**Group 1 / 2 = alkali metals (LEFT)**

**Group 2-3 = transitional metals**

**Group 4-7 = non-metals**

**Group 7 = halogens**

**Group 0 = noble gases (RIGHT)**

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**CC4B - ATOMIC NUMBER AND THE PERIODIC TABLE**

- End of 19th century, noble gases discovered
- Hadn’t predicted existence due to being inert
- Pair reversals (iodine and tellurium) not explained until 1913, Moseley

**Atomic number –**

- Moseley showed an elements position was based on its physical properties
- Fired high-energy electrons at elements, making them give off x-rays
- Discovered for every step increase of atomic number, change in energy of x-rays
- Realised atomic number was equal to number of positive charges in nucleus
- Proton then discovered shortly after
- Therefore proved atomic number to be equal to number of protons in nucleus

**CC4C - ELECTRON CONFIGURATION**

- Electrons occupy electron shells, arranged around the nucleus
- Arrangement is known as electron configuration

**For first 20 elements**

- 1st shell = 2 electrons
- 2nd shell = 8 electrons
- 3rd shell = 8 electrons

- **Vertical column (group)** indicates how many electrons are in outer shell
- **Horizontal column (period)** indicates how many shells there are
- Can be calculated using atomic number of an element – fill shells

**ATOM** - what makes up an element, consists of protons, neutrons, and electrons on the outer shells

**ELEMENT** – substance made out of the same atoms (O2)

**COMPOUND** – substance made out of different atoms that are chemically bonded (CO2, H2O)

**MIXTURE** - substance consisting of different atoms that aren’t chemically bonded (Salt water)
Single covalent bond – only one electron is shared

Double covalent bond – two electrons are shared
• Used to stabilise atoms by completing their outer shell

**MOLECULAR SUBSTANCES** – group of atoms held together by covalent bonds

**VALENCE** – number of covalent bonds formed by atoms / number of electrons needed to complete the outer shell

<table>
<thead>
<tr>
<th>Group</th>
<th>Outer Electrons</th>
<th>Bonds Formed</th>
<th>Valence</th>
</tr>
</thead>
<tbody>
<tr>
<td>5</td>
<td>N and P</td>
<td>4</td>
<td>4</td>
</tr>
<tr>
<td>6</td>
<td>O and S</td>
<td>2</td>
<td>2</td>
</tr>
<tr>
<td>7</td>
<td>F and Cl</td>
<td>1</td>
<td>1</td>
</tr>
</tbody>
</table>

Diagram E shows how molecular formulae can be worked out by matching up the valences, so that all atoms have the correct number of bonds (and so a complete outer electron shell).

S has a valency of 2 so it forms 2 bonds
C has a valency of 4 so it forms 4 bonds

So two S atoms each form a double bond with a single C atom. As a result all atoms form the correct number of bonds.

(You can also swap the valencies between elements)
PROPERTIES OF METALS

Properties -
- Malleable – used for infrastructure/building
- Good conductors of electricity – used as wires in circuits (copper)
- Shiny – making jewellery
- Very high melting and boiling points
- High density

Metallic bonding -
- Metals tend to have 1-3 electrons on their outer shell, which become delocalised meaning they are free to move around
- This establishes positive ions (due to loss of electrons) in a sea of negative, delocalised electrons
- Electrostatic attraction between cations and anions

Metallic structure –
- Atoms packed together in lattice structure
- Stacked layers of positive ions
- Sea of delocalised negative electrons that move around feely
**EMPIRICAL FORMULA** – simplest whole number ratio of atoms/ions of each element in it

1. MASS OF EACH COMPOUND/RELATIVE ATOMIC MASS
2. DIVIDE BOTH BY SMALLEST NUMBER
3. MULTIPLY NUMBERS IF NOT WHOLE NUMBERS

<table>
<thead>
<tr>
<th></th>
<th>CA</th>
<th>CL</th>
</tr>
</thead>
<tbody>
<tr>
<td>MASS</td>
<td>10g</td>
<td>17.8g</td>
</tr>
<tr>
<td>RELATIVE ATOMIC MASS</td>
<td>40</td>
<td>35.5</td>
</tr>
<tr>
<td>MASS/RAM</td>
<td>0.25</td>
<td>0.5</td>
</tr>
<tr>
<td>DIVISION BY SMALLEST NUMBER</td>
<td>0.25/0.25 = 1</td>
<td>0.5/0.25 = 2</td>
</tr>
<tr>
<td>CA = 1</td>
<td>CL = 2</td>
<td></td>
</tr>
</tbody>
</table>

= CaCl₂

**MOLECULAR FORMULA** - how many atoms of each element are actually in a molecule

**RELATIVE FORMULA MASS/ EMPIRICAL FORMULA MASS**

MULTIPLY EMPIRICAL FORMULA BY THE RESULT

1. GLUCOSE = CH₂O
2. RFM = 180
3. EFM = CH₂O = (12X1)+(1X2)+(16X1) = 30
4. 180/30 = 6
5. NOW TIMES EVERYTHING IN THE EMPIRICAL FORMULA BY 6 – C₆H₁₂O₆

CC9B – CONSERVATION OF MASS

THE LAW OF CONSERVATION OF MASS – mass can’t be created nor destroyed, only rearranged (total mass of reactants = total mass of products)

- When a solute is dissolved into a solvent, the mass of the resulting solution is the mass of both the solute + solvent – no atoms are lost long the way

**CONCENTRATION** – the amount of solute dissolved in a given volume of solution, measured in – GMD-3

1DM³ = 1 LITRE/ 1000CM³

**CONCENTRATION** = mass of solute (g)/ mass of solution

**CLOSED REACTIONS** – enclosed in a box, nothing can escape therefore mass will read the same as before reaction occurred

**UN-ENCLOSED REACTIONS** – gas can escape during reaction therefore mass can decrease due to it being elsewhere
IONIC COMPOUNDS DISSOLVED IN WATER –

- Water can be ionised too
- H+ and OH- ions present in water
- Products formed at electrodes depend on whether water ions discharge more easily than salt ions
- Ionic compound with metal more reactive than hydrogen will be replaced with H+ (produced instead of metal) due to being discharged more difficultly

Sodium chloride -> Na+ more reactive than H+ therefore H+ replaces Na+ and goes to negative electrode, Na+ stays in solution

CC11A – REACTIVITY

REACTIVITY SERIES – list of metals in order of reactivity, most reactive at the top
SPECTATOR IONS – ions that remain the same during a reaction
HALF EQUATIONS – way of representing the change of electrons (from ionic equation)
DISPLACEMENT REACTION – whereby a more reactive metal takes place of a less reactive metal compound, being a redox reaction (both oxidation and reduction occur)

Metals + cold water -> hydrogen + metal hydroxide
Metals + steam -> hydrogen + metal oxide
Metals + dilute acid -> hydrogen + salt solution (first name from metal, second from acid)

Equation – Zn + CuSO₄ -> Cu + ZnSO₄ (zinc displaced copper)

Ionic – Zn + Cu²⁺ + SO₄²⁻ -> Cu + Zn²⁺ + SO₄²⁻

Zn + Cu²⁺ -> Cu + Zn²⁺

Half-equation – Zn -> Zn²⁺ + 2e (OXIDATION)

Cu²⁺ + 2e -> Cu (REDUCTION)
Group 7 = halogens

- All have seven electrons on outer shell, meaning they gain one to become stable (1-)
- Bad conductors
- Diatomic structure (two atoms held by single, covalent bond)
- Often used as disinfectants/bleach/cleaning products
- Reactivity decreases as you go down series due to the increased distance making the electrostatic force of attraction stronger
  - harder to break and weaker reactivity (due to trying to gain electron)

(Fluorine, chlorine, bromine, iodine, astatine)

<table>
<thead>
<tr>
<th>HALOGEN</th>
<th>RELATIVE SIZE</th>
<th>MELTING POINT</th>
<th>BOILING POINT</th>
<th>STATE AT ROOM TEMP</th>
</tr>
</thead>
<tbody>
<tr>
<td>Fluorine</td>
<td>-1</td>
<td>-220</td>
<td>-118</td>
<td>Gas</td>
</tr>
<tr>
<td>chlorine</td>
<td>0</td>
<td>-101</td>
<td>-34</td>
<td>Gas</td>
</tr>
<tr>
<td>Bromine</td>
<td>1</td>
<td>-7</td>
<td>59</td>
<td>Liquid</td>
</tr>
<tr>
<td>Iodine</td>
<td>2</td>
<td>144</td>
<td>184</td>
<td>Solid</td>
</tr>
</tbody>
</table>

- React with metals forming ionic compounds (contain halide ions – x)

Reactions with hydrogen –

- Halogens react with non-metals by sharing electrons and forming covalent compounds
- These gases are extremely soluble in water and dissolve to produce acids (aqueous)
- Halogen + hydrogen = hydrogen halides
- Can convert hydrogen halide to its acid by dissolving it in water

\[ \text{H}_2 + \text{Cl}_2 = 2\text{HCl} \] – hydrochloric acid (aqueous when dissolved in water)

Test for chlorine –

- Turns blue litmus paper red, then white
- Turns bleach white
CC13C - HALOGEN REACTIVITY

Halogens + metal = halide salts

**DISPLACEMENT REACTION** – whereby a more reactive element replaces a less reactive element in a compound (more reactive halogen replaces less reactive to form halide compound)

"Chlorine + sodium bromide = bromide + sodium chloride"

**REDOX REACTION** – reaction in which both oxidation and reduction occurs – (OIL,RIG)

**OXIDATION** – the loss of electrons + gain of oxygen

**REDUCTION** – the gain of electrons

"2Na (s) + Br₂ (g) = 2NaBr (s)"

- Sodium oxidised, lost electron to bromine, bromine reduced, gained electron from sodium

<table>
<thead>
<tr>
<th>Halogen</th>
<th>Effect on iron wool</th>
</tr>
</thead>
<tbody>
<tr>
<td>fluorine</td>
<td>bursts into flames</td>
</tr>
<tr>
<td>chlorine</td>
<td>glows brightly</td>
</tr>
<tr>
<td>bromine</td>
<td>glows dull red</td>
</tr>
<tr>
<td>iodine</td>
<td>changes colour</td>
</tr>
</tbody>
</table>

CC13D - GROUP 0 – NOBEL GASES

- 8 electrons in outer shell – no delocalised electrons
- Monatomic (singular atoms)
- Inert (non-reactive) – no electrons to lose or gain
- Non-metals
- Low boiling and melting points
- Bad conductors of electricity (no free electrons to move around)
- Reactivity remains the same throughout series
- As you move down series, density increases due to more outer shells
- Used to be group zero however helium didn’t fit into requirement
- Weren’t in first periodic table due to being so unreactive, couldn’t be detected

1. KYRPOTON – used in fluorescent lights, photography flash and lasers (reacts with fluorine)
2. ARGON – used in wine barrels to prevent wine oxidising (more dense than air)
3. HELIUM – used in weather balloons and airships (low density so floats and non-flammable)
4. NEON – long lasting illuminated signs (produces red/orange light when current passed through)
CC14A - RATES OF REACTION

CHEMICAL REACTION - when one or more reactants form one or more products

RATE OF REACTION - speed of which reactants are turned into products (frequency of collisions and amount of energy needed)

Chemical reactions –

- Colour change
- New product formed (precipitate – two soluble substances producing insoluble solid)
- Effervescence (gas formed)
- Irreversible
- Temperature change

(No mass lost or gained in reaction, only rearranged)

- Rusting/eroding = slow reaction
- Explosions/ potassium & water = quick reaction

• Steeper the slope – faster the reaction (gradient)
• No slope/flat line – all energy used up and reaction complete
• Greatest speed of reactant in beginning due to highest concentration of reactants available
• Concentration of reactants decrease whilst products increase
• Reactions don’t proceed at steady rate
• Gas syringes used to measure rate of reaction as traps product and measures production in given time

CC14B - FACTORS THAT AFFECT RATE OF REACTION

• For reaction to occur, atoms of reactants must collide with one another with enough energy

ACTIVATION ENERGY – minimum amount of energy required for a reaction to occur