# 3.1 - The periodic table

## 3.1.1 - Periodicity

|  |  |
| --- | --- |
| **What is periodicity?** | A repeating pattern across different periods. |
| **Why do P4 and S8 have high boiling and melting points?** | They have more electrons ∴ more London forces which require more energy to overcome. |
| **Define first ionisation energy** | Energy required to remove an electron from each atom (1) in one mole of gaseous atoms (1). |
| **Why are there sharp and slight increases in successive ionisation energies?** | * Slight increases ∵ removing an electron ⇒ smaller atomic radius (∵ less repulsion between subshells) ⇒ stronger attraction. * Sharp increase ∵ less shielding and smaller atomic radius ⇒ stronger attraction. |
| **Why do electronegativites and ionisation energies decrease down groups?** | * Increased atomic radius and more shells ⇒ more shielding ⇒ less attraction ⇒ easier to remove. * Increased shielding outweighs the increase in proton no. and thus nuclear charge. |
| **Why does electronegativity and ionisation energy increase across periods?** | * Greater nuclear charge ⇒ reduced atomic radius ⇒ greater attraction ⇒ harder to remove * During this, shielding stay the same. |
| **Why is boron (1s22s22p1)’s first ionisation energy less than beryllium (1s22s2)’s?** | p-orbital e- has a higher energy than the s-orbital e- ∴ easier to remove.  *It’s all about a new factor now having an effect relative to the previous element.* |
| **Why is oxygen (1s22s22p4)’s first ionisation energy less than nitrogen (1s22s22p3)’s?** | Electrons now pair in the same orbital ⇒ repel one another ⇒ easier to remove. |
| **What is a metallic bond?** | Strong electrostatic force of attraction (1) between cations and delocalised electrons (1). |
| **What 3 things are metallic bonds affected by?** | * No. of protons (the more, the stronger). * No. of delocalised electrons per atom (the more, the stronger). * No. of shells (the more, the weaker ∵ electron shielding). |
| **Why are metals malleable?** | Layers of cations can easily slide over each other. |
| **What shape do group 4 elements form?** | Tetrahedral. |
| **Give 3 examples of giant covalent molecules** | * Diamond. * Graphite. * Silicon dioxide (SiO2). |
| **Give 2 reasons for why graphite is soft** | * Layers can easily slide of each other. * Weak London forces. |

## 3.1.2 - Group 2

|  |  |
| --- | --- |
| **Why does the reactivity increase down group 2?** | First ionisation energy decreases ⇒ less energy required to remove the outermost electron. |
| **Why does reactivity decrease down group 7?** | * Increased atomic radius and increased shielding due to more shells ⇒ less attraction between the nucleus ⇒ gains electrons less readily. * The increase in proton number and thus **nuclear charge** is outweighed by an increase in shielding. |
| **How are group 2 compounds used in farming?** | * To neutralise acidic soils. * E.g., CaO, Ca(OH)2, or CaCO3. |
| **Give a group 2 compound used in medicine, how it's used, and why it's used** | * Mg(OH)2 (milk of magnesia) is used to neutralise excess stomach acid ∵ not too strong an alkaline * E.g., |
| **Give 3 observations of a group 2 element in acid** | 1. Fizzing (more vigorous, further down). 2. Metal (ribbon) dissolving (faster, further down). 3. Solution heating (more, further down). |
| **What changes for group 2 hydroxides down the group and why?** | Solubility increases and alkalinity increases ∵ more HO- ions are released ∴ higher pH. |
| **What is decomposition (with an example equation)?** | * The separation of a single compound into two or more elements/compounds. * 2ClO2 → Cl2 + 2O2. |
| **What changes for group 2 carbonates down the group?** | Ease of thermal decomposition decreases. |

## 3.1.3 - Group 7

|  |  |
| --- | --- |
| **What is a halide?** | A binary compound of a halogen. |
| **Give the colours and states of the first 4 halogens at room temperature** | |  |  |  | | --- | --- | --- | | **Name** | **Colour** | **State** | | F2 | Pale yellow | Gas | | Cl2 | Green | Gas | | Br2 | Red-brown (like bromine water) | Liquid | | I2 | Grey | Solid |   You Guys Read Gore.  YELLOW GREEN RED GREY (down the group). |
| **Why must experiments with halogens be done in a well-ventilated area?** | They’re powerful oxidising agents ∴ toxic. |
| **How will Br2 and I2 appear when in cyclohexane compared to water?** | * Br2 is orange in both. * I2 is orange/brown in water and violet in cyclohexane.     *Br2 being orange can be remembered by bromine water being orange.* |
| **What is displacement of group 7 elements (with an example)?** | * More reactive halogen displacing less reactive halide from its aqueous solution. * E.g., |
| **What is disproportionation and what can it happen with?** | * A reaction where same element is both oxidised and reduced. * Any halogen (except fluorine ∵ very powerful oxidising agent). |
| **Why will chlorine and water bleach universal indicator?** | The ClO- ion from the HClO(aq) produced is a powerful oxidising agent. |
| **Give 1 pro and 1 con of adding Cl2 to water** | * It kills bacteria (1) * It could form chlorinated hydrocarbons which are toxic/carcinogenic (1) |
| **How does chlorine react with water?** | This is a disproportion reaction. |
| **How is bleach formed, why this way, and what is used as the bleach?** | * By reacting chlorine with **cold dilute sodium hydroxide solution** as it has a higher solubility. * The mixture of NaClO is used as bleach. The NaClO is just called sodium chlorate (I). |

## 3.1.4 - Qualitative analysis

|  |  |
| --- | --- |
| **What is the test for carbonates?** | 1. Add nitric acid. If you see effervescence, it **COULD** contain a carbonate ∵ a gas is produced. 2. Bubble this gas through limewater. If cloudy then gas was CO2 and thus carbonate was present. |
| **Give the test for sulfate ions** | 1. Add nitric acid to dissolve any carbonates to prevent the formation of BaCO3 precipitate (aka a false positive which can mask desired observations). 2. Add barium nitrate solution (or anything with Ba2+). 3. If white ppt forms, you have a sulfate present.   *This is because Ba2+(aq) + SO42-(aq) → BaSO4 (s)* |
| **Give the test for halides with a precaution and example equation** | 1. Add nitric acid to dissolve any carbonates to prevent the formation of Ag2CO3 precipitate (aka a false positive which can mask desired observations). 2. Add silver nitrate solution. 3. If the precipitate formed is **white** it’s silver **chloride**, if it’s **cream** then it’s silver **bromide**, and if it’s **yellow** then it is silver **iodide**.   *To remember it, learn the halides in alphabetical order* ***BCI*** *then the colours in alphabetical order* ***CWY****.* |
| **How can you be certain of the result of the halide test?** | 1. Add ammonia solution. 2. Silver **chloride** dissolves under dilute solution. Silver **bromide** dissolves under concentrated solution. Silver **iodide** will not dissolve at all. |
| **What is the test for ammonium ions?** | * Add sodium hydroxide solution and **HEAT GENTLY**. * If NH3 is produced then you can…   + Smell for it.   + Use damp litmus paper and hover it over the tube. This will turn blue under NH3 as NH4OH will form on the paper. |
| **Why should you not use certain acids when performing qualitative analysis?** | * As the acid releases the ions you're testing into the solution so you will **OBVIOUSLY** get a positive test for follow-up tests. * Only use these if you plan to perform **ONE TEST / NO FOLLOW-UP TESTS.**   *If you use HCl or H2SO4 then you will obviously get a positive test for SO4 and Cl2.* |
| **What is a precipitation reaction (with example)?** | When cations and anions in aqueous solution combine to form an insoluble ionic solid.  E.g., |
| **In what order should you carry out anion tests? Why?** | * Carbonate - Sulfate - Halide (Ca-S-H). * Continue if it is negative. * Adding Ba2+ to CO32- produces a white ppt of BaCO3. * Adding Ag+ to CO32- produces a yellow-grey ppt of Ag2CO3. * Adding Ag+ to SO42- produces a white ppt of Ag2SO4. |

# 3.2 - Physical chemistry

## 3.2.1 - Enthalpy changes

|  |  |
| --- | --- |
| **Why can ‘enthalpy change of formation’ not be measured directly?** | Many different compounds can form from the same constituent elements. |
| **How is enthalpy change calculated?** | The plus or minus depends on whether the reaction was exothermic or endothermic. |
| **Breaking bonds is … because ...** | * Endothermic. * Energy is put into the system.   *Atoms become less stable as they have no bonds and lots of energy.* |
| **Making bonds is … because ...** | * Exothermic. * Energy is released to the surroundings.   *Despite the little energy needed to make a bond, the molecule becomes more stable when bonded and thus releases energy.* |
| **What is ΔH°r equal to? (using bond enthalpies)** | ΔH°r = ΣΔH(bonds broken) - ΣΔH (bonds formed) |
| **What is activation energy defined as?** | The minimum energy required of particles to collide to start a reaction. |
| **Draw and label the energy profile diagram for an exothermic reaction** | ΔH is negative as energy is **EX**iting the system to surroundings and thus **EX**othermic. Taking up to be the positive direction, we have the arrow pointing downwards. |
| **Draw and label the energy profile diagram for an endothermic reaction** | ΔH is positive as energy is **EN**tering the system from the surroundings and thus **EN**dothermic. Taking up to be the positive direction, we have the arrow pointing downwards. |
| **What does Hess’ Law state?** | The enthalpy change of reaction is independent of the route taken. |
| **What 2 things are required of particles to react?** | * Minimum kinetic energy (i.e., activation energy). * Correct orientation. |
| **How does concentration affect the rate of reaction?** | Increasing concentration ⇒ more particles per unit volume (1) ⇒ more frequent successful collisions (1) ⇒ increased rate of reaction (1).  *If the question explicitly says double concentration, say double particles and thus double frequency of successful collisions.* |
| **How does temperature affect the rate of reaction?** | Increasing temperature ⇒ higher % of particles with KE greater than EA (1) ⇒ more frequent successful collisions (1) ⇒ increased rate of reaction (1). |
| **How does surface area affect the rate of reaction?** | Increasing surface area ⇒ more sites are exposed (1) ⇒ more frequent successful collisions (1) ⇒ increased rate of reaction (1). |
| **What are the limitations of calorimetry?** | * Heat loss to surroundings. * Non-standard conditions. * Heat capacity of container not included. * Contents of calorimeter may evaporate. |

## 3.2.2 - Reaction rates

|  |  |
| --- | --- |
| **Give 2 ways of improving an experiment that uses a gas syringe** | 1. Let the reaction carry on until the gas syringe has stopped moving. 2. Wait for the gas to cool to room temperature before reading. |
| **What do catalysts do?** | Provide an alternate reaction pathway with a lower activation energy without being used up. |
| **Give 2 reasons for why catalysts are useful** | 1. Allow reactions to occur at lower temperatures and pressure ∴ less (thermal) energy demand. 2. Allow for reactions with better atom economies ∴ less waste. |
| **What is a homogeneous catalyst and its benefits and drawbacks?** | * A catalyst in the same phase/state as the reactants. * It means more molecules can react. * It makes separating molecules at the end harder. |
| **What is a heterogeneous catalyst and its benefits and drawbacks?** | * A catalyst in a different phase/state to the reactants. * Can be easily separated. * Reactions only take place on the catalyst’s surface. |
| **Describe what occurs on the surface of the nickel catalyst in the hydrogenation of ethene** | 1. The reactants (H2 and C2H4) are **ADSORBED** onto the catalyst surface. 2. The bonds weaken (1) and they react to form C2H6. 3. The product (C2H6) is **DESORBED** from the surface.   Thus…  C2H4 + H2 → C2H6. |
| **What is the key idea behind the ‘Boltzmann Distribution’?** | Particles gain or lose energy in collisions with others and thus travel at different speeds. |
| **Give the key features of the ‘Boltzmann Distribution’** | * It starts at the origin because no particles have no energy. * The peak (Emp) is the most probable energy (which is different from the mean). * It never meets the x-axis as there is no maximum energy for particles.     *The area under the graph represents the number of molecules.* |
| **What does the ‘Boltzmann Distribution’ look like at different temperatures and why?** | Increasing the temperature shifts the distribution towards more particles having higher energies (more than any EA) **AND** wider range of energies due to more frequent collisions.    *The area under the graph represents the number of molecules.*  The x-axis is also energy / kinetic energy of the molecules. |

## 3.2.3 - Chemical equilibrium

|  |  |
| --- | --- |
| **What is a closed system?** | A container in which nothing cannot enter or leave except for heat. |
| **What is dynamic equilibrium?** | * When the forward and backward reactions are occurring at the same rates. * Thus, the concentrations stay constant. Yet, the amounts of each may differ. |
| **What is Le Chatelier’s Principle?** | If a dynamic equilibrium is changed, the position of the equilibrium to oppose this change. |
| **What happens to the equilibrium position if the temperature is increased and why?** | * Favours the endothermic side. * To minimise the effect of temperature rise by absorbing heat. |
| **What happens to the equilibrium position if pressure is increased?** | * Favous the side producing the least moles of gas (1). * To minimise the effect of pressure increase by reducing the number of particles.   *Remember that each mole takes 24dm3 under standard conditions!* |
| **What compromises must be made when altering the equilibrium position and why?** | **Pressure:**   * A lower pressure when favouring the side with less moles ∵ high pressures are difficult to produce (high electrical costs in pumping gas) and difficult to contain (to withstand high pressures). * A higher pressure when favouring side with more moles otherwise the rate of reaction will be too low.   **Temperature:**   * A higher temperature when favouring the exothermic side otherwise the rate of reaction will be too low. * A lower temperature when favouring the endothermic side ∵ expensive to use too much fuel. * Because this allows for a reasonable rate of reaction / costs without shifting the equilibrium too far to the left/right. * You will have a smaller equilibrium yield but more product produced in unit time. |
| **What happens to the equilibrium position if a catalyst is added and why?** | * No change as it increases the rate of reaction for both sides. * The time at which equilibrium is achieved is reduced. |
| **What is the equation for the equilibrium constant (Kc)?** | aA + bB ⇌ cC + dD |