

CfE Higher Chemistry – Unit One - Summary

1) Reaction Rates

a) Collision Theory

- Rates of chemical reactions can be controlled by chemists.
- If reaction rates are too low a manufacturing process will not be economically viable.
- If reaction rates are too high there is a risk of thermal explosion.
- The rates of reactions are affected by changes in concentration, particle size and temperature and the collision theory can be used to explain these effects.
- The relative rate of a reaction can be calculated using the formula Rate = $1/t$.

b) Reaction Profiles

- Graphs which use the same axes and place the results for different experiments in which the concentration or temperature are varied from one experiment to the next are common and show up how that variable affects the reaction progress.
- Temperature is a measure of the average kinetic energy of the particles of a substance.
- Activation energy is the minimum kinetic energy required by colliding particles before reaction can occur.
- Energy distribution diagrams can be used to explain how an increase in temperature or, in some chemical reactions the energy from light, increases the number of particles with energy greater than the activation energy (E_a/EA).
- Reactions increase their rate at higher temperatures because a higher proportion of the molecules involved have energy in excess of the activation energy and more successful collisions can occur.
- It is observed that a 10°C rise is responsible for an approximate doubling of rate in many reactions.
- The effect of temperature on reaction rate can be explained in terms of an increase in the number of particles with energy greater than the activation energy.

c) Catalysis

- In general, the lower the activation energy the faster the reaction.
- Catalysts are substances which speed up chemical reactions without being used up in the process. They are widely used in industrial processes.
- Catalysts allow chemical reactions to occur more quickly at lower temperatures and so reduce energy costs.
- Catalysts work by the adsorption of reactant molecules onto the surface of the catalyst with consequent weakening of the bonds in the reactant molecules.
- Catalytic converters are fitted to cars to catalyse the conversion of poisonous carbon monoxide and oxides of nitrogen to carbon dioxide and nitrogen. Cars with catalytic converters can only use 'lead-free' petrol to prevent poisoning of the catalyst.
- Enzymes catalyse the chemical reactions which take place in the living cells of plants and animals. They are also widely used in industrial processes.
- Cobalt(II) chloride speeds up the reaction between tartrate ions and hydrogen peroxide. Colour changes clearly show that the catalyst takes part in the reaction and is chemically unchanged at the end of the reaction.

c) Catalysis continued

- The activated complex is an unstable arrangement of atoms formed during a reaction at the maximum of a potential energy barrier.
- A potential energy diagram can be used to show the energy pathway for a reaction.
- The enthalpy change, which can be calculated from the potential energy diagram, is the energy difference between products and reactants.
- The enthalpy change has a negative value for exothermic reactions, which cause heat energy to be released to the surroundings.
- The enthalpy change has a positive value for endothermic reactions, which cause absorption of heat energy from the surroundings.
- The activation energy is energy required by colliding molecules to form the activated complex.
- The activation energy can be calculated from potential energy diagrams.
- Catalysts speed up reactions by providing an alternative pathway which has a lower activation energy.
- A potential energy diagram can be used to show the effect of a catalyst on activation energy.

2) The Periodic Table

- Elements are arranged in the Periodic Table in order of increasing atomic number.
- The Periodic Table allows chemists to make accurate predictions of physical properties and chemical behaviour for any element based on its position.
- There are periodic variations in the densities, melting points and boiling points of the elements across a Period and down a Group.

3) Bonding and Structure

- Metallic bonding is the electrostatic force of attraction between positively charged ions and delocalised outer electrons.
- A metallic structure consists of a giant lattice of positively charged ions and delocalised outer electrons.
- Atoms in a covalent bond are held together by electrostatic forces of attraction between positively charged nuclei and negatively charged shared electrons.
- A covalent molecular structure consists of discrete molecules held together by intermolecular forces.
- A covalent network structure consists of a giant lattice of covalently bonded atoms.
- Ionic bonding is the electrostatic force of attraction between positively and negatively charged ions.
- An ionic structure consists of a giant lattice of oppositely charged ions.
- Elements can be categorised into four classes according to their bonding and structure.
 - Metallic
 - Covalent molecular
 - Covalent network
 - Monatomic

4) Periodic Table Trends

- The covalent radius decreases across a Period because the increase in nuclear charge attracts the electrons more strongly.
- The covalent radius increases on going down a group as the number of occupied electron shells increases.
- The first ionisation energy is the energy required to remove one mole of electrons from one mole of gaseous atoms.
- The second and subsequent ionisation energies refer to the energies required to remove further moles of electrons.
- First ionisation energies increase across a Period and decrease down a group.
- This can be explained in terms of atomic size, nuclear charge and the screening effect due to inner shell electrons.
- Electronegativity is a measure of the attraction an atom in a bond has for the electrons of the bond.
- Electronegativity values increase across a Period and decrease down a group.
- This can be explained in terms of atomic size, nuclear charge and the screening effect due to inner shell electrons.
- Electronegativity is a measure of the attraction an atom in a bond has for the electrons of the bond.
- Electronegativity values increase across a Period and decrease down a group.

5) Bonding continuum and polar covalent bonding

- Polar covalent bonds occur when the atoms of the bond attract the bonding electrons unequally causing the atoms to have partial positive and negative charges.
- The polarity of a covalent bond depends on the difference in electronegativity between the bonded atoms, the most electronegative becoming more negative.
- Between pure covalent and pure ionic bonds there are polar covalent bonds.
- The type of bonding between two atoms depends mainly on the difference in electronegativity between the atoms.
 - When the difference is zero, the bond will be covalent.
 - With a small difference, a polar covalent bond is likely.
 - When the difference is large, fully charged ions are produced and ionic bonding will be predicted.
- Permanent dipole-permanent dipole interactions act in addition to London dispersion electrostatic attractions between polar molecules and are stronger than these attractions for molecules of equivalent size.
- Not all covalent molecules with polar bonds result in polar molecules.
- Molecules which are highly symmetrical tend to be non-polar.
- An electrostatically charged rod can be used to detect the presence of polar molecules in a liquid. Polar molecules are attracted to both a negative and positive rod.
- There is a complete range of bond types leading to a bonding spectrum mainly based on electronegativity.

6) Intermolecular Forces

- All molecular elements and compounds and monatomic elements condense and freeze at sufficiently low temperatures. For this to occur, some attractive forces must exist between the molecules or discrete atoms.
- Intermolecular forces acting between molecules are known as van der Waals' forces.
- There are several different types of van der Waals' forces such as London dispersion forces and permanent dipole: permanent dipole interactions, which include hydrogen bonding.

a) London Dispersion Forces

- London dispersion forces are forces of attraction that can operate between all atoms and molecules.
- These forces are much weaker than all other types of bonding.
- They are formed as a result of electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules.
- The strength of London dispersion forces is related to the number of electrons within an atom or molecule.

b) Permanent Dipole: Permanent Dipole Interactions

- Bonds consisting of a hydrogen atom bonded to an atom of a strongly electronegative element such fluorine, oxygen or nitrogen are highly polar.
- Hydrogen bonds are electrostatic forces of attraction between molecules which contain these highly polar bonds.
- A hydrogen bond is stronger than other forms of permanent dipole-permanent dipole interaction but weaker than a covalent bond.
- Melting points, boiling points and viscosity can all be rationalised in terms of the nature and strength of the intermolecular forces which exist between molecules.
- By considering the polarity and number of electrons present in molecules, it is possible to make qualitative predictions of the strength of the intermolecular forces.
- The melting and boiling points of polar substances are higher than the melting and boiling points of non-polar substances with similar numbers of electrons.
- The anomalous boiling points of ammonia, water and hydrogen fluoride are a result of hydrogen bonding.
- Boiling points, melting points, viscosity and solubility/miscibility in water are properties of substances which are affected by hydrogen bonding.
- Hydrogen bonding between molecules in ice results in an expanded structure which causes the density of ice to be less than that of water at low temperatures.
- Ionic compounds and polar molecular compounds tend to be soluble in polar solvents such as water and insoluble in non-polar solvents.
- Non-polar molecular substances tend to be soluble in non-polar solvents and insoluble in polar solvents.