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| **UNIT 2**  **PARTICLES, BONDING AND STRUCTURES**  **Answers** |

***Lesson 1 – What happens when atoms join together?***

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| Image result for test icon**Test your knowledge 1.1: Classifying materials** |
| Answers: (a) element; (b) compound (NH3); (c) mixture; (d) mixture; (e) compound (H2O); (f) compound (CO2); (g) mixture; (h) mixture; (i) element (an allotrope of carbon); (j) compound (SiO2); (k) element; (l) compound (CH4) |

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| Image result for test icon**Test your knowledge 1.2 – Demonstrating the laws of chemical combination** |
| 1. 7.5/2.5 = 3; 30/10 = 3 so consistent (ie composition is always the same) 2. 13.6/6.8 = 2 (a small whole number ratio, so consistent) |

***Lesson 2 – What is ionic bonding?***

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| **Summary Activity 2.1: Which atoms attract electrons most strongly?** |
| * As you cross a Period proton number increases but shielding stays the same, so electrons in the outer shell are more strongly attracted to the nucleus * As you descend a Group the number of shells increases so shielding increases, so electrons in the outer shell are less strongly attracted to the nucleus |

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| Image result for test icon**Test your knowledge 2.2: Forming simple ions** |
| 1. lithium ion Li+, potassium ion K+, magnesium ion Mg2+, calcium ion Ca2+, aluminium ion Al3+ 2. oxide ion O2-, nitride ion N3-, sulphide ion S2-, bromide ion Br-, phosphide ion P3- |

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| Image result for test icon**Test your knowledge 2.3: Forming simple ionic compounds** |
| **(a); (b) ; (c);**  **(d) ; (e)** |

***Lesson 3 - What is covalent bonding?***

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| Image result for test icon**Test your knowledge 3.1: Forming simple molecules** |
| 1. : H2, H-H; (b) HCl, ; (c) H2O, ; (d) CH4,;   (e) NH3, ; (f) O2,; (g) N2; (h) CO2, |

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| Image result for test icon**Test your knowledge 3.2: Understanding dative covalent bonding** |
| 1. N in NH3 on forms dative covalent bond with B in BH3; N provides both electrons      1. O in H2O forms dative covalent bond with H+; O provides both electrons |

***Lesson 4 - What is a metallic bond?***

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| Image result for test icon**Test your knowledge 4.1: Forming metallic bonds** |
| (a); (b) |

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| **Test your knowledge 4.2: Understanding ionic, covalent and metallic bonding** |
| (a) covalent; (b) metallic; (c) ionic; (d) ionic;  (e) metallic; (f) covalent |

###### *Lesson 5 – What shapes do different molecules have?*

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| **Summary Activity 5.1: Lewis dot structures for different molecules** |
| * Covalent bond: pair of electrons shared between two atoms * Molecule: small group of atoms held together by covalent bonds |

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg **Activity 5.2: Investigate the shapes balloons form when tied together** |
| Equipment needed: 4 balloons  Two inflated balloons should naturally sit 180o apart when tied; three inflated balloons 120o apart and four inflated balloons 109.5o apart, in the shape of a tetrahedron when tied |

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| Image result for test icon**Test your knowledge 5.3: Understanding Molecular Shapes** |
| Linear (180o) Trigonal Planar (120o) Tetrahedral (109.5o) Pyramidal (107o)  H-H  O=O O=C=O  Non-linear (104.5o) linear Tetrahedral (109.5o) Pyramidal (107o) linear Linear (180o) |
| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 5.4: Understanding molecular shapes** |
| Free choice question so no answers available |

***Lesson 6 – What are Intermolecular Forces?***

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| Image result for test icon**Test your knowledge 6.1: Distinguishing between temporary and permanent dipoles** |
| (a) BeCl2 no because dipoles on bonds cancel  (b) BF3 no because dipoles on bonds cancel  (c) SiCl4 no because dipoles on bonds cancel  (d) PCl3 yes because dipoles on bonds don’t cancel  (e) H2S yes because dipoles on bonds don’t cancel  (f) H2 no because the bond is not polar  (g) SiH4 no because the dipoles on bonds cancel  (h) PH3 yes because the dipoles on bonds don’t cancel  (i) O2 no because the bond is not polar  (j) CO2 no because dipoles on bonds cancel  (k) HCl yes because there is only one bond and it is polar |
| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 6.2: Distinguishing between temporary and permanent dipoles** |
| Free choice question so no answers available |

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| Image result for test icon**Test your knowledge 6.3: Understanding factors affecting the strength of Van der Waal’s forces** |
| 1. Ar, because there are more electrons in each atom, and so stronger Van der Waal’s forces 2. C2H6, because there are more electrons in each molecule, so stronger Van der Waal’s forces 3. Br2, because there are more electrons in each molecule, so stronger Van der Waal’s forces 4. CH3CH2CH2CH3, because the surface area of the molecule is larger, so stronger Van der Waal’s forces 5. SO2, because there are more electrons in each molecule, so stronger Van der Waal’s forces, and because it has a permanent dipole |

***Lesson 7 – What is hydrogen bonding?***

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| Image result for test icon**Test your knowledge 7.1: Recognising hydrogen bonding** |
| 1. (i) BeH2 no - no electropositive H; (ii) BH3 no - no electropositive H; (iii) CH4 no - no sufficiently electropositive H; (iv) PH3 no - no sufficiently electropositive H; (v) H2O yes - electropositive H because bonded to O; (vi) H2 no - no electropositive H; (vii) HCl no - no sufficiently electropositive H; (viii) NH3 yes - electropositive H because bonded to N; (ix) HF; yes - electropositive H because bonded to F; (x) CO2 no – no H; (xi) H2S no - no sufficiently electropositive H 2. Water molecules can form strong hydrogen bonds with other water molecules; but CO2 cannot; CO2 will have stronger Van der Waal’s forces between molecules but the effect of hydrogen bonding is more significant |

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| Image result for test icon**Test your knowledge 7.2: Using the Kelvin temperature scale** |
| 1. (i) 298 K; (ii) 308 K; (iii) 573 K; (iv) 223 K, (v) 73 K 2. (i) -223 oC; (ii) 20 oC; (iii) -73 oC; (iv) 35 oC; (v) 327 oC |

***Lesson 8 – What are solids and what are the properties of solids?***

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 8.1: Warm ice, sulphur and iodine** |
| Equipment needed: clay pipe triangle, three crucibles, tripod, Bunsen burner and gas, tongs, a block of ice, a spatula of sulphur powder, a spatula of iodine crystals  If possible, warm the sulphur and the iodine in a fume cupboard; if not, use an open space and keep the students a safe distance from the crucible  Place a clay pipe triangle on a tripod over a Bunsen burner connected to gas; take the ice block and place it into a crucible; then place the crucible into the clay pipe triangle; light the Bunsen burner and heat the crucible gently (medium flame) until it starts to melt; then remove the crucible, replace it with a crucible containing one spatula of sulphur, and repeat the warming process; then remove the crucible, replace it with a crucible containing one spatula of iodine, and repeat the warming process |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 8.2: Determine the melting point of naphthalene** |
| Equipment needed: thermometer with range up to 100 oC, capillary tube sealed at one end, a small quantity (less than 1 g) of finely powdered naphthalene, a spatula, a clamp, stand and boss, a 250 cm3 beaker half-filled with water, a tripod, gauze, Bunsen burner and access to gas  Fill the capillary tube with 1 cm of naphthalene and strap it to a thermometer using an elastic band; place a gauze onto a tripod over a Bunsen burner connected to a gas supply; half-fill the 250 cm3 beaker with water; use the boss to attach the clamp to the stand; place the thermometer into the beaker at a depth so that the napthalene is submerged in the water but the top of the capillary tube is not; fix the thermometer in place with the clamp; light the Bunsen burner and heat the water on a medium flame until the naphthalene starts to melt  The melting point of naphthalene is 80 oC |

***Lesson 9 – What are liquids and what are their properties?***

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 9.1: Heat water** |
| Equipment needed: gauze, tripod, Bunsen burner and gas, tongs, a 250 cm3 beaker, access to water  Pleace the gauze on the tripod; place the Bunsen burner under the tripod; half-fill the beaker with water; place the beaker on the gauze; light the Bunsen burner; heat the water on a medium flame   * as the water is heated, small bubbles start to form due to evaporation; the number of bubbles increases as the water gets hotter; then the bubbles start to get very large * the water is boiling and the bubbles become large * some of the water evaporates before it boils |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 9.2: Determine the boiling point of water** |
| Equipment needed: thermometer with range up to 150 oC, quickfit 250 cm3 round-bottomed flask, quickfit three-way head, quickfit bung with hole for thermometer, quickfit condenser, 100 cm3 beaker, tripod, gauze, Bunsen burner, access to gas, 2 clamp stands, each with boss and clamp, 2 pipes to connect water to tap and sink  Set up close to a sink; place a gauze on a tripod; place a Bunsen burner connected to gas underneath the tripod; pour 100 cm3 of water into the flask; place the flask on top of the gauze and clamp into position using a stand, boss and clamp; insert the three way head into the top of the flask, place the thermometer through the quickfit bung with hole, insert the bung into the top of the head; adjust the height of thermometer so it is well above the level of the water; attach condenser to head; clamp condenser into position using another stand, boss and clamp; place beaker underneath lower end of condenser to catch water; use water tube to connect tap to lower water valve on condenser; use another water tube to connect upper water valve on condenser to sink; turn tap on and ensure that condenser is full of water and that there is a steady flow of water into the sink; then light Bunsen burner and heat on medium flame until temperature has reached a steady level   * The boiling point of water is 100 oC |

***Lesson 10 - What are gases and what are their properties?***

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| Image result for test icon**Test your knowledge 10.1: Understanding the structure and properties of solid, liquids and gases** |
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| Image result for test icon**Test your knowledge 10.2: Understanding mixtures and solutions** |
| 1. Take melting point by placing a small sample in a capillary tube and strapping to a thermometer; immerse in a suitable liquid and heat until it melts; note temperature at which it starts and finishes melting; compare melting point with that of aspirin; if sample melts sharply and at same temperature it is pure; if it melts gradually over a range of lower temperatures it is not pure 2. Mg2+ ions, Cl- ions and water molecules; the Mg2+ and Cl- are distributed evenly throughout the water molecules; the solute is magnesium chloride (or Mg2+ and Cl-) and the solvent is water |

***Lesson 11 – What evidence is there to support the kinetic model of matter?***

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.1: Observe Brownian motion in pollen grains or powdered sulphur** |
| Equipment needed: microscope, slide, slide cover, smoke cell, pollen grains, dropping pipettes, access to water  Take a micro-spatula measure of pollen grains and place them on a slide; add a drop of water and then add the slide cover; remove any pollen/water which gets squeezed out when the slide cover is added.  Repeat using sulphur powder instead of pollen grains.  Light a small piece of wood or paper and collect some of the smoke using a dropping pipette; inject the smoke from the dropping pipette into the smoke cell and place a slide cover over it to keep the smoke in.  Place all three under a microscope in turn. Students should be able to see Brownian motion   * pollen grains, smoke particles and sulphur grains can be seen moving randomly and changing direction suddenly (Brownian motion) |

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg **Activity 11.2: Observe Brownian motion in the classroom** |
| If there is direct sunlight coming into the room, students should be able to observe Brownian motion in the dust particles   * dust particles moving randomly, changing direction constantly |

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 11.3: Observe bromine or nitrogen dioxide diffuse through air** |
| **Note: This demonstration should be carried out in a fume cupboard.**  Equipment needed: two gas jars with lids, access to liquid bromine or copper (II) nitrate  If using bromine: take one dropping pipette full of bromine and drop in into a gas jar inside a fume cupboard; when most of the gas jar is full of bromine, place a lid on the gas jar. Carefully place the gas jar horizontally and position another gar jar close by. Remove the lid and quickly connect the gas jars, ensuring that the bromine cannot escape.  If using NO2: Place 1 g of copper nitrate in a boiling tube. Heat the copper nitrate until it starts to decompose, giving off a brown gas. When the boiling tube is full of the brown gas, remove it from the heat. Then connect the boiling tube to another empty boiling tube upside down so that the air in the two tubes can mix.   * The orange/brown colour will gradually fill the other jar or tube, although it does take some time |

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg **Activity 11.4: Observe diffusion using smell** |
| **Note: this demonstration should be carried out either close to the end of a lesson or in another empty classroom so the room can be evacuated immediately afterwards. Take care that the smell does not overpower the students closest to the bottle. Replace the lid as quickly as possible.**   * Students closest to the ammonia bottle will smell it quickly; students further away will take longer * The ammonia molecules evaporate and diffuse through the room; they move slowly because they keep bumping into air particles |

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| Image result for test icon**Test your knowledge 11.5: Using Graham’s Law of Diffusion** |
| 1. He most quickly, Kr most slowly 2. √(40/4) = 3.2 times faster 3. √(84/20) = 2.0 times faster |

***Lesson 12 – What are the gas laws?***

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| Image result for test icon**Test your knowledge 12.1: Using the gas laws** |
| 1. 180,000 Pa; (b) 50 dm3; (c) 300,000 Pa; (d) 125,000 Pa; (e) 180 dm3; (f) 500,000 Pa |

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| Image result for test icon**Test your knowledge 12.2: Using the combined gas law** |
| 1. 229,000 Pa; (b) 100,000 Pa; (c) 45.8 dm3; (d) 125,000 Pa; (e) 396 K; (f) 11.1 dm3, (g) Gay-Lussac, Boyle, Charles, Combined |

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| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 12.3: Applying the gas laws** |
| 1. volume decreases as pressure increases;   when you reduce the volume of a balloon the pressure increases until it it more than the balloon can bear   1. volume increases as temperature increases,   at low temperatures the volume of the tyre is smaller, so it looks deflated   1. pressure increases as temperature increases,   as the air temperature in the tyre increases, so does the pressure in the tyre, on very hot days the pressure will be very high and maybe too much for the tyre to bear |

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| Image result for test icon**Test your knowledge 12.4: Using Dalton’s Law of Partial Pressures** |
| 1. H2O: 2.3/100 x 102 = 2.35 kPa; Ar: 1.0/100 x 102 = 1.02 kPa; N2 = 77.4/100 x 102 = 78.95 kPa ; O2 = 19.3/100 x 102 = 19.7 kPa |

***Lesson 13 – What are giant ionic structures?***

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| **Image result for test icon Test Your Knowledge 13.1: Understanding the physical properties of giant ionic structures** |
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| **Image result for test iconTest Your Knowledge 13.2: Deducing the unit formula of giant ionic structures** |
| 1. Li2O; (b) BaCl2; (c) Li3N; (d) NH4Cl; (e) KNO3; (f) Al(NO3)3; (g) Mg(OH)2; (h) MgSO4; (i) NH4NO3; (j) (NH4)2SO4 |

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| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 13.3: Comparing the melting points of giant ionic structures** |
| 1. NaCl because Na+ is smaller than K+ so stronger attraction to Cl- 2. MgCl2 because Mg2+ is smaller and more highly charged than Na+ so stronger attraction to Cl- 3. MgO because O2- is more highly charged than Cl- so stronger attraction to Mg2+ 4. NaF because F- is smaller than Cl- so stronger attraction to Na+ |

***Lesson 14 – What are giant metallic structures?***

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| **Image result for test iconTest Your Knowledge 14.1: Understanding the physical properties of giant metallic structures** |
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| https://image.freepik.com/free-icon/plus-sign_318-54005.jpg**Extension 14.2: Comparing the physical properties of giant metallic structures** |
| 1. Be, because the Be2+ ions are smaller and more highly charged than Li+, so attract the delocalised electrons more strongly 2. Be because B the Be2+ ions are smaller than Mg2+, so attract the delocalised electrons more strongly 3. Be will be harder/less malleable and ductile |

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| **Image result for test iconTest Your Knowledge 14.3: Exploring different simple molecular structures** |
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***Lesson 15 – What are the physical properties of simple molecular substances?***

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| **Summary Activity 15.1: Intermolecular forces** |
| * Temporary dipoles, caused by random movement of electrons in molecules, induce temporary dipoles in adjacent molecules; the resulting attraction between the molecules is a Van der Waal’s force * Hydrogen bonding is the attraction between an electropositive hydrogen atom on one molecule and an electronegative atom (O, N or F) on another molecule; it arises when H is attached to a very electronegative atom, which makes it very positive and very small * Br2 has more electrons in molecule so stronger Van der Waal’s forces between molecules * Water has hydrogen bonding between molecules but carbon dioxide does not |

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| **Image result for test iconTest Your Knowledge 15.2: Explaining the boiling points of simple molecular structures:** |
| 1. More electrons in one molecule of I2 than in one molecule of Br2, so stronger Van der Waal’s forces between molecules of I2, which are strong enough to hold I2 molecules together as a solid at room temperature 2. More electrons in one molecule of O2 than in one molecule of N2, so stronger Van der Waal’s forces between molecules of O2 3. More electrons in one molecule of HBr than in one molecule of HCl, so stronger Van der Waal’s forces between molecules of O2 4. Hydrogen bonding between HF molecules but not between HCl molecules, so overall intermolecular forces stronger between HF molecules, despite stronger Van der Waal’s forces between HCl molecules |

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| **Image result for test iconTest Your Knowledge 15.3: Explaining the physical properties of simple molecular structures:** |
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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg**Thinkabout 15.4: The special properties of water** |
| Strong hydrogen bonding causes strong surface tension in water, allowing some insects to walk on water  Strong hydrogen bonding causes ice to form an open lattice structure, causing it to be less dense than water |

***Lesson 16 – What are giant covalent structures?***

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| **Image result for test iconTest Your Knowledge 16.1: Explaining the physical properties of giant covalent structures:** |
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| **Test Your Knowledge 16.2: Understanding different types of chemical formula** |
| 1. Simplest whole number ratio of ions (in an ionic compound) or atoms (in a giant covalent compound); it is used for ionic compounds and giant covalent compounds; eg KCl, Na2CO3, SiO2 2. Number of atoms of each type in one molecule of the substance; it is used for elements and compounds with simple molecular structures; eg Cl2, CO2, O2, H2O 3. Elements with giant covalent, giant metallic or simple atomic structures (ie all elements except those with simple molecular structures) |

***Lesson 17 – Why do some substances dissolve in water but not others?***

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| https://image.freepik.com/free-icon/think-symbol-of-a-head-from-side-view-with-brain-shape-inside_318-61572.jpg**Thinkabout 17.1: Which substances dissolve in water and which do not?** |
| * sodium chloride (salt), sucrose (sugar), vinegar all dissolve in water * most oils, calcium carbonate (chalk) and most rocks do not dissolve in water * many paints use ethanol or methanol or a mixture of the two (methylated spirits) as a solvent |
| Image result for test icon**Test your knowledge 17.2: Predicting Solubility in water and in non-polar solvents** |
| BH3, CH4, H2, CO2 non-polar so more soluble in hexane  HCl highly polar so more soluble in water  NH3, HF can form hydrogen bonds with water so more soluble in water  NaCl ionic so more soluble in water |

***Lesson 18 – How can we relate the physical properties of a substance to its structure?***

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| cid:ii_jepnweck1_1621fa5f68bdf569 **Demonstration 18.1**: **Compare the physical properties of different materials** |
| Equipment needed: tongs, Bunsen burner, access to gas, three test tubes, three spatulas, three small beakers (100 cm3), two crocodile clips, two electrical wires, one small light bulb in socket connectible to wires, 12V power supply, strip of copper, 2 graphite electrodes, 1 graphite rod, a gram of sugar, a gram of sand, a gram of salt, access to water   * Hold the copper and graphite strips separately with tongs and place it in a medium Bunsen flame for one minute; nothing will happen * Pour a small amount (1 cm depth) of sugar, sand and salt separately into test tubes, hold the test tubes with tongs and place each in a medium Bunsen flame; the sugar will quickly melt but the sand and salt will not * Connect crocodile clips to each end of the copper strip and graphite rods, use a wire to connect one end of the strip to one side of the bulb, and use two more wires to connect the other end of the strip and the other side of the bulb to a 12V power supply (set the power lower than this); in both cases the bulb will light and shine brightly * Place the graphite electrodes 5 cm apart in a beaker half filled with water with the circuit otherwise complete and the power supply set at 12V; the bulb will not light; * Add a small spatula measure of sand to the water and stir gently; the sand will not dissolve and the bulb will not light * Repeat, but adding sugar; the sugar will dissolve in the water but the bulb will not light * Repeat, but adding salt; the salt will dissolve in the water and the bulb will light |

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| Image result for test icon**Test your knowledge 18.2: Summarising the relationship between physical properties and structure** |
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| Image result for test icon**Test your knowledge 18.3: Using structures to explain physical properties** |
| (a) giant ionic - NaCl; giant metallic - Mg; simple molecular - Cl2; giant covalent - SiO2; (b) strong ionic bonds require lots of energy to break; (c) ions can move when molten but not when solid; (d) Na+ and Cl- ions attracted to polar water molecules; (e) delocalised electrons can move; (f) layers of cations can slip over each other without breaking the metallic bonds; (g) strong metallic bonds require lots of energy to break; (h) intermolecular forces between Cl2 molecules don’t require much energy to break; (j) strong covalent bonds require lots of energy to break; (k) graphite has delocalised electrons, diamond does not; (l) layers of atoms in diamond cannot slip over each other without breaking covalent bonds; layers of atoms in graphite can slip over each other - only intermolecular forces need to be overcome; (m) it is very hard; (n) it is very soft and the layers can slip off |

***Lesson 19 – How does structure and bonding in the elements change across the Periodic Table?***

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| **Summary Activity 19.1: Periodic properties of atoms** |
| * Atomic size decreases across a Period; nuclear charge increases, shielding stays the same, no nuclear attraction increases and outer shell electrons are pulled in more strongly * Atomic size increases down a Group; number of shells increases so shielding increases; nuclear attraction decreases (despite increase in nuclear charge) and outer shell electrons are pulled in less strongly * Electronegativity increases across a Period; nuclear attraction increases so attraction to bonding electrons increases * Electronegativity decreases down a Grouo; nuclear attraction decreases so attraction to bonding electrons decreases |

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| Image result for test icon**Test your knowledge 19.2: Explaining the bonding and structure of the Period 2 elements** |
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| Image result for test icon**Test your knowledge 19.3: Explaining the bonding and structure of the Period 1 and 3 elements** |
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***Lesson 20 – How does structure and bonding in compounds change across the Periodic Table?***

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| Image result for test icon**Test your knowledge 20.1: Describe the structure and bonding in oxides and chlorides** |
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| Image result for test icon**Test your knowledge 20.2: Summarising structure and bonding in the Periodic Table** |
| 1. Electronegativity increases so less tendency to form metallic bonds 2. Metals - lithium, beryllium, sodium, magnesium, aluminium; Non-metals: any of the other 13! 3. LiCl, NaCl and MgCl2 4. Any oxide in the simple molecular column of LA 2.16 5. Sodium chloride is giant ionic but aluminium chloride is simple molecular 6. Aluminium oxide is giant ionic but silicon dioxide is giant covalent |

***Lesson 21 – What have I learned about Particles, Bonding and Structure?***

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| 21.1 END-OF-TOPIC QUIZ  UNIT 2 – PARTICLES, BONDING AND STRUCTURES  Image result for test icon |
| Difference in electronegativity between Mg and Cl is large; difference in electronegativity between N and H is much smaller  N on NH3 has a lone pair; it can form a dative covalent bond with H+ by providing both of the shared electrons; the resulting species is NH4+    , tetrahedral  , non-linear (104.5o)  Random movement of electrons causes temporary dipole, which induces a dipole in the adjacent molecule; the resulting dipoles attract  SO2 is non-linear so the dipoles on the two bonds do not cancel out; CO2 is linear so the dipoles on the two bonds do cancel out  Particles vibrate about fixed positions in a solid; on heating, the particles get more energy so vibrate more and faster; eventually the particles have enough energy to break some of the bonds between the particles and so the particles are able to move around  Brownian motion is the random movement in different directions of visible particles such as smoke, pollen or sulphur; it shows that the particles are constantly being hit by smaller particles and pushed in different directions  P1V1/T1 = P2V2/T2 so V2 = P1V1T2/P2T1 = 29.2 dm3  3D lattice of Mg2+ cations, held together by a sea of delocalised electrons; the delocalised electrons are free to move  3D lattice of C atoms, each attached to four others in a tetrahedral arrangement; all atoms are held in place by strong bonds and the atoms cannot move without breaking the bonds, so the structure is hard to break and cannot be bent  3D lattice of Na+ and Cl- ions; each Na+ surrounded by 6 Cl- and vice versa; water is a polar molecule; the Na+ ions are attracted to the negative pole and the Cl- ions are attracted to the positive pole  Water can form hydrogen bonds, H2S can only form Van der Waal’s forces, hydrogen bonds are generally stronger |