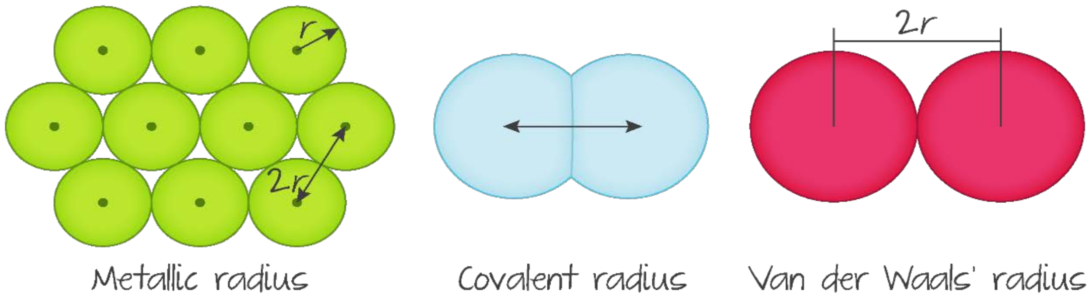
Periodicity in Pictures and Definitions

1. Different measurements of atomic radius

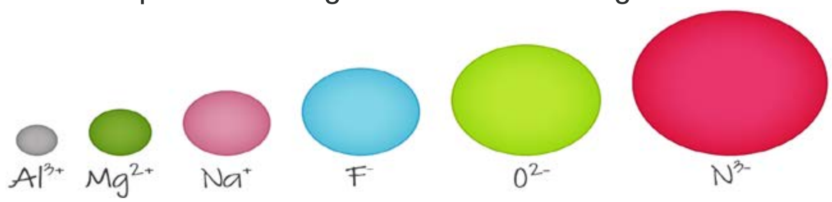


The **covalent radius** is measured as half the distance between two neighboring nuclei. The other two types of measurement are the **metallic radius** and the **van der Waals' radius**.

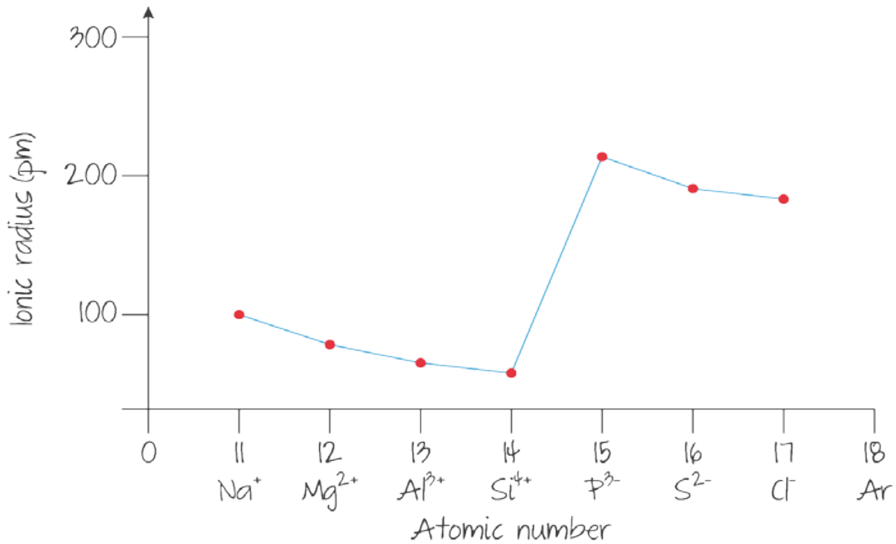
2. **Electron shielding** occurs when outer electrons are shielded from the attraction of the nucleus by inner electrons (known as shielding electrons). **Electron shielding** remains more or less constant across a period but nuclear attraction increases and atomic radius decreases across a period.

The overall trends in atomic radius in the periodic table

3. **Ionic radius**
   Across a period starting with smallest to largest



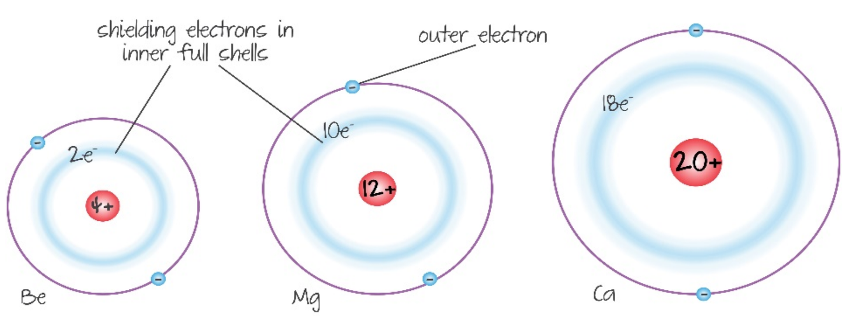
Isoelectronic species have the same number of electrons. Examples include cations such as Na^+, Mg^{2+} and Al^{3+} and anions such as N^{3-}, O^{2-} and F^-. These ions have different numbers of protons, as they are different elements, but the same number of electrons.



The ionic radii data for the ions formed by the elements in period 3. Note that there is a distinct change in the trend as we move from elements that form positive ions to those that form negative ions.

4. **Effective nuclear charge**

Because of shielding effect and change in atomic radius, valence electrons which are found in the outermost energy levels do not feel the full attraction from the protons in the nucleus. The attraction felt by the valence electrons, known as the **effective nuclear charge**, is less than the actual nuclear charge of the atom.

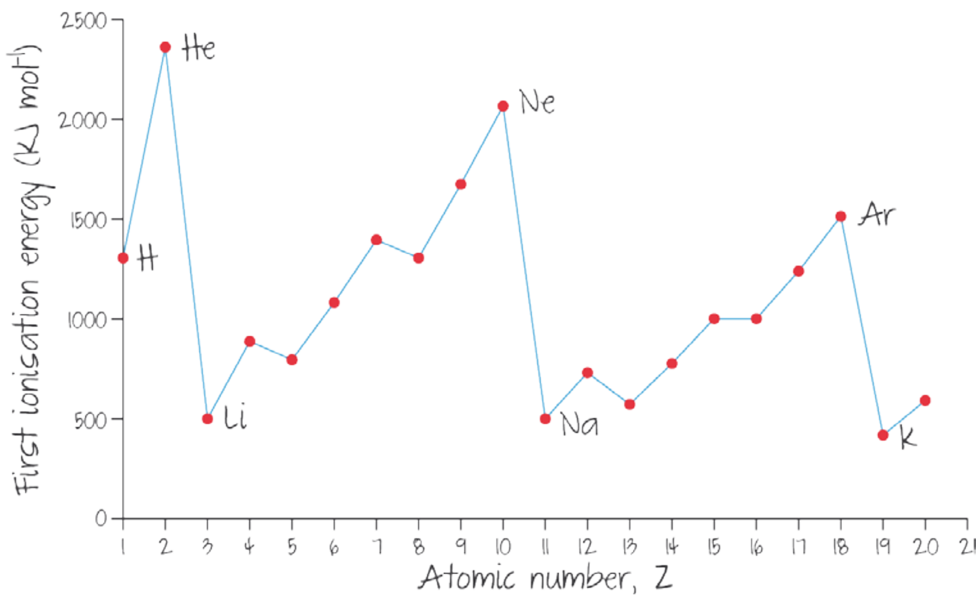


The effective nuclear charge remains approximately constant going down a group - in this case, group two.

5. **First ionisation energy**

The general equation for the first ionisation energy of an atom is shown below:

6. $X(g) \longrightarrow X^{+}(g) + e^{-}$ IE = + *x* kJ mol^{-}



Plot of the first ionisation energies of the first twenty elements

Exceptions to the trend

Beryllium has the electron configuration $1s^2$ $2s^2$

Boron has the electron configuration $1s^2$ $2s^2$ $2p^1$

Nitrogen has the electronic configuration $1s^2$ $2s^2$ $2p^3$

Oxygen has the electronic configuration $1s^2$ $2s^2$ $2p^4$

Exam Tip

When you are asked to explain the difference between the first ionisation energies of elements in a period, remember to mention:

- nuclear charge,
- energy level (or distance from the nucleus), and
- repulsion by other electrons (shielding).

7. **Electronegativity**

Electronegativity is defined as the attraction of an atom for a bonding pair of electrons. It is measured on the **Pauling scale**, a relative scale that assigns fluorine a value of 4.0 and francium a value of 0.7.

The trend in electronegativity in the periodic table

8. **Electron affinity**

The **first electron affinity** is the energy released when one mole of electrons are added to one mole of gaseous atoms to form one mole of gaseous 1^- ions. First electron affinities are **exothermic**, whereas **second electron affinities** tend to be **endothermic**, as in the case of oxygen. The equation for the first electron affinity of fluorine is shown below.

$F(g) + e^- \rightarrow F^-(g) \quad E_A = -328 \text{ kJ mol}^{-1}$

Definition

The terms exothermic and endothermic refer to the enthalpy change of a process. Exothermic processes give out heat and endothermic processes absorb heat.

The elements in group 17 have the highest electron affinities (most exothermic).

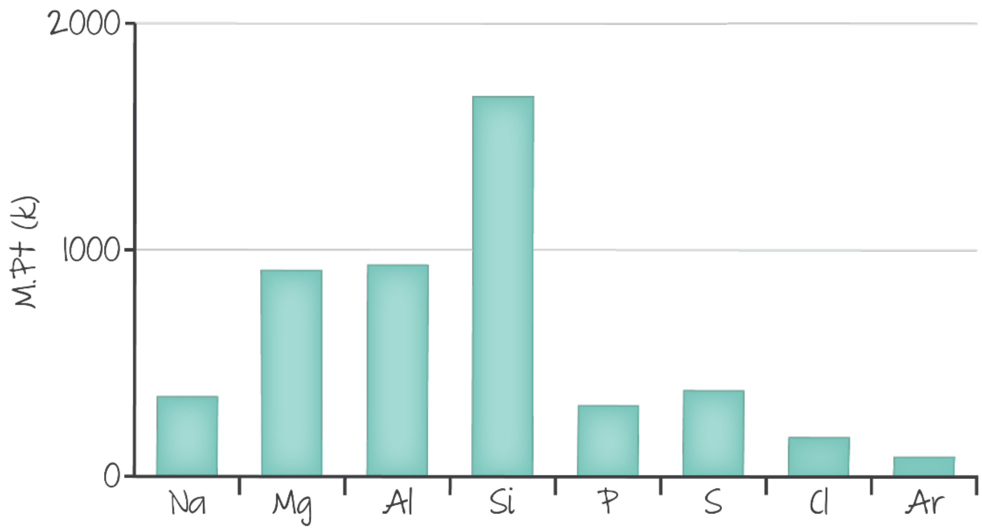
The general trend down a group is decreasing electron affinity. The additional electron gained is entering an energy level further from the nucleus.

9. **Trends in metallic character**

The metallic character of an element is strongly related to its ionisation energy, which depends on the nuclear charge and atomic radius of an atom. The higher the nuclear charge and smaller the atomic radius of an atom, the lower its metallic character, and vice versa.

10. **Trends in melting point**

The melting point of a substance depends on its structure and bonding. Across a period, the structure and bonding gradually change from **metallic** to **giant covalent** to **molecular covalent**.



The trend in melting point across period 3

11. **Trends in oxide behavior**

The **period 3 oxides** show a gradual trend of decreasing metallic character across the period. Metal oxides such as sodium oxide (Na_2O) and magnesium oxide (MgO) have giant ionic structures.

The acid–base properties of the period 3 oxides show a change from basic to acidic across the period. Aluminium oxide is **amphoteric** (it can act as either an acid or a base).

Basic oxides, such as those shown below, dissolve in water to produce alkaline solutions.

$$Na_2O\ (s) + H_2O\ (l) \rightarrow 2NaOH\ (aq)$$

$$MgO\ (s) + H_2O\ (l) \rightarrow Mg(OH)_2\ (aq)$$

Acidic oxides dissolve in water to produce acidic solutions according to the equations below.

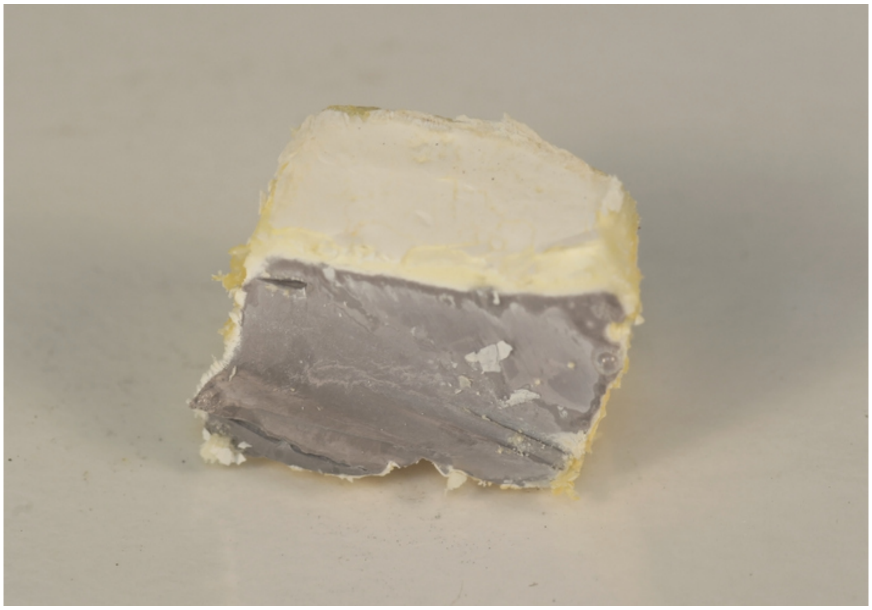
$$P_4O_{10}(s) + 6H_2O\ (l) \rightarrow 4H_3PO_4(aq)$$

$$SO_3(g) + H_2O\ (l) \rightarrow H_2SO_4(aq)$$

$$3NO_2(g) + H_2O\ (l) \rightarrow 2HNO_3(aq) + NO\ (g)$$

12. **The alkali metals**

    The alkali metals are located in group 1 of the periodic table, which is made up of the elements lithium, sodium, potassium, rubidium, caesium and francium.

    When cut, sodium quickly reacts with the oxygen in the air to form a layer of sodium oxide.

    The alkali metals are stored in oil to prevent them reacting with the oxygen in the air.

The reactions of sodium and potassium with water are shown below:

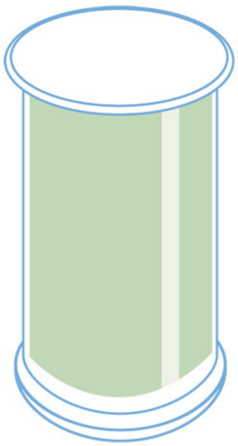
$$2Na\ (s) + 2H_2O\ (l) \rightarrow 2NaOH\ (aq) + H_2(g)$$

$$2K\ (s) + 2H_2O\ (l) \rightarrow 2KOH\ (aq) + H_2(g)$$

13. **The halogens**

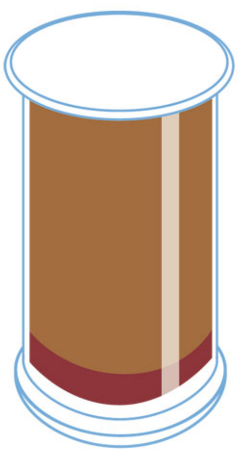
The elements in group 17, known as the halogens, are a very reactive group of non-metal elements.

The halogens are **diatomic,** which means they exist as two atoms bonded together, rather than separate atoms.



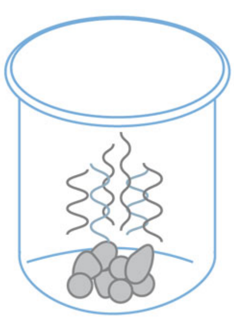
Chlorine (Cl_2)

- dense pale-green gas
- smelly and poisonous
- occurs as chloride in the sea
- relative atomic mass 35.5



Bromine (Br_2)

- deep-red liquid with red-brown vapour
- smelly and poisonous
- occurs as bromide in the sea
- relative atomic mass 80



Iodine (I_2)

- grey solid with purple vapour
- smelly and poisonous
- occurs as iodides and iodates in some rocks and in seaweed
- relative atomic mass 127

The halogens can also undergo **displacement reactions** with each other. Chlorine, for example, will displace those halogens beneath it from a solution of their salts. Chlorine will displace iodine from a colourless solution of potassium iodide, producing a brown-coloured solution according to the following equation:

$$Cl_2 (aq) + 2KI\ (aq) \rightarrow 2KCl\ (aq) + I_2 (aq)$$