CHEMISTRY PAPER 1: Atomic structure and the periodic table

Facts: Atoms, elements, compounds and mixtures

1.Atom	The **smallest particle** that makes up all substances.		
2. Atom radius	The radius **(r)** of an atom is **about 0.1nm ($1x10^{-10}$m)**.		
3. Subatomic particles	**Particles** that are **smaller than an atom**. Atoms are made of three subatomic particles **1. Protons 2.Neutrons 3.Electrons**		
4. Nucleus	**Central part** of an atom. It has a **radius less than 1/10,000** of that of the atom **($1x10^{-10}$m)**.		
5. Proton	Relative mass of 1.	Relative charge of +1.	Found in the nucleus.
6. Neutron	Relative mass of 1.	Relative charge of 0.	Found in the nucleus.
7. Electron	Relative mass of almost 0.	Relative charge of -1.	Found in the shells.
8. Electronic configuration	The **arrangement** of **electrons** in **shells** (energy levels). 2.8.1 = 2 electrons in 1st shell, 8 electrons in 2nd shell, 1 in 3rd shell.		
9. Filling shells	1st shell holds 2 electrons. 2nd and 3rd shell hold 8 electron max. Inner shells filled first.		
10. Mass number	The number of **protons and neutrons** in the **nucleus**.		
11. Atomic number	The number of **protons** in the **nucleus.** Atoms of **different elements** have **different numbers of protons** and therefore **atomic numbers**. E.g. Carbon = 6 and Oxygen = 8.		
12. Overall charge	Atoms are **neutral (0 charge)** because the **number of electrons = the number of protons** so the **positive charge** of the protons **cancels out** the **negative charge** of the electrons.		
13. Element	A substance made up of only **one type** of **atom**. Represented by a single capital letter.		
14. Isotopes	**Forms of an element** that have the **same number** of **protons (atomic number)** but **different numbers** of **neutrons (mass number)**.		
15. Compounds	**Two or more** elements **chemically bonded together** e.g. calcium carbonate, $CaCO_3$.		
16. Mixtures	**Two or more elements or compounds NOT chemically bonded.** They can be **separated** by the physical processes of filtration, evaporation, distillation and chromatography.		

Facts: Development of the model of the atom

17. John Dalton (1803)	**Solid sphere model:** Atoms were **tiny spheres** that could **not be divided**.
18. JJ Thompson (1904)	**Plum pudding model:** Discovery of the **electron** led to the idea that the atom was a ball of + charge with negative electrons throughout**.**
19. Rutherford (1911)	**Nuclear model:** The **alpha particle scattering experiment** showed that the **mass** of the atom was **concentrated** in the **centre** and that the **nucleus** which was **positive.**
20. Niels Bohr (1913)	**Nuclear model:** Calculations suggested **electrons** orbit at **set distances.**
21. Chadwick (1932)	**Nuclear model**: **Discovery** of **neutrons** in the nucleus.

Facts: Modern Periodic Table

22. Arrangement	Elements are arranged in order of **atomic number.**
23. Groups	**Vertical columns.** Elements in the s**ame group** have the **same number** of **electrons** in their **outer shell**. So elements in the **same group** have **similar chemical properties**.
24. Periods	**Horizontal rows**. Elements in the **same period** have the **same number of electron shells** e.g. period 3 = 3 electron shells.
25. Metals	On the left and in the centre of the periodic table. They lose electrons and form + ions. High melting and boiling points and good conductors of electricity/heat.
26. Non-metals	On the right of the periodic table. They gain electrons and form - ions. Low melting and boiling points and poor conductors of electricity/heat.

Facts: Group 0 – Noble gases

27. Structure	**Single atoms** with **full outer shell** of electrons. Elements include: Helium (He), Argon (Ar).
28. Unreactive	Also called **inert** due to **full electron shells**.
29. Boiling Points	I**ncrease down the group because t**he atoms get **larger in size.**

Facts: Group 1 – Alkali metals

30. Structure	One electron in the outer shell. Elements include: Lithium (Li), Sodium (Na), and Potassium (K).
31. Reactions	React with metals to form ionic compounds. They gain one electron to become a -1 ion.
32. Properties	Low density (float). React violently with water to form metal hydroxide (alkaline solutions).
33. Reactivity	**Increases down the group** because more electron shells → more shielding → less attraction from nucleus → easier to lose one electron.

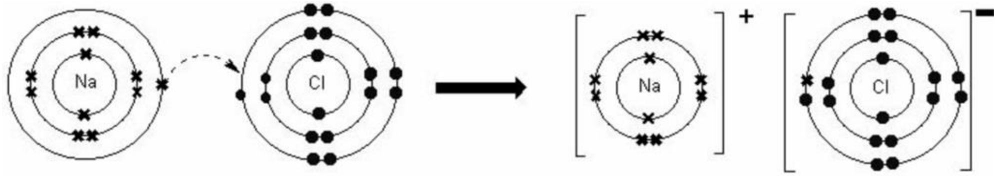
Facts: Group 7 – Halogens

34. Structure	They have seven electron in their outer shell so form diatomic molecules e.g.F_2, Cl_2, Br_2.
35. Reactions	React with non-metals to form ionic compounds. They gain one electron to become a -1 ion.
36. Properties	Boiling and melting points increase down the group.
37. Reactivity	**Decreases down the group** because more electron shells → more shielding → less attraction from nucleus → harder to gain one more electron.

CHEMISTRY PAPER 1: Bonding structure and properties of matter

Facts: Ionic bonding

1. Ionic bonding	Compounds formed when **metals react with non-metals** by **transferring electrons**.
2. Atoms	Atoms are **neutral (0 charge)** because the **number of electrons = the number of protons** so the **positive charge** of the protons **cancels out** the **negative charge** of the electrons.
3. Ions + or -	**Charged particles** formed when atoms **lose or gain electrons** to get a **full outer shell**.
4. Metals	**Lose electrons** and become **positive ions**.
5. Group 1	Metals with **1 electron in their outer shell**. So **lose 1 electron** to become +**1 ions**.
6. Group 2	Metals with **2 electrons in their outer shell**. So **lose 2 electrons** to become **+2 ions**.
7. Non-metals	**Gain electrons** and become **negative ions**.
8. Group 6	Non-metals that **gain 2 electrons** to become **-2 ions**.
9. Group 7	Non-metals that **gain 1 electron** to become **-1 ions**.
10. Ionic bond	The **force of electrostatic attraction between oppositely charged ions** e.g. $[Na]^+$ and $[Cl]^-$



Facts: Ionic compounds

11. Ionic compound	A **giant regular structure** (lattice) of ions held together **by strong electrostatic forces** of attraction **between oppositely charged ions**.
12. High melting + boiling points	A **large** amount of **heat energy (high temp)** is needed to **overcome (break)** the **strong electrostatic forces** of attraction between oppositely charged **ions**.
13. Conducting electricity	An ionic compound **cannot conduct electricity** when **solid** because the **ions** are **not mobile** (free to move). The compound must be **melted** or **dissolved** in water.

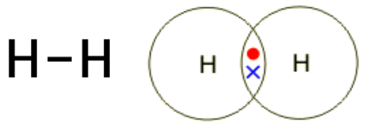
Facts: Metallic bonding

14. Metals	**Giant structures** of atoms arranged in a lattice.
15. Delocalised electrons	Electrons that have left their outer-shell and are **free to move** around.
16. + metal ion	The ion formed when a **metal atom loses an electron.**
17. Metallic bond	The **electrostatic force between** the **delocalised electrons** and the **+ metal ion.**
18. High melting + boiling points	A **large** amount of **heat energy (high temperature)** is needed to **overcome (break)** the **strong metallic bond between** the **delocalised electrons** and the **+ metal ion**.
19. Good conductors	Metals **conduct heat** and **electricity** because they have **delocalised electrons** that can **move** through the **whole structure.**
20. Malleable	Metals can be **bent** and **shaped** without breaking because their **atoms** are arranged in **layers** that can **slide** over each other.
21. Alloys	Metals are **too soft** for many uses so they are **mixed** with other **metals** to make **alloys.**
22. Alloy properties	Alloys are hard because the **different sized metal atoms distort/change** the **layers** and they are **not able to slide** over each other.

Facts: Covalent bonding

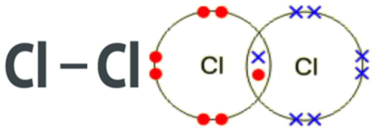
23. Covalent bonding	Bonds formed when **non-metal atoms react together** and **share a pair electrons**.

24. Hydrogen H_2



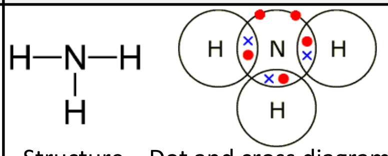
Structure Dot and cross diagram

25. Chlorine Cl_2



Structure Dot and cross diagram

26. Ammonia NH_3



Structure Dot and cross diagram

Facts: Simple covalent structures (small molecules formed from a few atoms)

27. Intermolecular forces	The **forces between molecules**. Shown by a dashed line **-------**.
28. Low melting + boiling point	A **small** amount of **heat energy (low temp)** is needed to **overcome (break)** the **weak intermolecular forces** so they are **liquids or gases** at room temperature.
29. Conducting electricity	**Cannot conduct electricity** because they **do not have delocalised electrons** that can **move** or no **overall electric charge**.

Facts: Giant covalent structures (large molecules formed from many atoms)

30. High melting and boiling points	A **large** amount of **heat energy (high temp)** is needed to **overcome (break)** the many **strong covalent bonds.**
31. Diamond	Made from **carbon**. **Each carbon atom is bonded to 4 other carbon atoms** by **strong** covalent bonds. Does not conduct electricity as there are **no delocalised electrons** to **carry charge** through the **whole structure.**
32. Graphite	Made from **carbon**. **Each carbon is bonded to 3 other carbon atoms**. The carbon atoms form a **hexagonal** layered structure. There are **delocalised electrons between the layers** so graphite can conduct electricity.

CHEMISTRY PAPER 1: Quantitative chemistry

Facts: Compounds	
1. Compounds	**Two or more** elements **chemically bonded together.**
2. Compound formulas	Show the **number of elements** and the **number of atoms** in the compound. **E.g.$CaCO_3$** has **3 elements**, calcium, carbon and oxygen and 5 atoms, 1 calcium, 1 carbon, 3 oxygen
3. Equations	**Word:** magnesium + oxygen → magnesium oxide **Symbol:** $2Mg + O_2 \rightarrow 2MgO$
4. Reactants	Substances that **react together**. Go **before** the arrow
5. Products	Substances **made** in the reaction. Go **after** the arrow.
6. Arrow	Shows a reaction is taking place and in which **direction**.
7. Conservation of mass	**No atoms are lost or made** during a chemical reaction. **Mass of products = mass of reactants**.
8. Balanced equations	The **law of conservation of mass** is why we have to **balance symbol equations** (the number of atoms **must be the same on both sides** of the equation).
9. Multiplier numbers	-The **normal size** number **before** a formula. -Show **how many molecules** of a substance there are: $\underline{2}H_2 + O_2 \rightarrow \underline{2}H_2O$. -The reaction has: 2 hydrogen molecules, 1 oxygen molecule and 1 water molecule.
10. Subscript numbers	-The **small**, **low** number **within** a formula. -Shows the **number of atoms** of an element **in a molecule** $H_{\underline{2}} + O_{\underline{2}} \rightarrow H_{\underline{2}}O$. -There are 2 hydrogen atoms in **each hydrogen molecule**, 2 oxygen atoms in the **one oxygen molecule** and 2 hydrogen and 1 oxygen atom in **each water molecule**.

Facts: Relative formula mass		
11. Relative atomic mass **(A_r)**	**Relative mass of an atom.** The top number of each atom.	$\underline{12}$ C 6
12. Relative formula mass **(M_r)**	**The relative atomic masses** of **all** the **atoms** in a formula added together.	
13. Example: Calculate the relative formula mass **(M_r)** of water, **H_2O.**	**H_2O**: There are 2 atoms of hydrogen so 1 + 1 = 2 plus 1 atoms of oxygen so 16 The total = 2 + 16 = 18	$\underline{1}$ H 1 $\underline{16}$ O 8

Facts: Relative formula mass	
14. Percentage mass	The mass of one element in a compound represented as a percentage
15. Equation	% mass of element in a compound = $\frac{\text{(number of atoms of element) x } (A_r \text{ of element})}{M_r \text{ of compound}}$

13. Example: Calculate the percentage of hydrogen in of water, **H_2O.**	The M_r of H_2O is 18. So, $\frac{2 \times 1}{18}$ = 0.11 0.11 x 100 = 11. 11%

Facts: Mass changes when there is a gas in the reaction	
14.Closed systems	A test tube with a bung on. Nothing can enter or leave, **mass of reactants=mass of products**
15. Non enclosed systems	A test tube **without** a **bung**. **Substances** (like **gases)** can **enter or leave**. Mass **appears** to **increase** or **decrease**.
16. Mass increase	The mass of magnesium when burnt appears to increase because **oxygen** gas in **added**. Magnesium **+ oxygen$_{(g)}$** → magnesium oxide
17.Mass decrease	The mass of the reactants decrease when calcium carbonate is added to acid because **carbon dioxide** is **released**. $CaCO_3 + 2HCl \rightarrow CaCl_2 + \mathbf{CO_{2(g)}} + H_2O$

Facts: Chemical measurements: Uncertainty		
18. Uncertainty	**In science there is always uncertainty** when measurements are taken.	
19. How to calculate uncertainty	1. Calculate the mean/average. 2. Calculate the range (highest-lowest). 3. Divide the range in half. 4. Uncertainty is given as the value ± half of the range.	Data set = 50, 51, 49, 50 1. Mean = $\frac{50 + 51 + 49 + 50}{4}$ = **50** 2. Range = 51 - 49 = 2 3. 2 divided by 2 = 1 4. Uncertainty = **50** ± 1

Facts: Concentration	
1 – Solution	A **mixture** formed by a **solute** and a **solvent**.
2 – Solute	The **dissolved substance** (usually a **solid**) in a solution.
3 - Solvent	The **liquid** (usually **water**) in which the solute dissolves to form a solution.
4 - Concentration	The **mass of solute dissolved in a given volume** of solvent.
5 – Units	Grams per dm^3 (g/dm^3) or moles per dm^3 (mol/dm^3).
6 - dm^3	Cubic decimetres.
7 - Conversions	1 dm^3 = 1000 cm^3. To convert cm^3 to dm^3 ÷ by 1000. To convert dm^3 to cm^3 x by 1000.
9 – Increasing concentration	1. Increase the mass of the dissolved solute (by dissolving more in a given volume of solution). 2. Reducing the volume of solvent (adding less water).
10 – Equation	Concentration (g/dm^3 or mol/dm^3) = $\frac{\text{mass (g or mol)}}{\text{volume } (dm^3)}$

CHEMISTRY PAPER 1: Chemical changes

Facts: Metal extraction

1. Metal ore	A rock that contains enough of a metal or a metal compound (usually metals oxides) to make extracting the metal worthwhile.
2. Reactivity series	A list of metals in **order** of how **reactive** they are based on their reactions with water and acid. The series is Used to determine **extraction** method.
3. Non-metals	**Carbon** and **hydrogen** are often included in the reactivity series.

4. Extraction method

potassium (most reactive)	**K**	Extracted using **electrolysis**
sodium	**Na**	
calcium	**Ca**	
magnesium	**Mg**	
aluminium	**Al**	
carbon	**C**	
zinc	**Zn**	Extracted by heating with **carbon**
iron	**Fe**	
tin	**Sn**	
lead	**Pb**	
hydrogen	**H**	
copper	**Cu**	Found in their **natural state**, not bonded in an ore.
silver	**Ag**	
gold	**Au**	
platinum (least reactive)	**Pt**	

Facts: Heating with Carbon

5. Oxidation	The **addition of oxygen** to a substance. E.g. $2Ca + O_2 \rightarrow 2CaO$. Ca has been **oxidised.**
6. Reduction	The **removal of oxygen** from a substance. E.g. $CuO + H_2 \rightarrow Cu + H_2O$. Cu has been **reduced**
7. Extracting iron	iron oxide + carbon → iron + carbon dioxide The iron is **reduced** and the carbon **oxidised**.

Facts: Electrolysis

8. Ionic compound	A **giant lattice** structure of **positive** and **negative ions** held together by strong **electrostatic forces**.
9. Electrolysis	A process that breaks down an **ionic compounds** using an **electrical current.**
10. Conducting electricity	An ionic compound **cannot conduct electricity** when **solid** because the **ions** are **not mobile** (free to move). The compound must be **melted** or **dissolved** in water.
11. Electrolyte	The **liquid** or **solution** formed from melting or dissolving the ionic compound that can **conduct electricity** because the **ions are mobile** (free to move).
12. Electrode	**Conducts electrical current** to and from a power source. Often made from carbon.
13. Anode +	**Positive** electrode = **O**xidation **I**s **L**oss of electrons.
14. Cathode -	**Negative** electrode = **R**eduction **I**s **G**ain of electrons.

Facts: Electrolysis of aluminium oxide

15. Anode	O^{2-} is **attracted** to the **anode** and collected as oxygen gas. The oxygen **reacts with** the **carbon electrode** , forming **carbon dioxide.** This wears the electrodes away = **replacing.**
16. Cathode	**Al^{3+} is attracted to the cathode and** collected as aluminium metal.

Facts: Electrolysis of aqueous solution

17. Aqueous solution (aq)	A solution that uses **water** to **dissolve** a substance e.g. an ionic compound**.**
18. Cathode (-ve)	1. **Hydrogen** is formed **if the metal is more reactive** than **hydrogen**. 2. **Metal ions** are formed **if the metal is less reactive** than **hydrogen**.
19. Anode (+ve)	1. A halogen is formed if a **halide ion** e.g. F-,Cl-, Br-, I- is present. 2.Oxygen is formed if there is **no halide ion** e.g. F-,Cl-, Br-, I- is present.

Facts: Acids and alkalis

20. Dissociate	The **breaking up** of a molecule into **ions** when **dissolved in water**.
21. Acids	**Dissolve in water** to form an **aqueous solution** containing **hydrogen ions H^+.** pH **1-6**.
22. Acid examples	Hydrochloric acid, **HCl**, nitric acid, **HNO_3** and sulphuric acid, **H_2SO_4**.
23. Alkalis	**Dissolve in water** to form an **aqueous solution** containing **hydroxide ions OH^-**. pH **8-14.**
24. Alkali examples	Most **metal hydroxides** e.g. copper hydroxide and **ammonia**.
25. Neutral	A substance that has a **pH of 7** e.g. water.

Facts: Metal and acid reactions

26. Equation 1	Acid + alkali → salt + water
27. Equation 2	Acid + metal oxide/metal hydroxide → salt + water
28. Equation 3	Acid + metal carbonate → salt + water + carbon dioxide
29. Hydrochloric	Produces a salt ending in **chloride**. Sodium chloride.
30. Nitric	Produces a salt ending in **nitrate**. Copper nitrate.
31. Sulphuric	Produces a salt ending in **sulphate**. Lead sulphate.

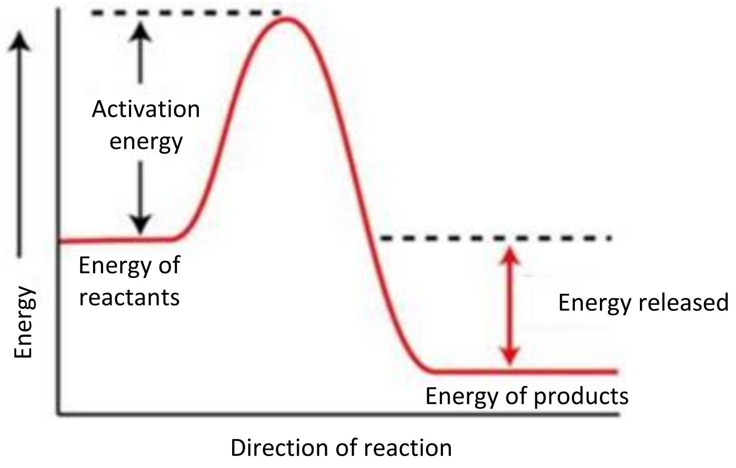
CHEMISTRY PAPER 1: Energy changes

Facts: Chemical reactions	
1. Chemical reaction	Occurs when reacting particles **collide** with each other with **sufficient energy,** making a **successful** collision.
2. Conservation of energy	The amount of energy in the universe at the end of a chemical reaction is the same as before the reaction takes place.
3. Activation energy	The **minimum** amount of **energy** that particles must have to react (successful collision).
4. Catalyst	**Speeds up** a chemical reaction without being used up. It **lowers** the **activation energy** needed for a reaction to take place.

Facts: Exothermic reactions		Facts: Endothermic reactions	
5. Exothermic	A reaction that transfers energy to the surroundings.	9. Endothermic	A reactions that **take in energy from** the **surroundings**.
6. Reactions	Combustion, many oxidation and most neutralisation reactions.	10. Reactions	Thermal decomposition, electrolysis, photosynthesis.
7. Uses	Hand-warmers and self-heating cans.	11. Uses	Instant icepacks to treat sport injuries.
8. Observations	**Temperature increases**, a glow.	12. Observations	**Temperature decrease**.

Facts: Energy level diagrams (reaction profiles)	
13. Energy level diagram	Diagrams that shows if a reaction is **exothermic** or **endothermic**. It shows the **energy** in **reactants** and **products**.
14. Exothermic	The **reactants** are **above** (have more energy) the **products** as **energy is given out.**
15. Endothermic	The **reactants** are **below** (have less energy) the products, as **energy is taken in.**

16. Energy level diagram: Exothermic reaction



17. Energy level diagram: Endothermic reaction