Higher Still
Notes
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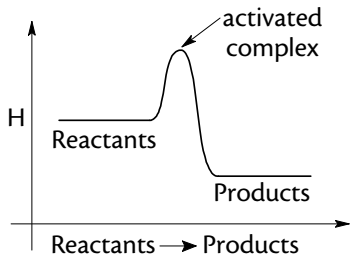
Higher Chemistry

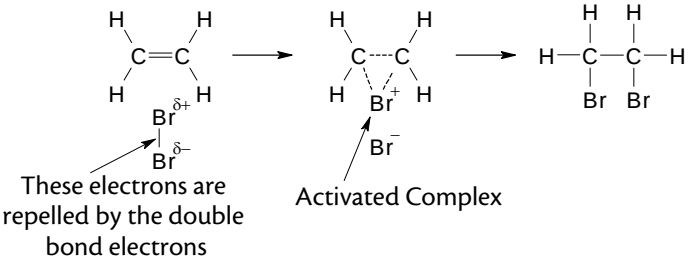
HSN14140
Unit 1 Vocabulary

Vocabulary

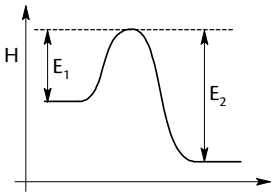
Activated complex



The high energy species formed during a reaction as existing bonds break and new bonds form.



Activation energy

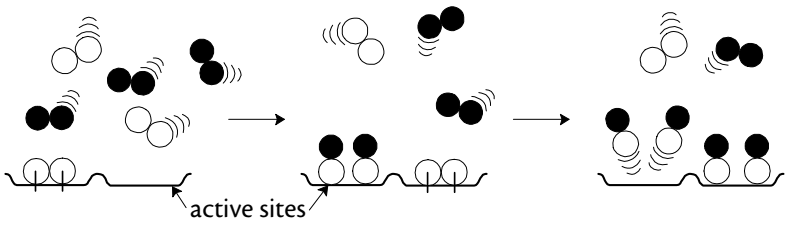


The energy required to cause a reaction to take place – the minimum energy required by reactant molecules to form the activated complex.

E_1 is the forward activation energy
E_2 is the backward activation energy.

Adsorbed

In a heterogeneous catalyst, reactant molecules are adsorbed at active sites on the surface; they form temporary bonds with the catalyst, making it easier for the reaction to take place (at a lower energy).



Avogadro's constant

Given the symbol L, this is the number 6.02×10^{23}. It is defined as the number of atoms in 12g of carbon-12 (ie 1 mole).

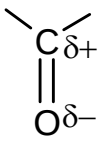
This number is often used in mole calculations.

Catalyst poisoning

When other molecules are preferentially adsorbed on the catalyst's active sites, preventing the reactant molecules from doing so.

Dipole

An unequal distribution of charge can cause a dipole to establish; one part becoming positive, the other negative.



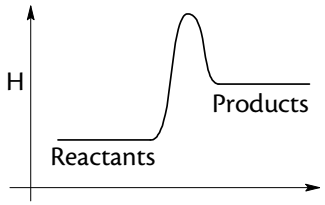
Dipoles can cause permanent dipole-permanent dipole interactions in substances, which have the same effect as hydrogen bonds, but are slightly weaker.

Electronegativity

A measure of an atom's attraction for the shared electrons of a bond; a large difference in electronegativities producing the most polar (or ionic) bonds. Trends in electronegativity:

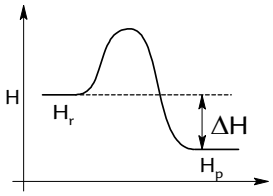
Increases along a period; decreases down a group

Endothermic reaction



One in which the products have higher enthalpy than the reactants; the reaction takes in energy (usually in the form of heat).

Enthalpy change



The difference in the enthalpies of the reactants and products:

$$\Delta H = H_p - H_r$$

For an exothermic reaction, ΔH is negative; for an endothermic reaction ΔH is positive.

Enthalpy of combustion

The enthalpy change when 1 mole of a substance burns completely in oxygen, eg:

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(l)} \qquad \Delta H = -892 \text{ kJ mol}^{-1}$$

Enthalpy of neutralisation

The enthalpy change when 1 mole of water is produced by a neutralisation reaction, eg:

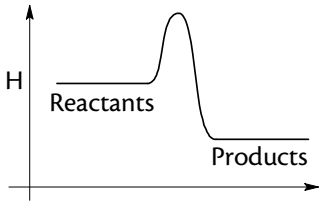
$$H^+_{(aq)} + OH^-_{(aq)} \longrightarrow H_2O_{(l)} \qquad \Delta H = -57.3 \text{ kJ mol}^{-1}$$

Enthalpy of solution

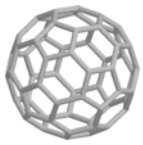
The enthalpy change when 1 mole of a substance dissolves completely in water, eg:

$$NaCl_{(s)} \longrightarrow Na^+_{(aq)} + Cl^-_{(aq)} \qquad \Delta H = +5 \text{ kJ mol}^{-1}$$

Exothermic reaction



One in which the products have a lower enthalpy than the reactants; the reaction gives out energy.



Fullerenes

A recently discovered third form of carbon, existing as large discrete molecules (eg C_{60} or C_{70}). Fullerenes are the subject of a great deal of research as applications are sought.

Heterogeneous catalyst

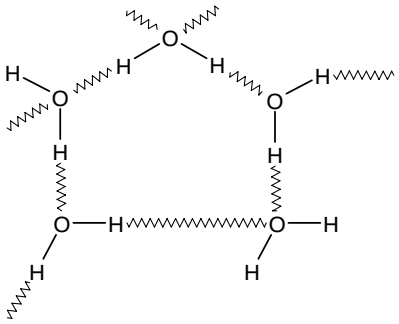
A catalyst in a different physical state from the reactants, examples are iron in the Haber process and catalytic converters in cars. See adsorption for an explanation of how they work.

Homogeneous catalyst

A catalyst in the same physical state as the reactants.

Hydrogen bonds

An intermolecular force which occurs between molecules in which a hydrogen atom is directly bonded to an atom of nitrogen, oxygen or fluorine. Water is an interesting example, as hydrogen bonding is responsible for ice floating on water – the hydrogen bonds give ice a very regular structure.



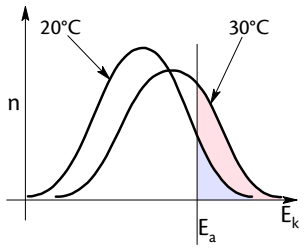
Ionisation enthalpy

The first ionisation enthalpy is the energy required to remove a mole of electrons from a mole of the element's atoms in the gas state. The second and third ionisation enthalpies are the energies required to remove subsequent moles of electrons

1st ionisation enthalpy $Na_{(g)} \longrightarrow Na^+_{(g)} + e^-$ $\Delta H = +494$ kJ mol^{-1}

Kinetic energy

A form of energy possessed by all moving objects (kinetic energy in a gas can be interpreted as heat). Kinetic energy can be converted to other forms of energy, such as chemical potential energy. Kinetic energy relates to activation energy, since reactant molecules must have this amount of kinetic energy in order to react. This explains the effect of an increase in temperature on reaction rate, as many more molecules will have energy greater than the activation energy.



Molar gas volume

The volume of a mole of a gas, which is equal for all gases at equal pressure and temperature. It is usually about 24 l mol^{-1}, but if you require it in a calculation, you will be told the value in the question.

Monatomic

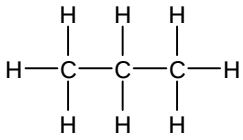
The unreactive noble gases are monatomic – their molecules consisting of a solitary atom. Van der Waals' forces are the only forces which act between the molecules.

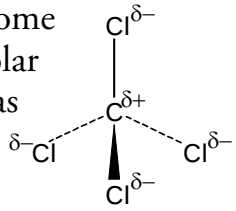
Network structure

Covalently bonded atoms which are not discrete molecules form covalent network structures; a giant lattice of covalently bonded atoms. Because the bonds holding the atoms together are covalent, these substances have very high melting points, and are often used as abrasives because of their strength. Examples include diamond, silicon dioxide, silicon carbide (carborandum) and boron nitride.

Non-polar molecule

Covalently bonded atoms which have identical, or very similar, electronegativities will result in a non-polar molecule, as the distribution of the electrons is equal. An example is the alkanes since the C-H bonds are non-polar. Some molecules contain polar bonds, but are still non-polar molecules, because the polarities of the bonds cancel out, as in tetrachloromethane.





Non-polar solvent

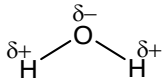
A non-polar substance which can be used to dissolve other non-polar molecules.

Photochemical reactions

Reactions whose activation energy is supplied by light, such as photography:

$$Ag^+ + e^- \longrightarrow Ag_{(s)}$$

Polar molecule



One in which the polarities of polar bonds do not cancel, giving the molecule an overall polarity.

Polar solvent

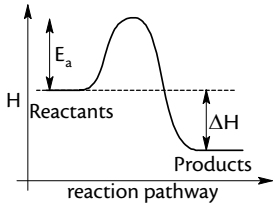
A substance consisting of polar molecules, which can be used to dissolve other polar molecules or ionic lattices. The most common solvent is the polar water molecule.

Polar-covalent bonding

A type of covalent bond in which the sharing of the electrons is unequal, due to the electronegativity difference.

$\delta^+ H — I^{\delta-}$

Potential energy diagrams



Graphs which display the chemical potential energy, or 'enthalpy', of the chemicals involved in a reaction. The example is an exothermic reaction.

Van der Waals' Forces

A weak intermolecular force present in all molecular substances, caused by the existence of temporary dipoles. These result from the molecules' electrons momentarily aligning to make a part of the molecule slightly more negative, so the more electrons present in the molecule (ie the bigger the molecule) the bigger the temporary dipole, hence the stronger the van der Waals' forces.