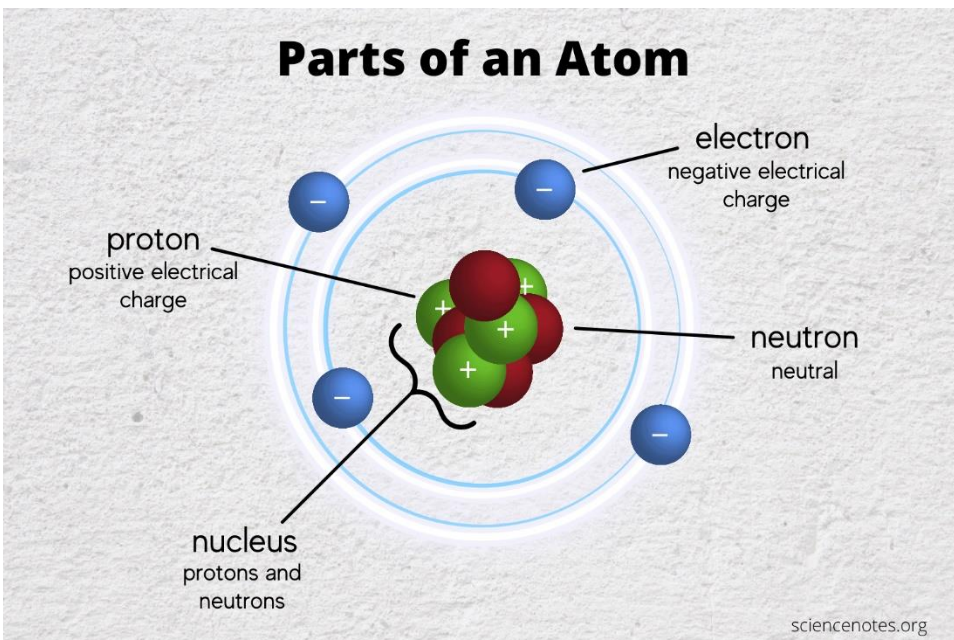
Part 1:

Atomic Structure

Positions, charges and masses of subatomic particles

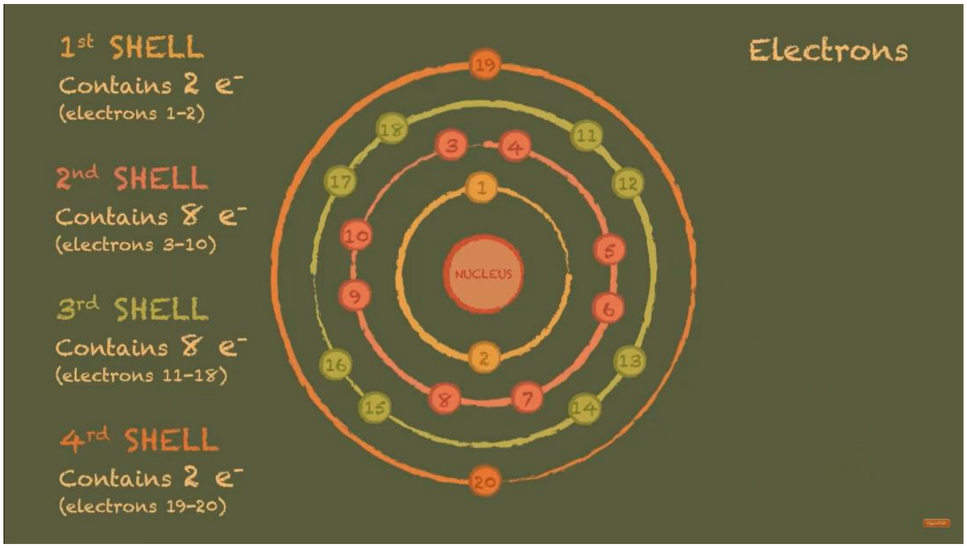
Subatomic particle	Symbol	Relative charge	Location	Amu mass
Proton	p	+1	inside the nucleus	1
Neutron	n	0	inside the nucleus	1
Electron	e^-	-1	outside the nucleus	0.0005 (~0)



Mass and atomic number

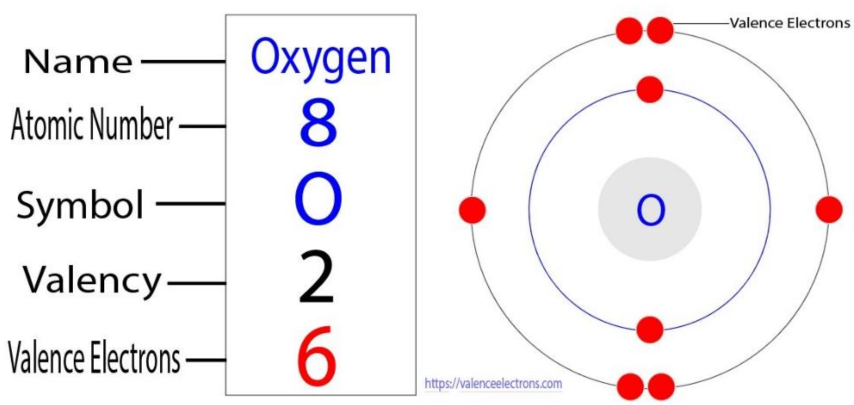
The atomic number is the number of protons in an element, the element is determined by the number of particles (e.g. all atoms with 6 protons are carbon atoms). The mass number is the number of protons + neutrons. However, the number of neutrons can vary, forming atoms known as isotopes.

Drawing electronic configurations (until Z=20)



Determine valency of electrons

The number of valence electrons = the number of electrons in the outer shell of an atom e.g



Evaluate the models for atomic theory

Dalton – Proposed that all matter is made of atoms

Thompson – Discovers electron and proposes plum pudding model (positively charges sphere with electrons scattered throughout)

Rutherford – Proposes nuclear model (positive charge and most of the mass is concentrated in a small nucleus with electrons orbiting around)

Bohr – Introduces his model which includes the idea of quantified energy levels

Schrodinger – develops wave equation, describing the wave-like behavior of particles leading to the development of quantum physics

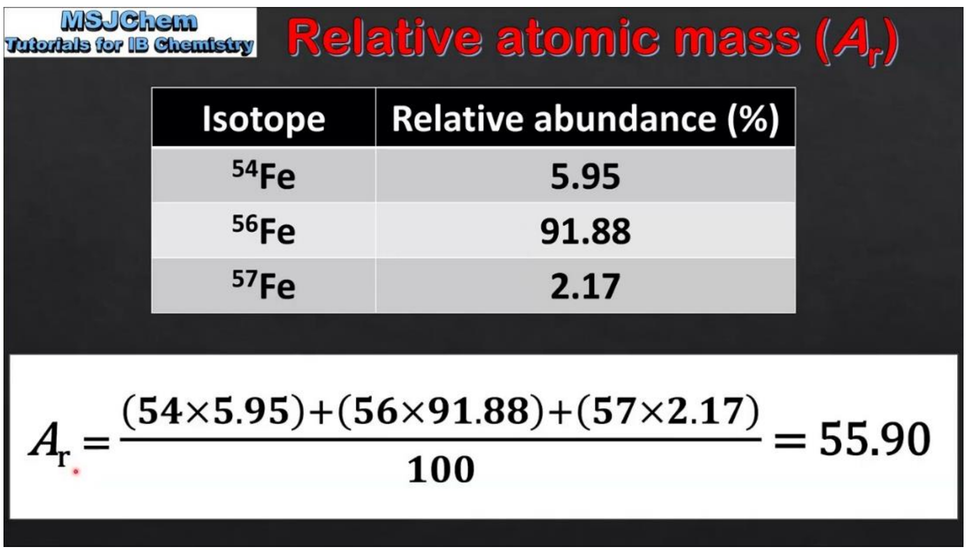
Chadwick – Discovers the neutron completing the model of the atom

Isotopes

Isotopes are different forms of an element that have the same number of protons but different numbers of neutrons. An example of an isotope would be the isotopes of Uranium such as U-235, which are used as a fossil fuel in nuclear reactors.

Relative atomic mass based on percentage abundance of isotopes

AVG=$(m_1p_1)+(m_2p_2)+(m_3p_3)/100$



Isotope	Relative abundance (%)
^{54}Fe	5.95
^{56}Fe	91.88
^{57}Fe	2.17

$$A_r = \frac{(54\times5.95)+(56\times91.88)+(57\times2.17)}{100} = 55.90$$

<u>Part 2</u>

<u>The periodic table</u>

Groups and Periods

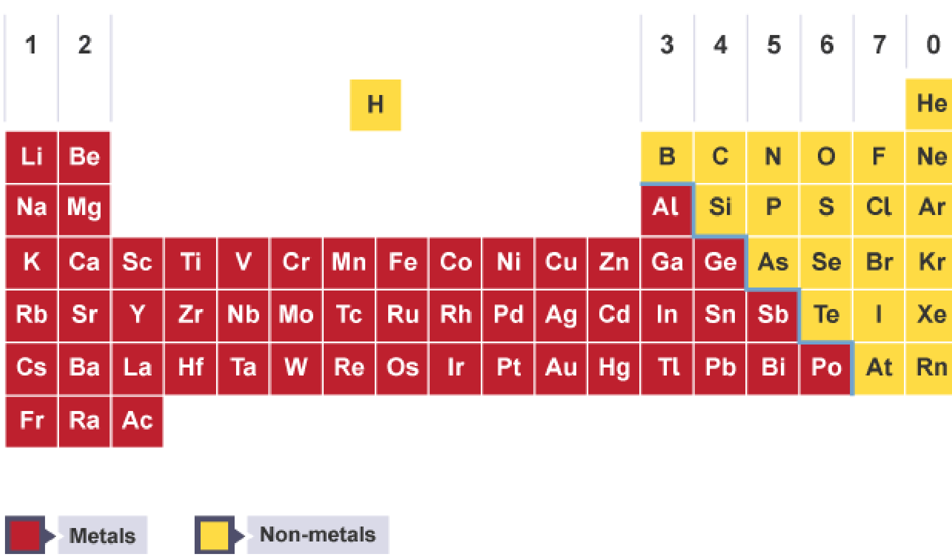
The vertical columns in the periodic table are known as groups. The elements of the groups exhibit similar chemical and physical properties. The horizontal rows are called periods, the periodic number correlates to the energy level of the element.

the number of circles in the electronic configuration of an element is represented in the periodic table as the period number that element is situated in. The number of electrons in the outermost shell of an element is represented in the periodic table as the group number that element is situated in.

History of periodic table

- Lavoisier: developed the first list of elements
- Döbereiner: observed that some elements could be grouped into triads
- Newlands: proposed the law of octaves, which arranged elements in groups of eight with similar properties
- Mendeleev: developed the first periodic table based on atomic weight and left gaps for missing elements
- Moseley: discovered the concept of atomic number and rearranged the periodic table based on this fundamental property
- Modern Periodic Table: based on the periodic law that the physical and chemical properties of elements are periodic functions of their atomic number, arranged in rows and columns with similar properties.

<u>Metals and Non-metals</u>



Properties of metals and non-metals compared to metalloids

Physical properties:

Metals	Non-metals
Shiny	Dull
High melting points	Low melting points
Good conductors of electricity	Poor conductors of electricity
Good conductors of heat	Poor conductors of heat
High density	Low density
Malleable and ductile	Brittle

Chemical properties:

The most common chemical property is the type of **oxide** that the element forms. Metals form oxides that are **basic**, but non-metals form oxides that are **acidic**. For example, sulfur and carbon are both non-metals. They react with oxygen to form sulfur dioxide and carbon dioxide. These **compounds** are both gases present in the air and which dissolve in rain water, making it acidic.

Metalloids:

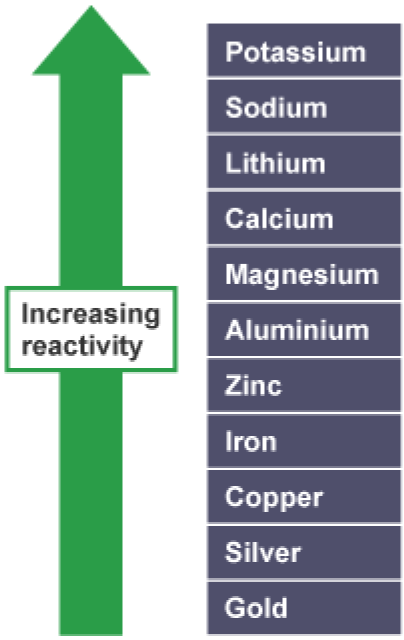
Some elements between the metals and non-metals in the periodic table have properties which are a mixture of the properties of metals and non-metals. These elements are called **metalloids** or **semi-metals**.These elements are found close to the zig-zag line that separates the metals from the non-metals.

Examples of metalloids

- Carbon – a dull, dark grey solid and is brittle (non-metallic properties) but has a high melting point and is a conductor of electricity and heat (metallic properties).
- Silicon – a shiny grey solid, an electrical conductor and has a high melting point (metallic properties) but is very brittle and has a low density (non-metallic properties).

Reactivity series

In a reactivity series, the most reactive element is placed at the top and the least reactive element at the bottom. More reactive metals have a greater tendency to lose electrons and form positive ions.



- **Very reactive** metals, such as aluminium, form **stable** oxides and other compounds. **<u>Electrolysis</u>** is commonly used to extract these metals and requires a lot of electric current (energy) to **<u>reduce</u>** them to extract the metal.
- **Less reactive** metals, such as iron, form **less stable** oxides and other compounds. Reduction with carbon is often used to extract these metals and requires less energy to reduce them to extract the metal.

Groups in the periodic table

Properties

Group 1: This group consists of alkali metals that have low electronegativity and readily lose their outermost electron to form a +1 ion. They exhibit high electrical conductivity and malleability due to the presence of a single valence electron that is loosely bound to the nucleus, enabling easy movement of electrons within the metal lattice.

Group 7: This group consists of halogens that exist in all three states of matter at room temperature depending on their atomic number. They have high electronegativity and reactivity with metals, making them strong oxidizing agents. They can readily form ionic bonds with metals to produce salts, with the halogens acting as the anion.

Group 8: This group consists of noble gases that have a stable octet electronic configuration and are chemically inert. This is because their valence shell is complete, making them unreactive and stable in their elemental form. They have low boiling and melting points, and are mostly found in the gaseous state at room temperature.

Periodic trends

Group 1: As you move down Group 1, the atoms get larger, and the outermost electron becomes further away from the nucleus. This makes it easier for the electron to be lost, increasing the reactivity of the group. As a result, the alkali metals in Group 1 become increasingly reactive and have lower ionization energies, making it easier to remove the outermost electron. They also exhibit increasingly lower melting and boiling points.

Group 7: As you move down Group 7, the atoms get larger, and the outermost electron becomes further away from the nucleus. This decreases the electronegativity of the group, making it harder to attract an electron, and hence making the halogens less reactive. At the same time, the melting and boiling points of the halogens increase down the group.

Group 8: The elements in Group 8 are called noble gases, and they have a complete outermost electron shell, making them inert. As you move down the group, the atomic size increases, but there are no changes in reactivity as the outermost electron shell remains complete. The boiling points of the noble gases increase as you move down the group, reflecting the increase in the number of electrons and the larger size of the atoms.

Ions and valency

Deduce the ions formed when groups 1, 2, 3 lose electrons:

- Group 1 elements (such as sodium and potassium) will lose one electron to form a +1 cation.
- Group 2 elements (such as magnesium and calcium) will lose two electrons to form a +2 cation.
- Group 3 elements (such as aluminum) will lose three electrons to form a +3 cation.

Deduce the ions formed when groups 5, 6, 7 gain electrons:

- Group 5 elements (such as nitrogen) will gain three electrons to form a -3 anion.
- Group 6 elements (such as oxygen) will gain two electrons to form a -2 anion.
- Group 7 elements (such as chlorine) will gain one electron to form a -1 anion.

Identify the charges on atoms in a compound:

The charges on atoms in a compound are determined by their valence electrons, which are the outermost electrons involved in chemical bonding. Generally, atoms gain or lose electrons to achieve a full valence shell of eight electrons, or two electrons for hydrogen and helium. The charges on atoms in a compound can be determined by balancing the positive and negative charges in the compound, ensuring that the compound is neutral overall.

State that transition elements can form more than one ion, including examples:

Transition elements can form more than one ion because they have partially filled d-orbitals, which can lose or gain electrons to form ions with different charges. For example, iron can form both Fe2+ and

Fe3+ ions, copper can form both Cu+ and Cu2+ ions, and manganese can form Mn2+, Mn3+, Mn4+, and Mn7+ ions.

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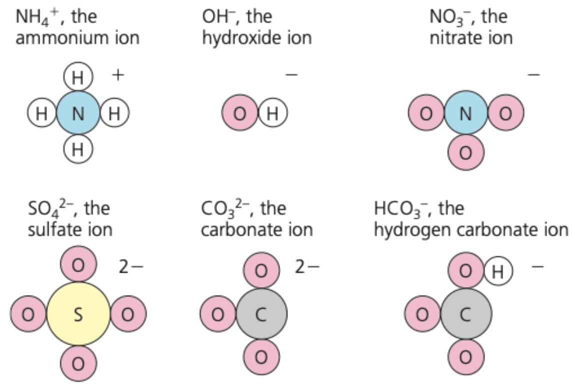
Charges of polyatomic ions

Compound ions

All the ions you met so far have been formed from single atoms. But ions can also be formed from a **group** of bonded atoms. These are called **compound ions**.

The most common ones are shown on the right. Remember, each is just one ion, even though it contains more than one atom.

The formulae for their compounds can be worked out as before. Some examples are shown below.



Example 3

1. Sodium carbonate.
2. The ions are Na^+ and CO_3^{2-}.
3. Two Na^+ are needed to balance the charge on one CO_3^{2-}.
4. The formula is Na_2CO_3.

Example 4

1. Calcium nitrate.
2. The ions are Ca^{2+} and NO_3^-.
3. Two NO_3^- are needed to balance the charge on one Ca^{2+}.
4. The formula is $Ca(NO_3)_2$. Note that brackets are put round the NO_3, before the $_2$ is put in.