reaction, the energy needed to break bonds in the reactants is greater than the energy released from forming new bonds in the products.

Enthalpy of hydration: The enthalpy change when one mole of gaseous ions is dissolved in water to form one mole of aqueous ions under standard conditions.

Enthalpy of solution: The enthalpy change when one mole of solute is dissolved in water to infinite dilution so that the ions no longer interact under standard conditions.

Exothermic reaction: A reaction which releases energy to the surroundings so the temperature of the surroundings increases. In an exothermic reaction, the energy needed to break existing bonds is less than the energy released from forming new bonds.

Ionic compound: A compound made up of anions and cations which are held together by ionic bonds, which arise due to the electrostatic attraction between oppositely charged ions. These structures are neutral overall.

Lattice dissociation enthalpy: The enthalpy change when one mole of a solid ionic compound is completely dissociated into its gaseous constituent ions under standard conditions.This is an endothermic process.

Lattice formation enthalpy: The enthalpy change when one mole of a solid ionic compound is formed from its gaseous constituent ions under standard conditions. This is an exothermic process.

Solubility: The ability of a given substance to dissolve in a solvent.

PI4.2 - Entropy and Feasibility of Reactions

Endothermic reaction: A reaction that takes in energy from the surroundings. The energy of the products is higher than the reactants.

Entropy: A measure of the disorder of a system. The units of entropy are $JK^{-1}mol^{-1}$. On a molecular level, gases are more disordered than liquids, which are more disordered than solids. A reaction that produces a greater number of molecules than the number of reactants molecules will have a positive entropy change, as more random arrangements of these molecules will exist, in other words, the system will become more disordered.

Entropy change: This can be calculated by finding the difference between the standard entropies of the products and the reactants:

$$\Delta S_{total} = \Sigma \Delta S_{products} - \Sigma \Delta S_{reactants}$$

If the entropy change for a reaction is positive, the products are more disordered than the reactants. If the entropy change for a reaction is negative, the products are less disordered than the reactants.

Exothermic: A reaction that releases energy to the surroundings. The energy of the reactants is higher than the products.

Feasible reaction: For a reaction to be feasible at a given temperature it must occur spontaneously. This means no extra energy is required for the reaction to occur.

Gibbs free-energy change: A measure of the feasibility of a chemical reaction. This equation relates Gibbs free energy to enthalpy change, entropy change and temperature. A negative Gibbs free energy change means the reaction is feasible and a positive Gibbs free energy change means it is not feasible.

$$\Delta G = \Delta H - T\Delta S_{system}$$