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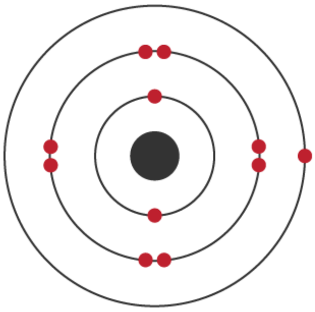
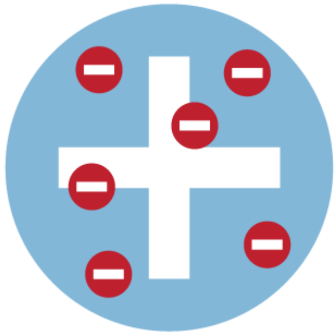
The electronic structure of an atom can be predicted from its atomic number. For example, the atomic number of sodium is 11. Sodium atoms have 11 *protons* and so 11 electrons:



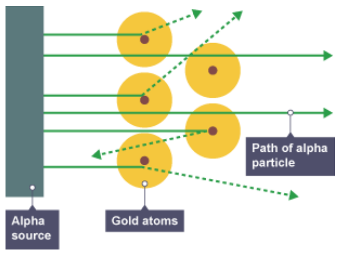
- two electrons occupy the first shell
- eight electrons occupy the second shell
- one electron occupies the third shell

This electronic structure can be written as 2,8,1 (each comma, or dot, separates one shell from the next). This electronic structure can also be shown as a diagram. In these diagrams: each shell is shown as a circle, each electron is shown as a dot or a cross.

Mendeleev left gaps in his table for elements not known at the time. By looking at the properties of the elements next to a gap, he could also predict the properties of these undiscovered elements.
J J Thomson carried out experiments and discovered the *electron*. This led him to suggest the *plum pudding model* of the atom. In this model, the atom is a ball of positive charge with negative electrons embedded in it - like currants in a Christmas pudding.

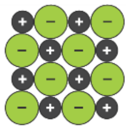


<u>The plum pudding model</u>
In 1909 Ernest Rutherford designed an experiment to test the plum pudding model. In the experiment, positively charged *alpha particles* were fired at thin gold foil. Most alpha particles went straight through the foil. But a few were scattered in different directions.
The alpha particle scattering experiment
This evidence led Rutherford to suggest a new model for the atom, called the *nuclear model*. In the nuclear model:

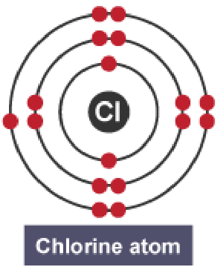


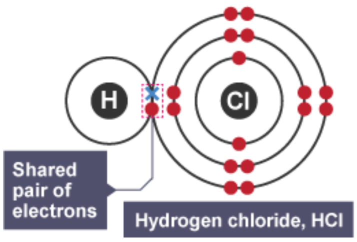
- the mass of an atom is concentrated at its centre, the *nucleus*
- the nucleus is positively charged

An *ionic compound* is a giant structure of ions. The ions have a regular, repeating arrangement called an *ionic lattice*. The lattice is formed because the ions attract each other and form a regular pattern with oppositely charged ions next to each other.

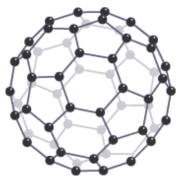


An ionic lattice is held together by strong *electrostatic forces* of attraction between the oppositely charged ions. The forces act in all directions in the lattice. This is called *ionic bonding*





A *covalent bond* is formed when two *atoms* share a pair of *electrons*. Covalent bonding occurs in most *non-metal elements*, and in *compounds* formed between non-metals.
These shared electrons are found in the outer shells of the atoms. Usually each atom contributes one electron to the shared pair of electrons. *Graphene* and *fullerenes* are forms of carbon. Their structures are different from those of *diamond* and *graphite*, which are also forms of carbon.
Graphene Graphene is a single layer of graphite. The strong *covalent bonds* between the carbon atoms mean that graphene: has a very high *melting point* is very strong
Like graphite, graphene conducts electricity well because it has *delocalised electrons* that are free to move across its surface.
These *properties* make graphene useful in electronics and for making *composites*.



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The rate of a reaction is a measure of how quickly a *reactant* is used up, or a *product* is formed.
Collision theory For a chemical reaction to happen:

- *reactant particles* must collide with each other
- the particles must have enough energy for them to react

A collision that produces a reaction is called a *successful collision*. The *activation energy* is the minimum amount of *energy* needed for a collision to be successful. It is different for different reactions. The greater the *frequency* of *successful collisions*, the greater the rate of reaction. If the *temperature* of the reaction mixture is increased:

- *reactant particles* move more quickly
- the *energy* of the particles increases
- the **frequency** of successful collisions between reactant particles increases
- the **proportion** of collisions which are successful increases
- the rate of reaction increases

Neutralisation

Bases and alkalis

A *base* is any substance that reacts with an *acid* to form a *salt* and water only. This means that metal oxides and metal hydroxides are bases. Bases that are *soluble* in water are called *alkalis* and they *dissolve* in water to form *alkaline solutions*. For example: copper oxide is a base, but it is not an alkali because it is *insoluble* in water, sodium hydroxide is a base, and it dissolves in water so it is also an alkali

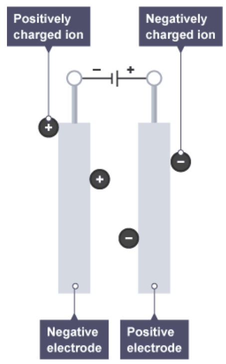
A *neutralisation* reaction is a reaction between an acid and a base. Remember:
acids in solution are sources of hydrogen *ions*, H^+ alkalis in solution are sources of hydroxide ions, OH^-
In acid-alkali neutralisation reactions, hydrogen ions from the acid react with hydroxide ions from the alkali: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$ Pure water is *neutral* (its *pH* is 7). A neutral solution can be produced if the correct amounts of acid and alkali react together. The change in pH during a neutralisation reaction can be measured using a pH probe and meter, or estimated using *universal indicator* solution and a pH colour chart.

Naming salts

The name of a salt has two parts. The first part comes from the base, alkali or metal carbonate. The second part comes from the acid: **hydrochloric acid produces chloride salts** nitric acid produces nitrate salts **sulfuric acid produces sulfate salts**

Electrodes and ions

The negatively charged electrode in electrolysis is called the *cathode*. Positively charged ions move towards the cathode.
The positively charged electrode in electrolysis is called the *anode*. Negatively charged ions move towards the anode.
Ions move towards their oppositely charged electrode



Products of electrolysis

When ions reach an electrode, they gain or lose *electrons*. As a result, they form *atoms* or *molecules* of *elements*:

- positive ions gain electrons from the negatively charged cathode
- negative ions lose electrons at the positively charged anode

When a chemical reaction happens, energy is transferred to or from the surroundings. When energy is transferred to the surroundings, this is called an *exothermic* reaction, and the temperature of the surroundings increases. Examples of exothermic reactions include: *combustion* reactions, many *oxidation* reactions, most *neutralisation* reactions
Everyday uses of exothermic reactions include self-heating cans and hand warmers. When energy is taken in from the surroundings, this is called an *endothermic* reaction and the temperature of the surroundings decreases. Examples of endothermic reactions include: *thermal decomposition* reactions, the reaction of citric acid and sodium hydrogen carbonate