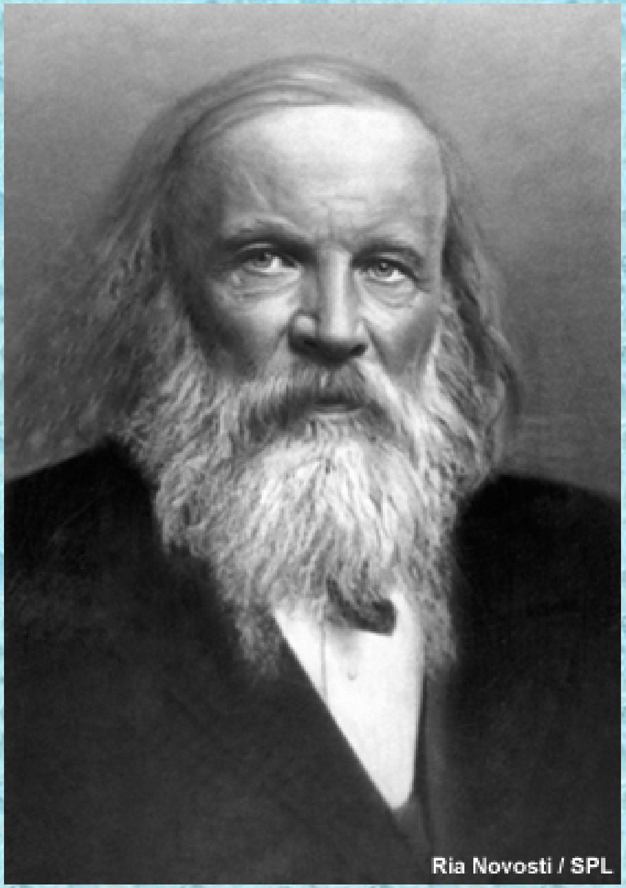
What is periodicity?

The term **periodicity** describes a repeating pattern in properties of elements across periods of the periodic table.

The Russian chemist Dmitry Mendeleev is credited with being the creator of the first version of the periodic table. He observed that when the elements are arranged in order of atomic mass, there are recurring patterns in certain properties.



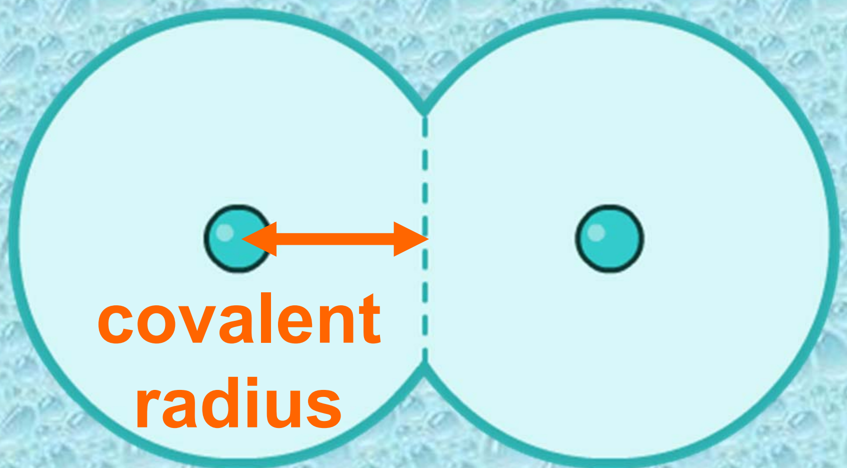
The modern periodic table can be used to analyse trends in properties such as **atomic radius** across periods and down groups.

What is atomic radius?

The **atomic radius** of an element is difficult to precisely define because of the uncertainty over the size of the electron cloud. Several definitions are used.

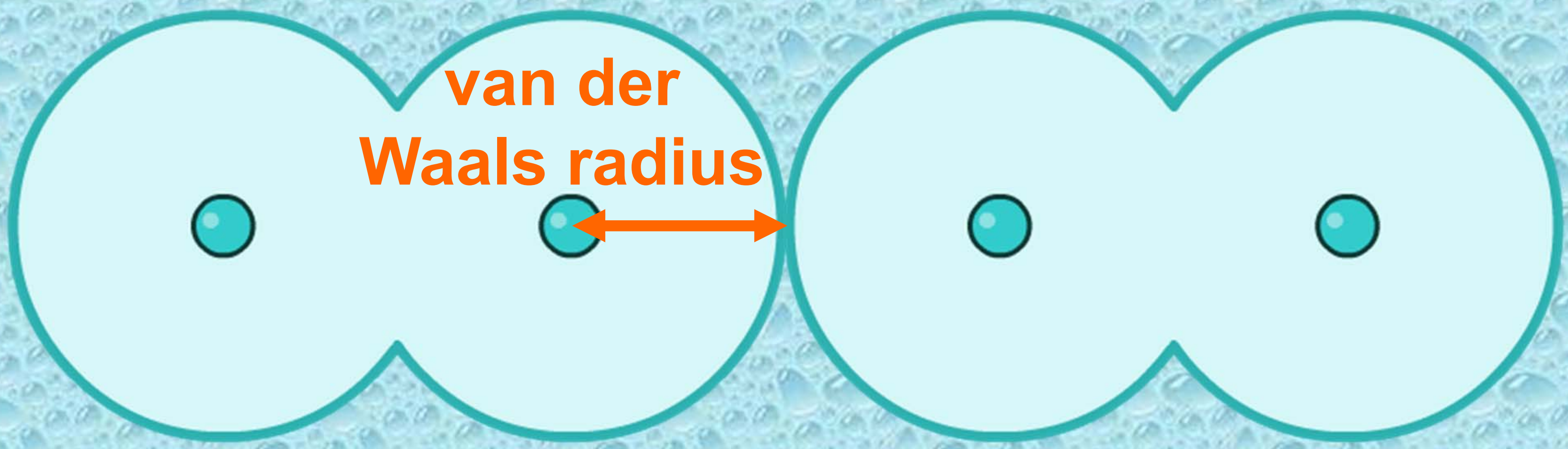
One definition is half the shortest internuclear distance found in the structure of the element.

For non-metallic elements, the **covalent radius** is often used as the atomic radius. This is half the internuclear distance between two identical atoms in a single covalent bond.



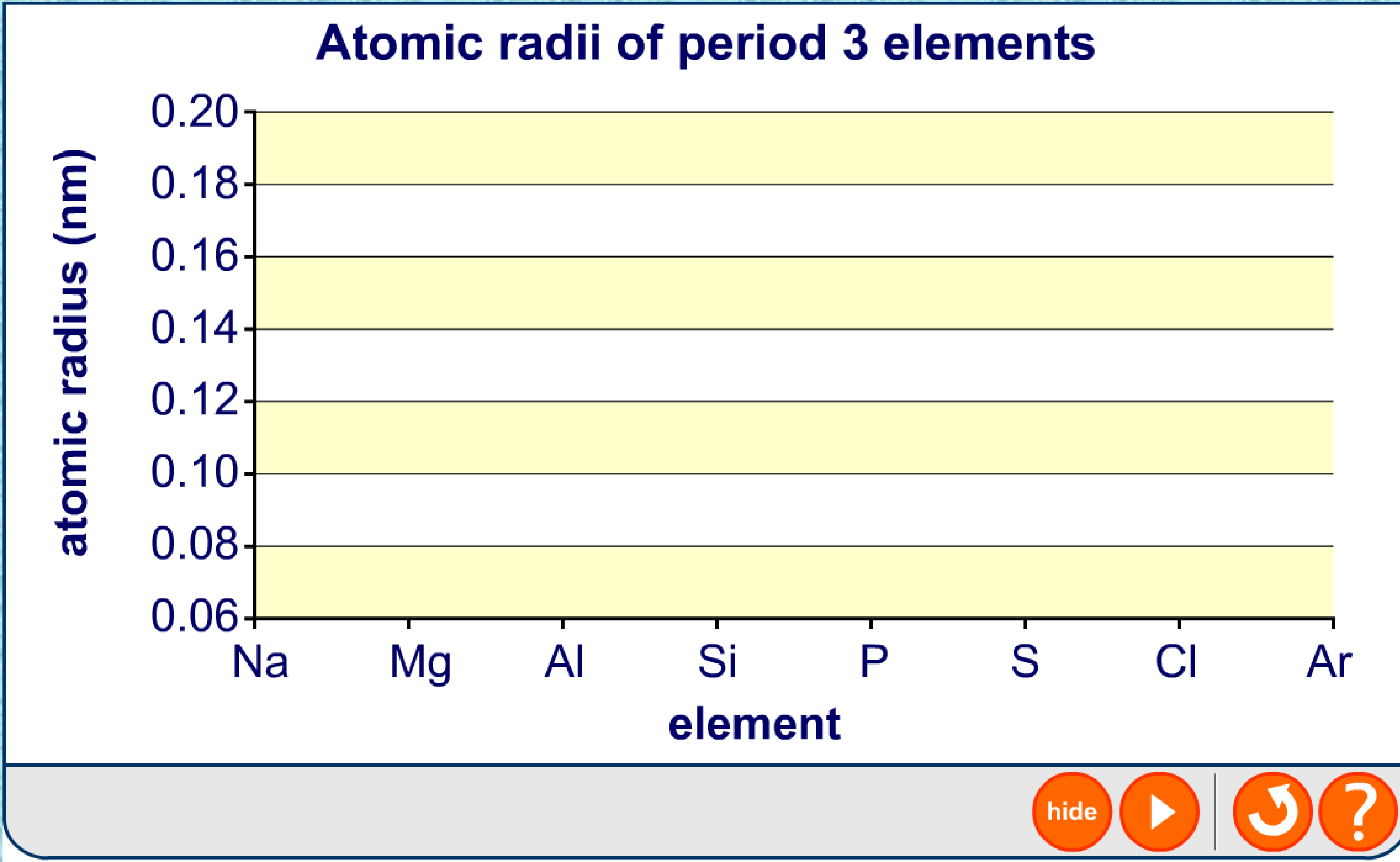
More on atomic radius

For non-bonded adjacent atoms (e.g. in a covalent crystal of a non-metallic element), the **van der Waals radius** is used as a value for atomic radius. This is half the shortest internuclear distance between two similar non-bonded atoms.



For metallic elements, the **metallic radius** is often used as the atomic radius. This is half the shortest internuclear distance between two adjacent atoms in a metallic bond.

Trends in atomic radius in period 3



Trends in atomic radius in period 3

Element	Atomic radius (nm)
Na	0.190
Mg	0.145
Al	0.118
Si	0.111
P	0.098
S	0.088
Cl	0.079
Ar	0.071

The atomic radius of the elements across period 3 decreases.

This might seem counter-intuitive, because as the numbers of sub-atomic particles increase, the radius might be expected to also increase.

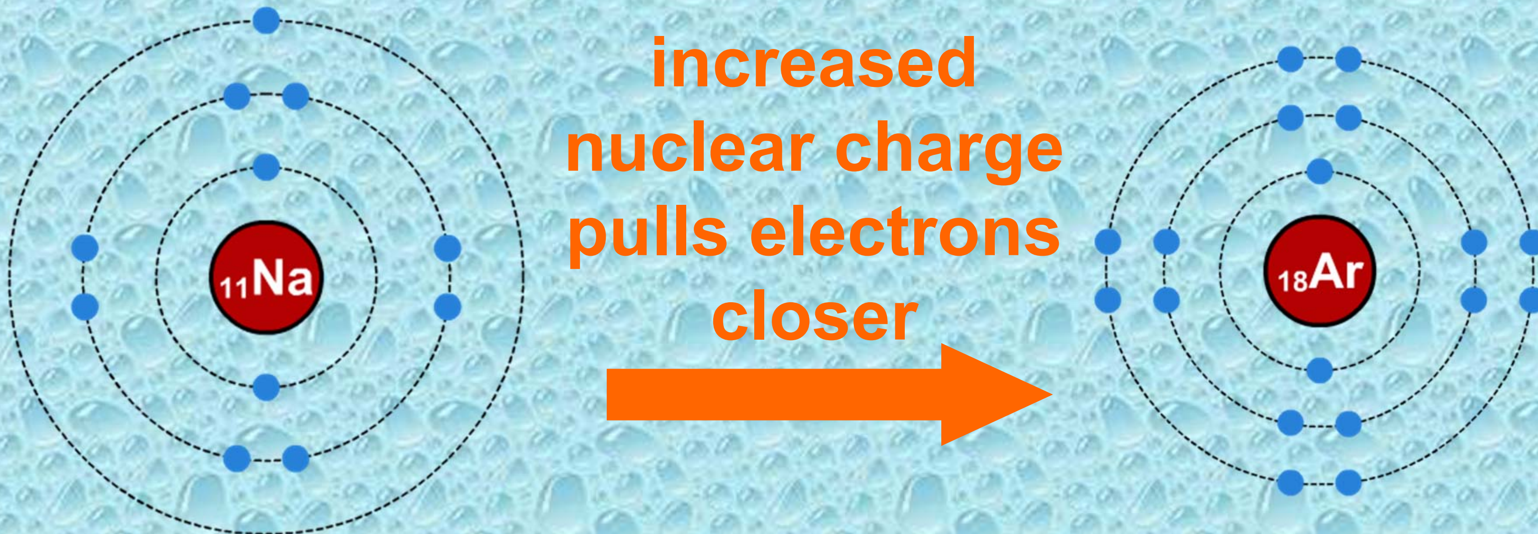
However, more than 99% of the atom is empty space – the nucleus and electrons themselves occupy a tiny volume of the atom.

Increase in proton number

The number of protons in the nucleus of the atoms increases across period 3.

proton number **Element**	$_{11}Na$	$_{12}Mg$	$_{13}Al$	$_{14}Si$	$_{15}P$	$_{16}S$	$_{17}Cl$	$_{18}Ar$

This increase in the number of protons increases the **nuclear charge** of the atoms. The nucleus has stronger attraction for the electrons, pulling them in closer and so the atomic radius decreases across the period.

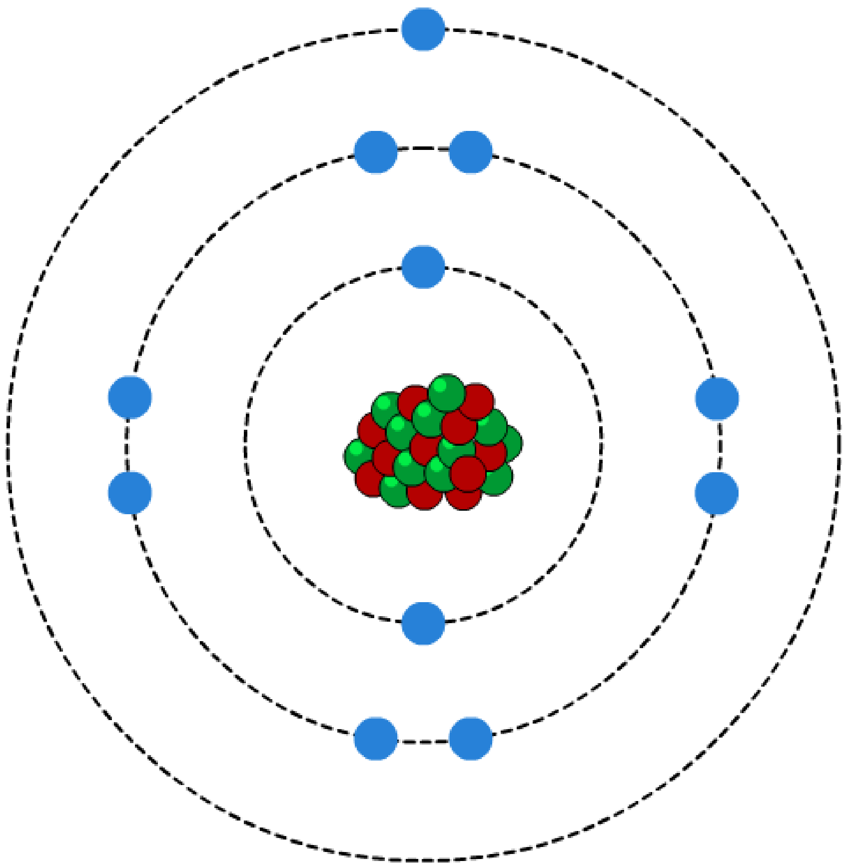


What is shielding?

What is the shielding effect?

The attraction between the outermost electrons of an atom and its nucleus is not normally as large as might be expected.

Click "**play**" or the atom to find out why.



Explaining atomic radius in period 3

Element	Proton number	Atomic radius (nm)
Na	11	0.190
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Proton number increases across period 3, but shielding remains approximately constant.

This causes an increase in **effective nuclear charge**, leading to a greater attraction between the nucleus and the outermost electrons.

This pulls these electrons closer to the nucleus and results in a smaller radius.

Atomic radius in period 3

Arrange the period 3 elements in order of increasing radius

1. chlorine
2. magnesium
3. phosphorus
4. silicon
5. aluminium
6. argon
7. sodium
8. sulfur

Atomic radius: true or false?

Are these statements about atomic radius true or false?

1.	The covalent radius is the distance between two identical atoms joined by a covalent bond.	?
2.	The atomic radius of period 3 elements decreases from sodium to argon.	?
3.	Increasing the number of protons, neutrons and electrons always increases the size of the atom.	?
4.	Effective nuclear charge is the resultant attractive force from the nucleus experienced by the outer electrons.	?
5.	Effective nuclear charge decreases from sodium to argon due to increased shielding by extra electrons.	?

true false

S ?

What is first ionization energy?

Ionization is a process in which atoms lose or gain electrons and become ions.

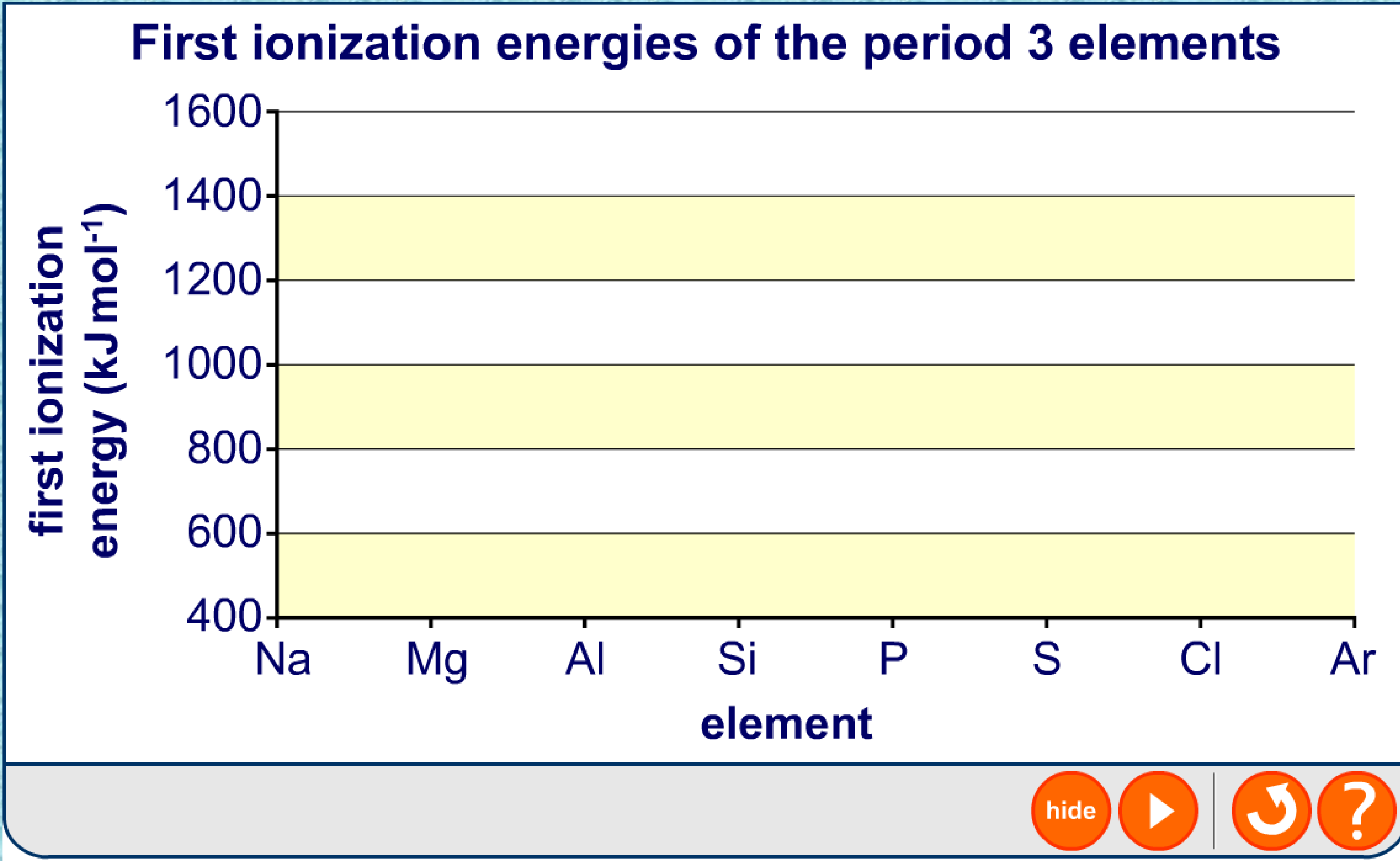
The first ionization energy of an element is the energy required to *remove* one electron from a gaseous atom.

$$M(g) \rightarrow M^+(g) + e^-$$

The first ionization energy is therefore a measure of the strength of the attraction between the outermost electrons and the nucleus.

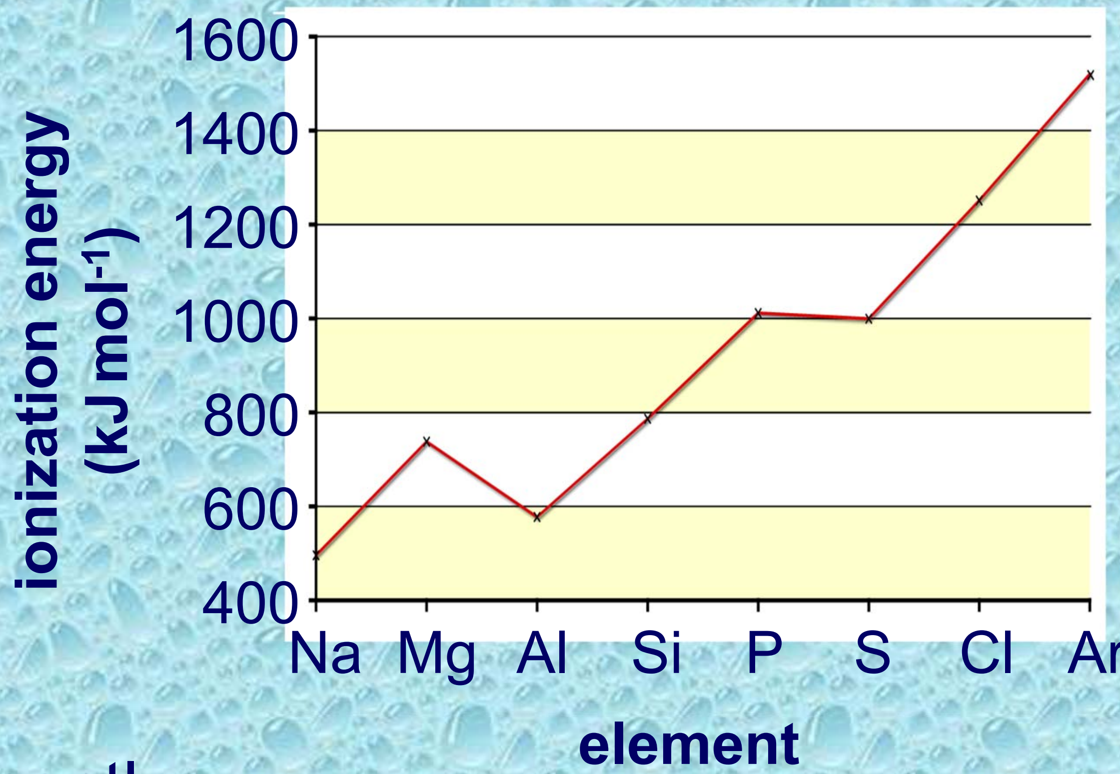
The first ionization energies of the elements in periods 2 or 3 can give information about their electronic structure.

Plot of the first ionization energies



General trend in first ionization energy

There is a **general increase** in the first ionization energies across period 3.



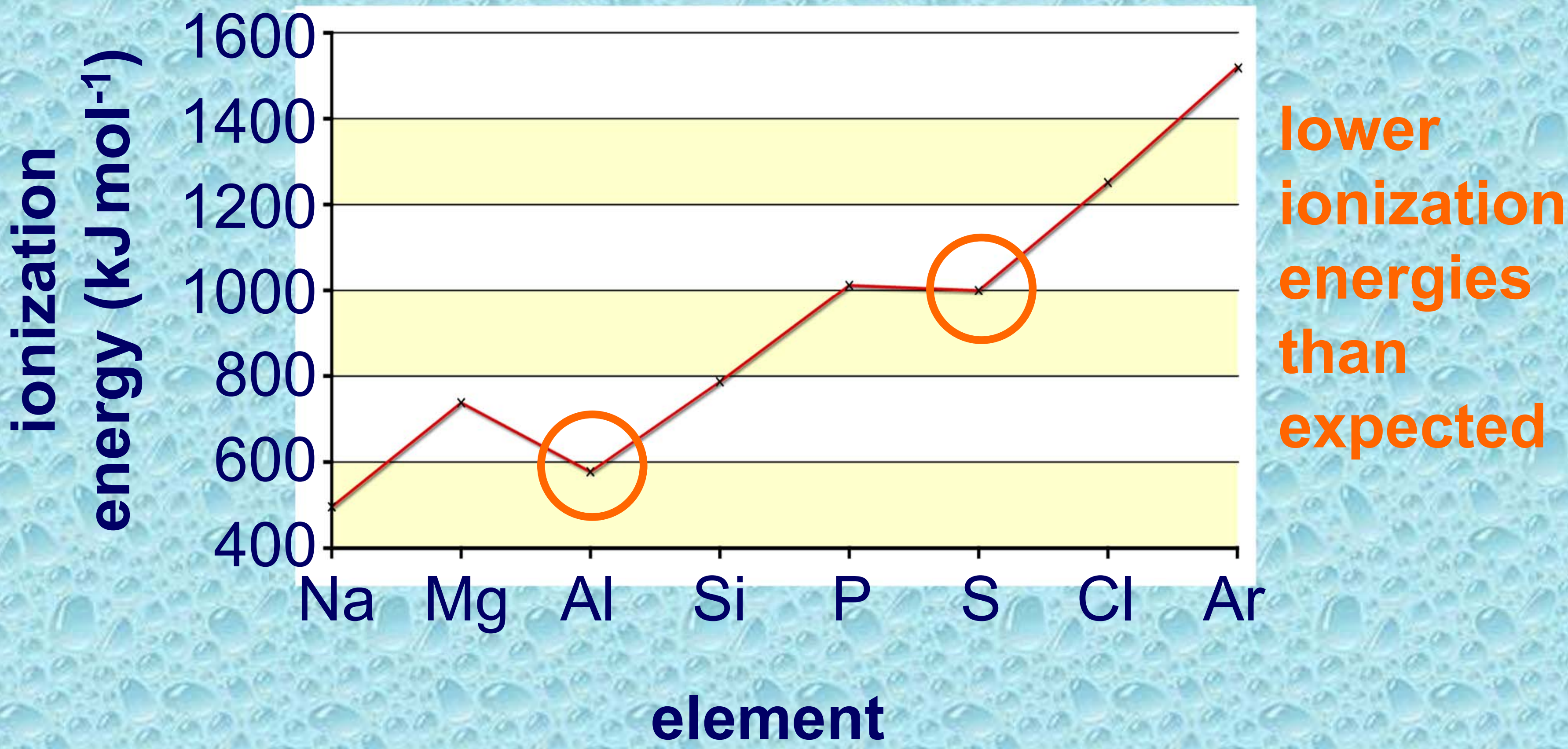
Across period 3, the proton number increases but the amount of shielding does not change significantly.

The effective nuclear charge therefore **increases**.

The greater attraction between the nucleus and the outermost electrons means that more energy is required to remove an electron.

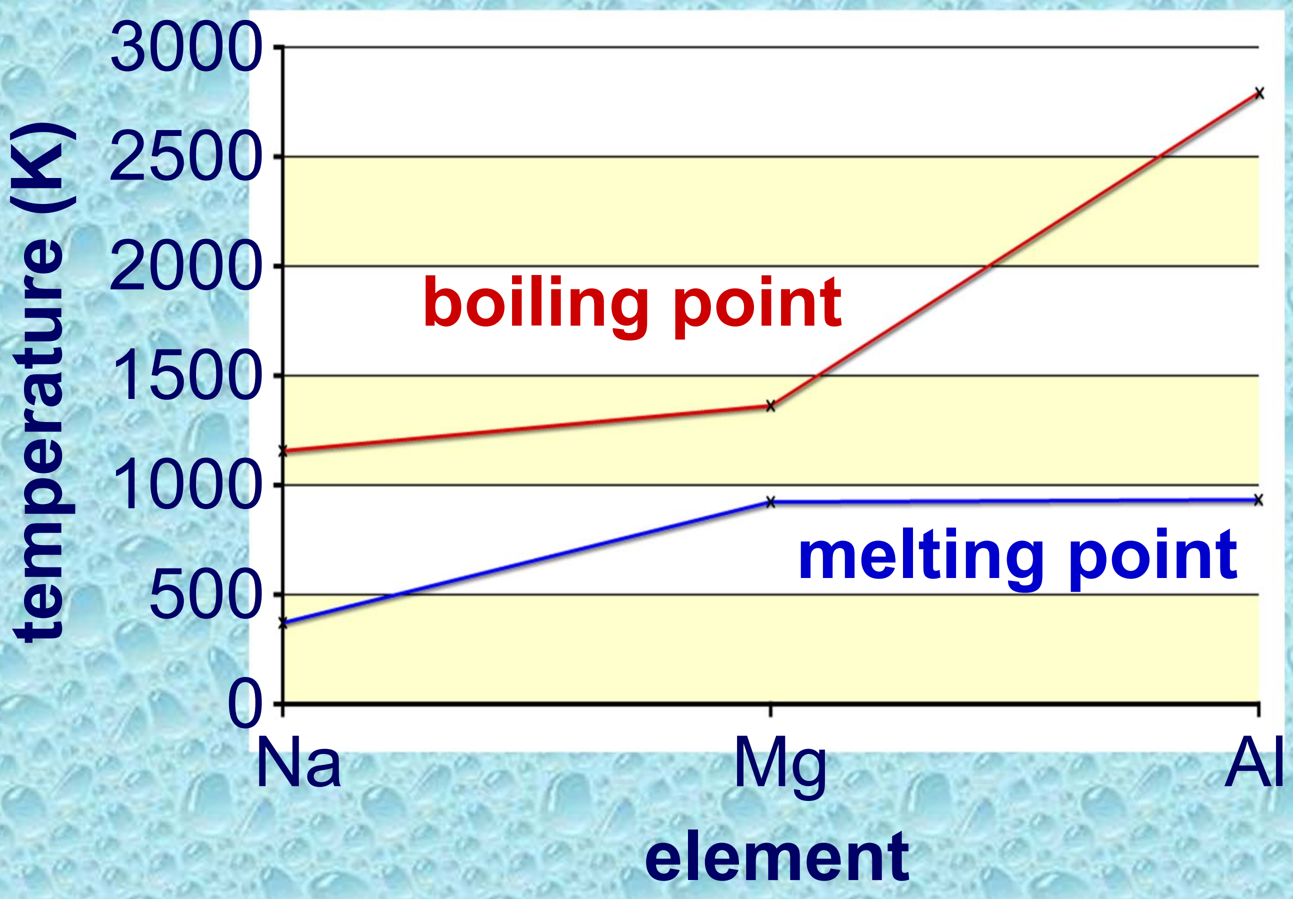
Trend in first ionization energy: exceptions

There are two exceptions to the general trend in first ionization energy: both aluminium and sulfur have lower ionization energies than might be expected.



Na, Mg and Al: melting and boiling points

The melting and boiling points increase for the three metallic elements from sodium to aluminium.

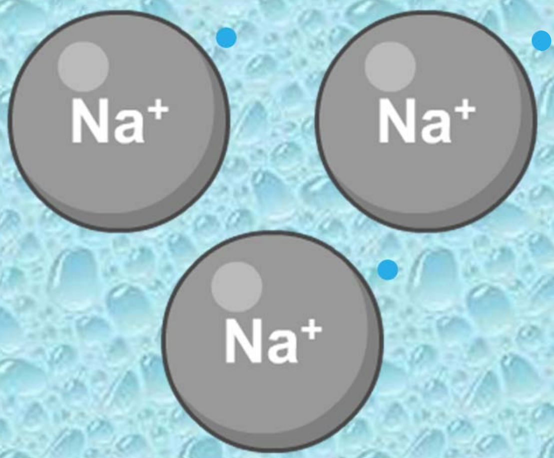


This is because the strength of the metallic bonds increases. More energy is needed to break the stronger metallic bonds, so melting and boiling points are higher.

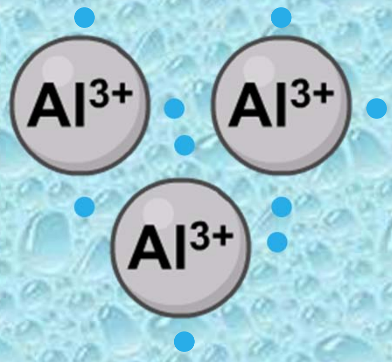
Na, Mg and Al: metallic bond strength

The increase in metallic bond strength from sodium to aluminium is due to two factors:

1. **Charge density.** This is the ratio of an ion's charge to its size. Na^+ ions are large with a small charge, so have a low charge density. Al^{3+} ions are smaller with a larger charge, and so have a higher charge density. They are therefore more strongly attracted to the delocalized electrons.



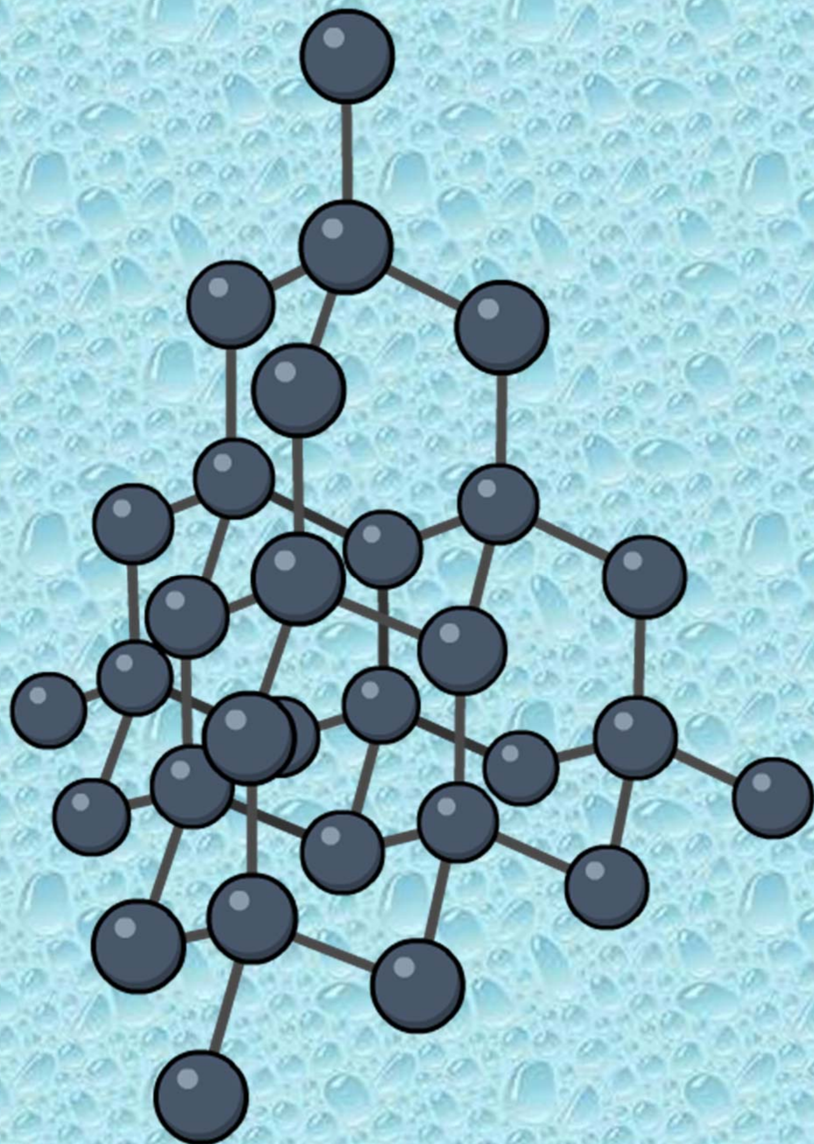
2. **Number of free electrons.** Sodium has one free electron per metal ion, whereas aluminium has three. This leads to more attractions that must be broken in aluminium.



Silicon

Silicon has a **macromolecular structure** similar to that of diamond.

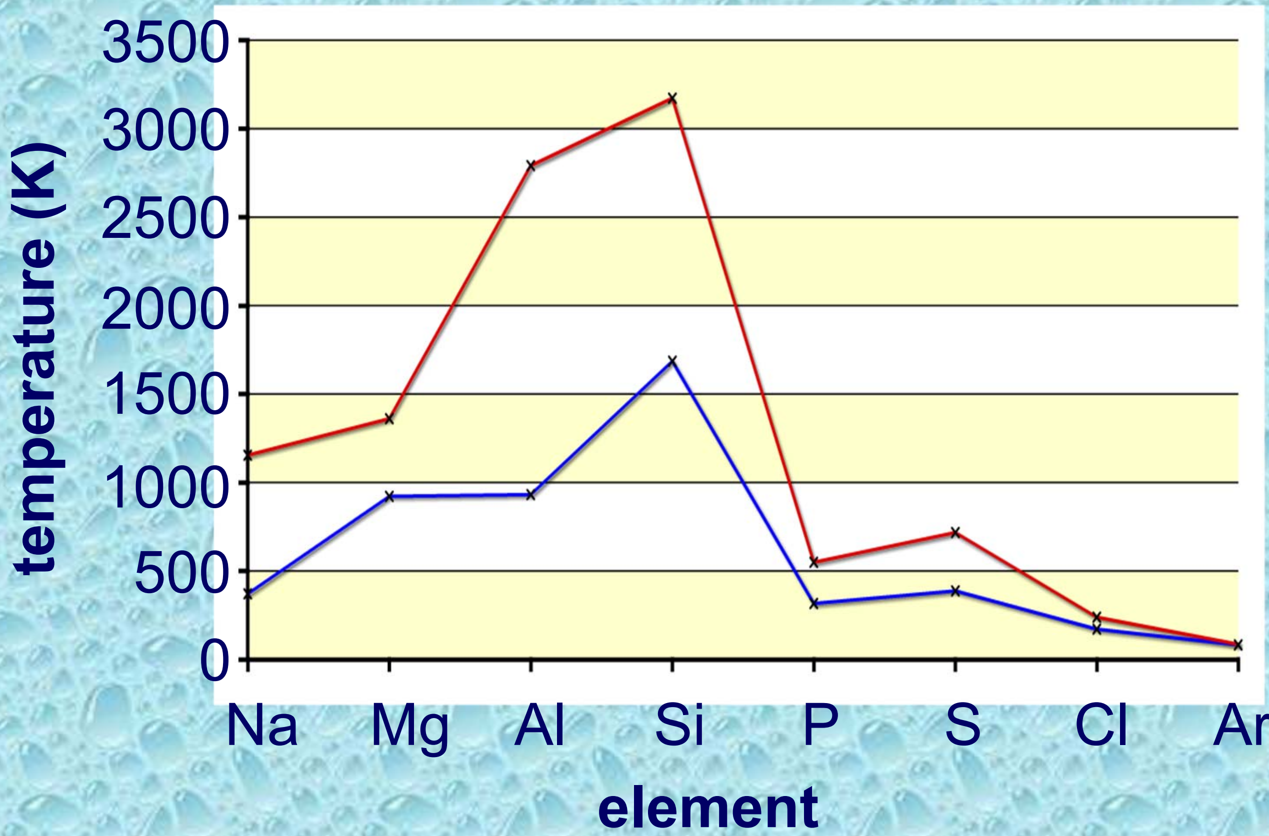
Each silicon atom is bonded to four neighbouring silicon atoms by strong covalent bonds. These must be broken in order for silicon to melt. This requires a lot of energy, so silicon's melting and boiling points are high.



Period 3 non-metals

The melting and boiling points of phosphorus, sulfur and chlorine are much lower than those of silicon.

This is because they have a **simple molecular structure** with weak van der Waals forces holding the molecules together.



Breaking these forces of attraction requires much less energy than breaking covalent bonds.

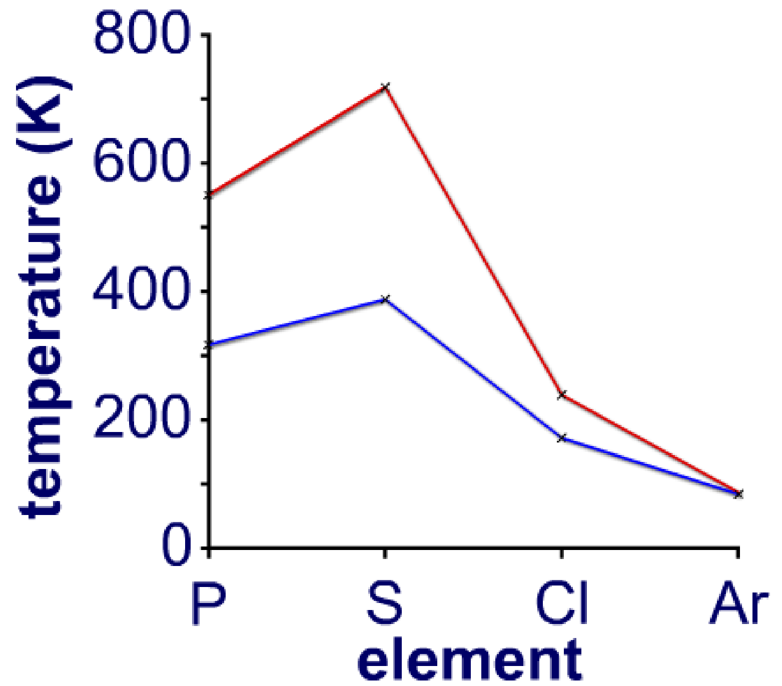
Period 3 non-metals: structure

Structure and melting points of period 3 non-metals

Phosphorus, sulfur and chlorine all have a molecular structure, but their melting and boiling points are different.

This is because the strength of the van der Waals forces is dependent on the size of the molecules and the number of electrons.

Click an element to find out more.



31.0 P 15	32.1 S 16	35.5 Cl 17	39.9 Ar 18

Melting points in period 3

What are the missing words about melting points?

1. Boiling points in period 3 follow **A** trends to melting points in period 3.

2. The strength of **B** bonding **C** from sodium to aluminium. Melting and boiling points therefore **D** because **E** energy is needed to break the bonds. The bonds are stronger when charge density is **F** and number of delocalized electrons is **G**.

C S ?