

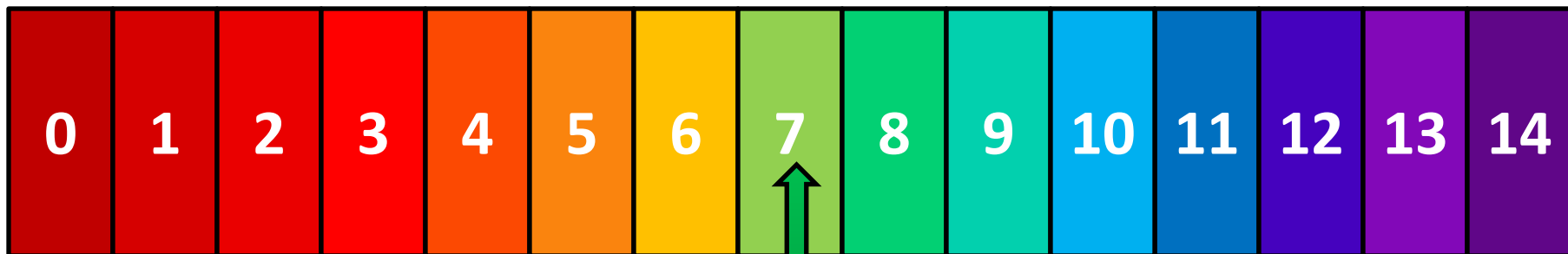
3rd Form Revision Notes

Indicators

pH Scale

← Acids get stronger

Alkalis get stronger →



← More acidic

More alkaline →

↑
Neutral

Indicators

Universal indicator is made from a range of dyes which change colour in a gradual way over a **range of pHs**. It changes through a variety of colours from **pH 1 to pH 14**, but is **not very accurate**.

Litmus is one of the **commonest indicators**; in **acidic solutions it is red**, in **alkaline solutions it is blue**, and in **neutral solutions** it is an equal mixture of the two forms, and shows **purple**.

Reactions of Acids

Acid + **alkali** → Salt + water

Acid + **carbonate** → Salt + carbon dioxide + water

Acid + **metal** → Salt + hydrogen

Naming Salts

When naming a salt, the first part of the name comes from the **metal ion** present. The second part comes from the **parent acid**.

e.g. Copper + Sulfuric Acid → Copper Sulfate + Hydrogen

Parent Acid	Salts
Hydrochloric(HCl)	Chlorides (Cl ⁻)
Sulfuric (H ₂ SO ₄)	Sulfates (SO ₄ ²⁻)
Nitric, (HNO ₃)	Nitrates (NO ₃ ⁻)
Phosphoric, (H ₃ PO ₄)	Phosphates (PO ₄ ³⁻)

Balancing Equations

- 1) **Determine** number of atoms for each element.
- 2) **Pick** an element that is not equal on both sides of the equation.
- 3) **Add** a coefficient in front of the formula with that element and adjust your counts.
- 4) **Continue** adding coefficients to get the same number of atoms of each element on each side.

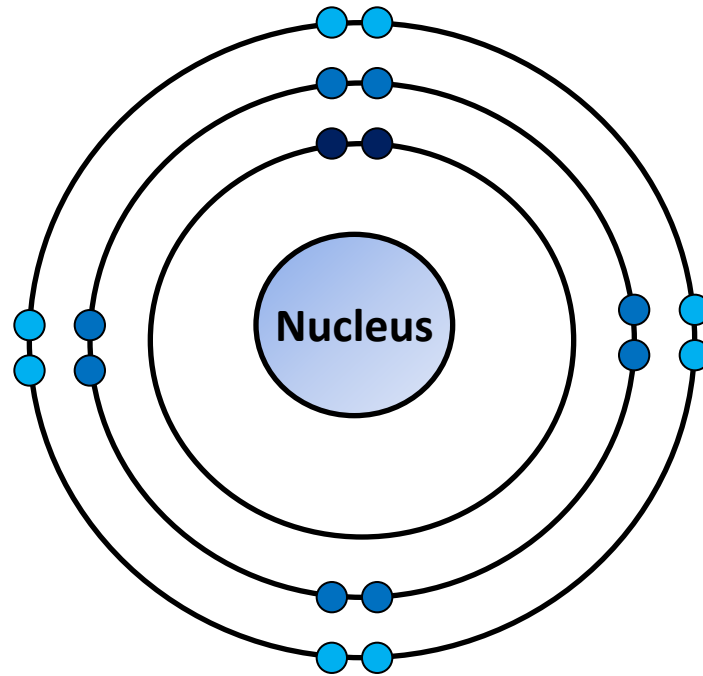
Atomic Structure

Sub-atomic Particle	Relative Mass	Relative Charge
Proton	1	+1
Neutron	1	0
Electron	$\sim 1/2000$	-1

Atomic number = number of protons
= number of electrons (in a neutral atom)

Mass number = number of protons + number of neutrons

Atomic Structure



Nucleus contains protons and neutrons – held together by strong nuclear force

Electrons exist in energy levels (orbits/shells) – 2, 8, 8 etc.

Groups 1: The Alkali Metals

1 electron in outer energy level

Down the group, **atomic radius increases**

Electrostatic attraction \propto atomic radius⁻²

Down the group, the number of **energy levels increase, shielding** of positive nucleus **increases**

Groups 1: The Alkali Metals

To obtain **stable electronic configuration** must **lose 1 electron**.

Down group, reactivity increases:

- **Shielding increases → easier to lose electron**
- **Less electrostatic attraction between outer electron and positive nucleus**

Groups 7: The Halogens

7 electrons in outer energy level

Down the group, **atomic radius increases**

Electrostatic attraction \propto atomic radius⁻²

Down the group, the number of **energy levels increase, shielding** of positive nucleus **increases**

Groups 7: The Halogens

To obtain **stable electronic configuration** must **gain 1 electron**.

Down group, reactivity decreases:

- **Shielding increases → harder to gain electron**
- **Less electrostatic attraction** between **incoming electron** and **positive nucleus**

Solubility

Solubility is the amount of solute (in grams) that will dissolve in 100g of solvent

$$\text{Solubility} = \frac{\text{mass of solute}}{\text{mass of solvent}} \times 100$$

Solubility

As **temperature increases** solubility of solute dissolved in solvents **increases**.

As **temperature increases** solubility of gases dissolved in solvents **decreases**

Solubility Rules

Nitrates: All nitrates are soluble

Sulfates: All sulfates are soluble except **lead and barium**

Chlorides: All chlorides (and other halides) are soluble except **lead and silver**

Hydroxides: All hydroxides are **insoluble** except **potassium, sodium and ammonium**

Carbonates: All carbonates are **insoluble** except **potassium, sodium and ammonium.**

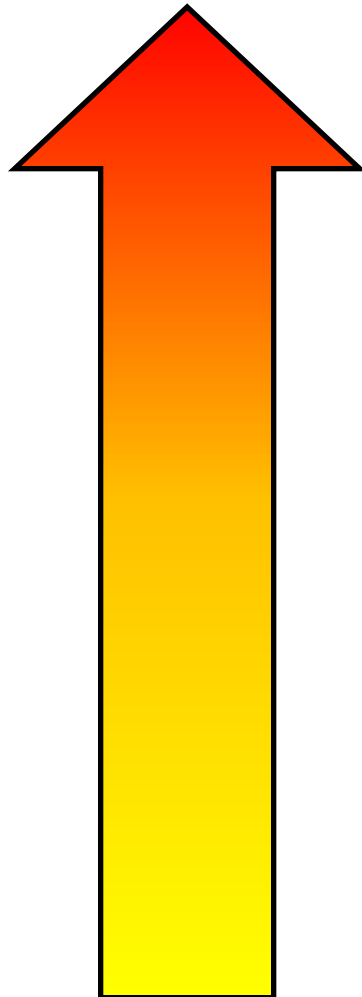
Reactivity Series

More reactive metals will displace less reactive salts from their counter ions.

e.g.

Magnesium + copper sulfate \rightarrow Copper + Magnesium sulfate

Reactivity Series



Potassium
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Tin
Lead
Copper
Silver
Gold
Platinum

Most Reactive

Least Reactive

Simple Electrolysis

Electrolysis is the process of **electrically splitting** an **ionic compound** to produce **new products**.

Positive ions (cations) move to **cathode** (-ve electrode)

Negative ions (anions) move to **anode** (+ve electrode)

Simple Electrolysis

OIL RIG

Oxidation is **loss** of electrons

Reduction is **gain** of electrons

Oxidation occurs at **anode**

Reduction occurs at **cathode**

(Remember, vowels together, consonants together)



Simple Electrolysis

Half equations for electrolysis of $\text{CuCl}_{2(\text{aq})}$

At the **CATHODE** (-):



At the **ANODE** (+):



Tips for half equations:

1. Always start by writing reactants and products
2. Balance in terms of number of atoms
3. Balance in terms of charge by adding electrons on the appropriate side

Extraction of Metals

Rocks from which metals can be extracted are called **ores**. These rocks are usually a mixture of **compounds**. These have to be broken down in order to **extract** the metal.

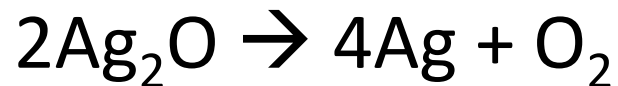
Only very unreactive metals are found **native**, in their elemental state. Gold, silver and copper are sometimes found this way.

Extraction of Metals

Extraction By Heat Alone – Thermal Decomposition

Heat alone can **reduce** some less stable metal ores. Only the **least** reactive metals can be extracted using this method:

Silver oxide \rightarrow silver + oxygen



Extraction of Metals

Extraction by Heating with Carbon

Most ores are **oxides** or can easily be made into **oxides** by **heating in air**. Therefore, they only need to be **reduced** (have the oxygen removed).

Carbon acts as a **reducing agent**. The carbon and the oxygen react to form carbon dioxide, leaving the metal.

Rates of Reaction

Greater than the activation energy

Chemical reactions occur when particles of reactant **collide** with **enough energy** to react.

If they do not have enough energy they simply bounce apart.

The more **concentrated** the solutions, the faster the reaction, due to the increased number of collisions.

Increasing the **surface area** of a reactant, increases the rate of reaction. The reactions of solids can clearly only take place at the surface of the solid.

For gaseous reactions the same arguments apply for **pressure**

At higher **temperatures** molecules move faster. As a result there are more collisions per second and so a faster reaction occurs.

Rates of Reaction

Chemical reactions involve the formation of bonds between atoms but often before new bonds can be formed old ones have to be broken.

This means that there has to be enough energy (**activation energy**) to start breaking the old bonds before a reaction can occur.

Catalysts increase the rate of reaction without being used up themselves in the reaction.

They do this by providing an alternative, lower energy reaction pathway.

