Year 9 – Autumn 2 – Investigative Chemistry

Type of bonding	Occurs between...	Movement of electrons	Bond	Structure	Example	Properties	Reason for property
Ionic 	Metals and non-metals	Electrons transferred from outer shell of metal to outer shell of non-metal. Both metals ion and non-metal ion have a full outer shell	Strong electrostatic force between positive metal ion and negative non-metal ion	Giant ionic lattice	Sodium chloride Lithium bromide Calcium oxide	Solid at room temperature	They have high melting points
						High melting and boiling points	Ions are held together by strong electrostatic forces - a lot of energy is needed to overcome these forces
						Do not conduct electricity when solid	Ions (charged particles) cannot move
						Do conduct electricity when molten or in solution	Ions (charged particles) are free to move
Covalent 	Non-metals	Atoms share pairs of electrons	A shared pair (or multiple pairs) of electrons between atoms	Simple covalent molecules	Chlorine Cl_2 Oxygen O_2 Water H_2O Methane CH_4 Ammonia NH_3 Hydrochloric acid HCl	Liquids or gases at room temperature	Low melting and boiling points
						Low melting and boiling points NB: the larger the molecules the higher the melting/boiling point (intermolecular force increases with size of molecule)	Weak intermolecular forces between the molecules
						Do not conduct electricity	No free/delocalised electrons nor are the molecules charged
				Giant covalent lattice		Solid at room temperature High melting and boiling point	All atoms are linked to others by strong covalent bonds
					Diamond	Hard	Each carbon atom formed four strong covalent bonds with other carbon atoms
						Does not conduct electricity	No free (no delocalised) electrons
						High thermal conductivity	There are strong covalent bonds between atoms. When one carbon atoms vibrates it causes all four neighbouring atoms to vibrate
					Graphite	Soft and slippery	Each carbon atom is bonded with three carbon atoms resulting in a layered structure. There are weak intermolecular forces between the layers meaning they can easily slide over each other
						Conducts electricity	One electron from each carbon atom is delocalised
					Silicon dioxide	Hard	Strong covalent bonds hold the oxygen and silicon atoms together
						Does not conduct electricity	No free (no delocalised) electrons
Metallic	Metals (elements) Alloys (mixtures of metals)	Metals in outer shell are delocalised and so are free to move throughout the whole structure	Electrostatic attraction between the delocalised electrons and positive metal ions	Regular arrangement (lattice) of positive ions held together by strong electrostatic	All metals All alloys	High melting and boiling points	Strong metallic bonds between positive ionic lattice and delocalised electrons
						Good electrical	Delocalised electrons are charge carriers
						Good thermal conductors	Delocalised electrons can transfer thermal energy

2. Graphene and Fullerenes

Material	Details	Uses
Graphene	Single layer of graphite. 	Electronics, composites
Fullerenes 	Molecules of carbon with hollow shapes. e.g. Buckminster fullerene, C_{60} - spherical e.g. Carbon nanotubes - cylindrical fullerenes with very high length to diameter ratios	Caron nanotubes: nanotechnology, electronics, materials

3. Reactions of Metals

React with	Word equation	Example
Oxygen	Metal + oxygen → metal oxide	Magnesium+oxygen → magnesium oxide Mg (s) +O_2 (g) → MgO (s)
Water	Metal + water → metal hydroxide + hydrogen	Potassium + water → potassium hydroxide + hydrogen K (s)+H_2O (l) → KOH (aq) + H_2 (g)
Acid	Metal + acid → metal salt + hydrogen NB: salt is an ionic compound (metal bonded to non-metal)	Magnesium + sulphuric acid → magnesium sulphate + hydrogen Mg (s) +H_2SO_4 (aq) → $MgSO_4$ (aq) +H_2 (g)

Oxidation: the addition of oxygen to an element.
Reduction: the removal of oxygen from a compound.

4. Conservation of Mass and Balancing Equations

Law of conserva - tion of mass	No atoms are lost or made during a chemical reaction. Mass of the products = mass of the reactants. If a gas is produced and escaped, it may **seem** like the mass has decreased.	E.g. Hydrogen and chlorine react together to form hydrogen chloride. If 5g of hydrogen and 5g of chlorine react together, there will be 10g (5g+5g) of hydrogen chloride produced
Balancing equations	In a symbol equation the numbers of atoms of each elements on each side of the equation must be equal. Rules: 1. Count number of atoms of each element on left and right hand side of equation 2. If are NOT the same on each side, need to balance the equation 3. Only add BIG numbers in FRONT of each compound/element, never small number afterwards	Example: 1. H_2 + O_2 → H_2O NOT balanced $2H_2$+O_2→$2H_2O$ IS balanced 2. Na +O_2 → Na_2O NOT balanced 4Na + $2O_2$ → $4Na_2O$ Is balanced

5. Acids and Bases

Statement	More detail
When a compound dissolves in water, it dissociates (splits up) into its individual ions	When sodium chloride NaCl dissolves, it splits up in Na^+ and Cl^-
Hydroxide ions OH^- make solutions alkaline	Sodium hydroxide NaOH is alkaline because it contains hydroxide OH^- ions
Hydrogen ions H^+ make solutions acidic	Hydrochloric acid HCl is acidic because it contains hydrogen ions H^+
pH scale tells us how acidic or alkaline something is	The pH scales runs from 0-14. pH of 0 = the most acidic pH of 7 = neutral (neither acidic nor alkaline) pH 14 - the most alkaline.
Measure the pH using: 1.)An indicator changes colour depending on whether it is in an acidic or alkaline solution 2) A pH probe - a digital metre than display the pH on a screen	1) Types of indicator Universal indictor = continuous scales from red (most acidic) to green (neutral) to purple (more alkaline) Phenolphthalein = colourless in acid, pink in alkali Methyl orange = red in acid, orange in alkali

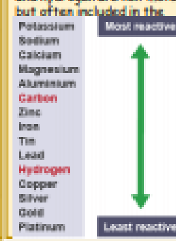
6. Neutralisation

Statement	More detail
Acids are neutralised by: 1. alkalis (soluble metal hydroxides) and by bases (insoluble metal hydroxides) 2. Metals carbonates	1. Acid + alkali/base → salt + water e.g. Hydrochloric acid + metal hydroxide → metal chloride + water 2. Acid + metal carbonate → salt + water + carbon dioxide e.g. hydrochloric acid + metal carbonate → metal chloride + water + carbon dioxide
The salt produced in the neutralisation reaction depends on: - Type of acid - Positive ion in the alkali, base or carbonate	e.g.: - hydrochloric acid produces chlorides nitric acid produces nitrates, sulfuric acid produces sulfates

NB. Bases and alkalis are both defined by their hydroxide ions. The difference is: alkalis are soluble while bases are insoluble

7. Reactivity Series

-Metals react to form positive ions
-A metal is more reactive, the easier it can lose electrons
-**Reactivity series** is the metals placed in order of their reactivity (NB: carbon and hydrogen are non-metals but often included in the



8. Displacement Reactions

A more reactive metal with displace (swap with) a less reactive metal in a compound.
The more reactive metals wants to be in a compound!
e.g.
1. Lithium is more reactive than aluminium so I will displace it:
Lithium + aluminium bromide → aluminium + lithium bromide
2. Potassium is less reactive than lithium so it will not displace it:

9. Extraction Methods

Unreactive metals (e.g. gold) occur naturally in their pure form.
Majority of metals occur naturally in a compound.
Extraction methods are ways to "extract" (obtain) the metal from the compound.
Metals that are LESS reactive than carbon are extracted from their oxides by reacting with carbon (see 7. Reactivity Series and 8. Displacement Reactions)
E.g. copper is extracted from copper oxide by using carbon because copper is lower down in the reactivity series.

10. Making Soluble Salts

Soluble salts can be made by reacting:
- **Acid with**
- **Solid insoluble substances (e.g. metals, metal oxides, hydroxides or carbonates)**

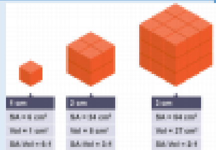
Procedure:
1. **Add solid to acid** until no more of the solid reacts with the acid
2. **Filter** the excess solid (unreacted solid)
3. The filtrate is a solution of the soluble salt
4. The salt produced can be retrieved by **crystallisation**

Year 9 – Autumn 2 – Investigative chemistry (Triple content)

11. Nanoscience

a. Size of particles and their properties

- Nanoscience studies three types of small particle: **coarse, fine and nano**
- Small particles have a **LARGE surface area to volume ratio** (i.e. their surface is very big compared to their volume)
- Changing the particle size dramatically effect this ratio.
  E.g. If reduce length of side of cube by a factor of 10 then
  - **Surface area decreases by 10x10=100**
  - **Volume decreases by 10x10x10=1000**
  - **Surface area to volume ratio increases by 10**



Particle	Size of diameter
Coarse	1×10^{-5} - 2.5×10^{-6} m
Fine	100nm-2500nm (1×10^{-5} - 1×10^{-7}m)
Nano	1nm-100n (1×10^{-9} - 1×10^{-7})m
An atom	0.1 nm 1×10^{-10}

b. Nanoparticles

- **A few hundred atoms** in size
- Large surface area to volume ratio means that:
  - nanoparticles have different properties to same material in bulk
  - smaller quantities are needed than in bulk
- Applications of nanoparticles:
  - Sun cream
  - Cosmetics
  - Electronics
  - Medicine
  - As catalysts

12. Percentage Yield

No atoms are gained or lost in a chemical reaction (see 5. *Conservation of Mass*). Yet, cannot always obtained the amount of product calculated because:

- Reaction may be **reversible**
- Some product may be lost when separated from reaction mixture
- Reactant may react in a different way to expected reaction

Yield: the actual amount of product obtained
Percentage yield: amount of product obtained compared to maximum theoretical amount.
Equation to know:

Percentage yield (%) = $\dfrac{\text{mass of product actually made}}{\text{maximum theoretical amount of product}}$

13. Atom Economy

Atom Economy	Size of diameter
Definition	Measure of amount of starting material that end up as **useful** product. (IAso called: atom utilisation)
Why is it important?	For sustainable development and for economic reasons Reason: more products made means less waste produced
Equation	**Atom economy** = $\dfrac{\text{relative formula mass of desired product}}{\text{sum of relative formula mass of all reactants}} \times 100$

14. Weak and Strong Acids

Strong acid	Completely **ionised** (split up into ions) in **aqueous solution** (dissolved in water)
Examples of strong acids	Hydrochloric acid, $HCl_{(aq)}$ Sulfuric acid, $H_2SO_{4(aq)}$ Nitric acid, $HNO_{3(aq)}$
Example: HCl	$HCl_{(g)} + aq \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$
Weak acid	Partially ionised in aqueous solution
Examples of weak acids	Ethanoic acid Citric acid Carbonic acid
Example: Ethanoic acid	$CH_3COOH_{(l)} + aq \rightleftharpoons CH3COO^-_{(aq)} + H^+_{(aq)}$ The double arrow here shows that there is only **partial dissociation** - some of the molecules stay as CH_3COOH instead of its ionic form CH_3COO^-.

For a fixed concentration: the **stronger the acid, the lower the pH**.
As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.

15. Tests for Ions

Ion	Test	Observation if ion present
Lithium, Li^+	Flame test	Crimson flame
Sodium, Na^+		Yellow flame
Potassium, K^+		Lilac flame
Calcium, Ca^{2+}		Orange-red flame
Copper, Cu^{2+}		Green flame
Aluminium, Al^{3+}	Sodium hydroxide solution added dropwise	Form white precipitates Precipitate dissolves in excess sodium hydroxide solution
Calcium, Ca^+		Forms white precipitate Does not dissolve in excess sodium hydroxide solution
Magnesium, Mg^{2+}		Forms white precipitate Does not dissolve in excess sodium hydroxide solution
Copper (II), Cu^{2+}		Forms blue precipitate
Iron (II), Fe^{2+}		Forms green precipitate
Iron (III), Fe^{3+}		Forms brown precipitate
Carbonate, $CO3^{2-}$	Add dilute acid	Carbon dioxide produced (carbon dioxide turns limewater cloudy)
Chloride, Cl^-	Dilute nitric acid, add silver nitrate dropwise	White precipitate (compound formed: silver chloride)
Bromide, Br^-		Cream precipitate (compound formed: silver bromide)
Iodide, I-		Yellow precipitate (compound formed: silver iodide)
Sulfate, SO_4^{2-}	Dilute hydrochloric acid, add barium chloride dropwise	Forms white precipitate

16. Redox Reactions

Oxidation	Loss of electrons
Reduction	Gain of electrons
Redox reaction	Chemical reaction where both reduction and oxidation occurs Example: $Fe + Cu^{2+} \rightarrow Cu + Fe^{2+}$ • Fe has lost two electrons to form Fe^{2+}→Fe has been oxidised • Cu^{2+} has gained two electrons to form Cu→Cu^{2+} has been reduced

O xidation **I** s **L** oss

R eduction **I** s **G** ain

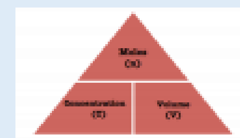
17. Ionic and Half Equations

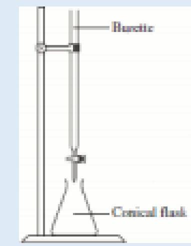
Ionic equation	Chemical equation showing only the ions that are involved in displacement reaction
Half equation	An ionic equation focusing on ONE species including the electrons that are transferred (An ionic equation can be split into two half equations → one for species that gain electrons and one for species that lose)
Spectator ion	Ions that do not change their electronic state (ionic charge) during the reaction. Spectator ions are NOT included in the ionic equation
Example	**Word equation:** Potassium + lithium chloride → potassium chloride + lithium (spectator ion: Cl^-, species involved in displacement: potassium and lithium) **Symbol equation:** K + LiCl → KCl + Li **Ionic equation:** $K + Li^+ \rightarrow K^+ + Li$ **Half equation for potassium** - potassium loses one electron to form potassium ion: $K - e^- \rightarrow K^+$ (OR $K \rightarrow K^+ + e^-$) **Half equation for lithium** - lithium ion gains one electron to form lithium: $Li^+ + e^- \rightarrow Li$

Metals can be extracted from molten compounds using electrolysis. This technique is used if the metal is too reactive to be extracted by reduction with carbon or if the metal reacts with carbon.
This extraction technique is expensive because large amounts of energy are needed to melt the ionic compounds and to produce

18. Titration

Titration: Experimental technique to find out how much acid is required to neutralise an alkali.
When neutralisation takes place, the hydrogen ions from the acid bond with the hydroxide ions from the alkali to produce water:
H^+ (aq) + OH^- (aq) → H_2O (l)
A suitable indicator is necessary→ need one that is one colour in acids and another in alkalis (universal indictor is NOT a good choice because it is hard to determine the exact point at which the acid is neutralised).
Equation to know:
Concentration (mol/dm³) = $\dfrac{\textbf{amount of substance (mol)}}{\textbf{volume (dm}^3\textbf{)}}$





Procedure:
1. Strong acid of known concentration is in burette
2. Alkali to be neutralised of known volume is in conical flask
3. Add indictor to conical flask
4. Note start point on burette
5. Add acid drop-wise to conical flask.
6. As soon as indicator changes colour stop
7. Note end point on burette and calculate volume added
8. Carry out analysis using equation triangle

NB procedure above can be used with an alkali of known concentration in burette and acid to be neutralised in conical flask.

19. Electrolysis

What is electrolysis?	**Electrolysis** is the process by which ionic substances are decomposed (broken down) into simpler substances when an electric current is passed through them.
What happens to the motion of ions when melted or in solution?	The are free to move
What happens when electricity is passed through the solution or molten ionic compound?	Ions are charged therefore the: - **Cations** (positive ions) move towards **the cathode** (negative electrode) - **Anions** (negative ions) move towards the **anode** (positive electrode)
What do you call the ions in solution that conduct electricity?	Electrolytes
What happens once the ions get to the electrodes?	Ions are **discharged** - forming elements
What is produced at the anode?	Non-metal is produced (Anions transfer electrons onto the anode)
What is produced at the cathode?	Metal is produced (Cations gain electrons from the cathode)

Electrolyte	Ions (molten or in solution) that are free to move and can therefore pass electricity through them
Electrode	The conductors through which electrical current is transferred to the ionic solution or molten ionic compound
Anode	Positive electrode
Cathode	Negative electrode
Anion	Negatively charged particles
Cation	Positively charged particles
Discharge	Transferal of charge (electrons) at an electrode result in ion returning to elemental (neutral) form