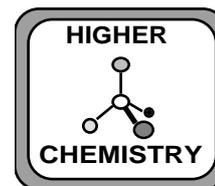


Unit 1 - Energy Matters



Reaction Rates

1. Reactions can be followed by measuring changes in concentration, mass and volume of reactants and products.
2. The average rate of a reaction, or stage in a reaction, can be calculated from initial and final quantities and the time interval.
3. The rate of a reaction, or stage in a reaction, is proportional to the reciprocal of the time taken.
4. The rates of reaction are affected by changes in concentration, particle size and temperature.
5. The collision theory can be used to explain the effects of concentration and surface area on reaction rate.
6. Temperature is a measure of the average kinetic energy of the particles of a substance.
7. The activation energy is the minimum kinetic energy required by colliding particles before reaction will occur.
8. Energy distribution diagrams can be used to explain the effect of changing temperature on the kinetic energy of the particles.
9. 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.
10. With some chemical reactions, light can be used to increase the number of particles with energy greater than the activation energy.
11. The reactant that is in excess can be calculated.
12. Catalysts can be classified as either heterogeneous or homogeneous.
13. Catalysts are used in many industrial processes.
14. Heterogeneous catalysts work by the adsorption of reactant molecules.
15. The surface activity of a catalyst can be reduced by poisoning.
16. Impurities in the reactants result in industrial catalysts having to be regenerated or renewed.
17. Catalytic convertors are fitted to cars to catalyse the conversion of poisonous carbon monoxide and oxides of nitrogen to carbon dioxide and nitrogen.
18. Cars with catalytic convertors use only unleaded petrol to prevent poisoning of the catalyst.
19. Enzymes catalyse the chemical reactions which take place in the living cells of plants and animals.
20. Enzymes are used in many industrial processes.

Enthalpy

1. Exothermic changes cause heat to be released to the surroundings; endothermic changes cause absorption of heat from the surroundings.
2. A potential energy diagram can be used to show the energy pathway for a reaction.
3. The enthalpy change is the energy difference between products and reactants.
4. The enthalpy change can be calculated from a potential energy diagram.
5. The enthalpy change has a negative value for exothermic reactions and a positive value for endothermic reactions.
6. The activated complex is an unstable arrangement of atoms formed at the maximum of the potential energy barrier.
7. The activation energy is the energy required by colliding molecules to form an activated complex.
8. The activation energy can be calculated using potential energy diagrams.
9. The effect of a catalyst can be explained in terms of alternative reaction pathways with lower activation energy.
10. A potential energy diagram can be used to show the effect of a catalyst.
11. Enthalpy of combustion is the enthalpy change when one mole of a substance burns completely in oxygen.
12. Enthalpy of solution is the enthalpy change when one mole of a substance dissolves completely in water.
13. Enthalpy of neutralisation of an acid is the enthalpy change when the acid is neutralised to form one mole of water.
14. Enthalpy changes can be calculated from $cm\Delta T$.

Periodic Table

1. The modern Periodic Table is based on the work of Mendeleev who arranged the known elements in order of increasing atomic masses in conjunction with similar chemical properties, leaving gaps for the yet to be discovered elements.
2. There are variation in the densities, melting points and boiling points of the elements across a period and down a group.
3. The atomic size decreases across a period and increases down a group.
4. The first ionisation energy is the energy required to remove one mole of electrons from one mole of gaseous atoms.
5. The second and subsequent ionisation energies refer to the energies required to remove further moles of electrons.
6. The trends in the first ionisation energy across periods and down groups can be explained in terms of the atomic size, nuclear charge and the screening effect due to inner shell electrons.

7. Atoms of different elements have different attractions for bonding electrons.
8. Electronegativity is a measure of the attraction an atom involved in a covalent bond has for the shared electrons of the bond.
9. Electronegativity values increase across a period and decrease down a group.

Bonding Structure and Properties

1. Metallic bonding is the electrostatic force of attraction between positively charged ions and delocalised outer electrons.
2. Atoms in a covalent bond are held together by electrostatic forces of attraction between positively charged nuclei and negatively charged shared electrons.
3. The polarity of a covalent bond depends on the difference in electronegativity between the bonded atoms.
4. Ionic bonding is the electrostatic force of attraction between positively and negatively charged ions.
5. The type of bonding in a compound is related to the positions of its constituent elements on the Periodic Table.
6. Van der Waals forces are forces of attraction which can operate between all atoms and molecules.
7. Van der Waals forces are much weaker than all other types of bonds.
8. Van der Waals forces are as a result of electrostatic attraction between temporary dipoles and induced dipoles caused by movement of electrons in atoms and molecules.
9. The strength of Van der Waals forces is related to the size of the atoms and molecules.
10. A molecule is described as polar if it has a permanent dipole.
11. Permanent dipole-permanent dipole interactions are stronger than Van der Waals forces for molecules of equivalent size.
12. The spatial arrangement of polar covalent bonds can result in a molecule being polar.
13. Bonds consisting of a hydrogen atom bonded to an atom of a strongly electronegative atom such as nitrogen, oxygen or fluorine are highly polar.
14. Hydrogen bonds are electrostatic forces of attraction between molecules which contain these highly polar bonds.
15. A hydrogen bond is stronger than the other forms of permanent dipole-dipole interaction but weaker than a covalent bond.
16. A metallic structure consists of a giant lattice of positively charged ions and delocalised outer electrons.
17. A covalent molecular structure consists of discrete molecules held together by intermolecular forces.
18. A covalent network structure consists of a giant lattice of covalently bonded atoms.
19. An ionic structure consists of a giant lattice of oppositely charged ions.

20. A monatomic structure consists of discrete atoms held together by Van der Waals forces.
21. The first 20 elements in the Periodic Table can be categorised according to bonding and structure:
metallic (Li, Be, Na, Mg, Al, K, Ca)
covalent molecular (H_2 , N_2 , O_2 , F_2 , Cl_2 , P_4 , S_8 and C (fullerenes))
covalent network (B, C (diamond and graphite), Si)
monatomic (noble gases)
22. Compounds each adopt one of three structures in the solid state:
covalent molecular
covalent network (including silicon dioxide and silicon carbide)
23. The melting points, boiling points and hardness/softness of elements and compounds are related to their bonding and structures.
24. The melting and boiling points of polar substances are higher than the melting and boiling points of non-polar substances of similar molecular size.
25. Ionic compounds and polar molecular compounds tend to be soluble in polar solvents such as water, and insoluble in non-polar solvents.
26. Non-polar molecular substances tend to be soluble in non-polar solvents and insoluble in polar solvents.
27. The anomalous boiling points of ammonia, water and hydrogen fluoride are as a result of hydrogen bonding.
28. Boiling points, melting points, viscosity and miscibility in water are properties of substances that are affected by hydrogen bonding.
29. 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.
30. The uses of diamond, graphite and silicon carbide are related to their structures and properties.
31. Fullerenes are the subject of current research and applications are being sought.

The Mole

1. One mole of any substance contains 6.02×10^{23} formula units.
2. Equimolar amounts of substances contain equal numbers of formula units.
3. The molar volume (in units of $dm^3 mol^{-1}$) is the same for all gases at the same temperature and pressure.
4. The volume of a gas can be calculated from the number of moles and vice versa.
5. The volumes of reactant and product gases can be calculated from the number of moles of each reactant and product.