



Chapter 3: Periodicity – fast facts

3.1 The Periodic Table

The arrangement of elements in the Periodic Table helps to predict their electron configuration.

- The Periodic Table arranges elements according to increasing atomic number / increasing number of protons.
- The horizontal rows are called **periods** and vertical columns are called **groups**.
- The period number (n) is the outer energy level that is occupied by electrons.
- Elements in the same period have outer electrons in the same energy level.
- The groups are numbered from 1 to 18.
- Elements in the same group have the same number of outer electrons and have similar chemical properties.
- The Periodic Table is arranged in 4 blocks – s, p, d and f – which are associated with the highest sub-level occupied by electrons.
- The position of an element is related to the electron configuration of its atoms. Magnesium, for example, is in Period 3, as it has three occupied energy levels, and in Group 2, as there are two electrons in its outer energy level. It is in s block as it has outer electrons in the s sub-level.
- Metals are found on the left of the Periodic Table and non-metals on the right. Metalloids form a diagonal staircase between the metals and non-metals.
- Group 1 = alkali metals, Group 17 = halogens, Group 18 = noble gases.
- The **transition metals** are in the large section of **d-block** elements in the middle of the Periodic Table from Sc to Zn, etc. Zn is not a transition metal because it does not form ions with incomplete d sub-levels.
- Lanthanoids and actinoids are metals in the first and second row of the f block.

3.2 Periodic trends

Elements show trends in their physical and chemical properties across periods and down groups.

- The chemical and physical properties of elements arranged in order of increasing atomic numbers vary periodically. **Periodicity** is the regular repetition of properties of elements arising from patterns in their electron arrangement.
- Effective nuclear charge refers to the nuclear pull experienced by the outer electrons in an atom, taking into account the shielding effect of inner full shells of electrons.
- **Atomic radius:** atomic radii decrease along a period as the nuclear charge increases and electrons are added to the same outer shell. The attraction between the outer electrons and nucleus increases.
- As a group is descended and the number of occupied energy levels increases, the atomic radii increase.
- **Ionic radius:** Cations are smaller than their parent atoms, as the formation of positive ions involves the loss of the outer shell.
 - Anions are larger than their parent atoms, as their formation involves the addition of electrons into the outer shell. The increased electron repulsion between the electrons in the outer energy level increases the radius of the outer shell.
 - Cation < atom < anion



- Patterns in ionic radii across a period are more complex.
 - The ionic radii decrease from Groups 1 to 14 for the positive ions. The ions Na^+ , Mg^{2+} , Al^{3+} and Si^{4+} are **isoelectronic** and have the same electron configuration ($1s^2 2s^2 2p^6$). The decrease in ionic radius is due to the increase in nuclear charge with atomic number across the period, which increases the attraction between the nucleus and the outer electrons.
 - The ionic radii decrease from Groups 14 to 17 for the negative ions. The ions Si^{4-} , P^{3-} , S^{2-} and Cl^- are isoelectronic and have the same electron arrangement ($1s^2 2s^2 2p^6 3s^2 3p^6$). The decrease in ionic radius is due to the increase in nuclear charge across the period.
 - The positive ions are smaller than the negative ions in the same period, as the former have one less occupied electron shell.
- The **first ionization energy (IE)** of an element is the minimum energy required to form one mole of singly charged positive ions (M^+) by removing an electron from each atom (M) in the gaseous state: $\text{M}(\text{g}) \rightarrow \text{M}^+(\text{g}) + \text{e}^-$ (units: kJ mol^{-1}).
- First ionization energies decrease down a group. This is due to the increased distance between the nucleus and the outer energy level that reduces the force of attraction between the nucleus and the outer electrons.
- Ionization energies increase along a period due to the increase in effective nuclear charge, which causes the outer electrons to be held more tightly.
- There are regular discontinuities in the trend of increasing ionization energies along a period, due to the existence of sub-levels within the main energy levels.
- **Electron affinity** is the energy change when one mole of electrons is added to one mole of gaseous atoms to form one mole of gaseous ions: $\text{X}(\text{g}) + \text{e}^- \rightarrow \text{X}^-(\text{g})$
- First electron affinities are usually endothermic, and increase along a period due to increasing nuclear charge.
- **Electronegativity** is a measure of the attraction of a nucleus for bonding electrons.
- Electronegativity increases along a period and decreases down a group.
- The noble gases are not assigned electronegativities as they do not readily form bonds with other elements.
- The electronegativities of diagonal elements remain approximately the same as both the group and period number increase. Boron and aluminium, for example, both have electronegativities of 1.6.
- The electronegativity of H is the same as that of P.
- Metals have low ionization energies and electronegativities; non-metals have high ionization energies and electronegativities.
- The **melting points** decrease down Group 1 as there is a decrease in the strength of the **metallic bonding**. The increase in ionic radii reduces the force of the attraction between the M^+ ions and the **delocalized electrons**.
- The melting points increase down Group 17, as there is an increase in the strength of London (dispersion) forces with increasing number of electrons.
- The chemical properties of elements are generally due to the number of electrons in the outer energy level of their atoms.

Group 1

- All the metals are too reactive to be found native in nature.
- They generally donate electrons and act as **reducing agents**: $\text{M} \rightarrow \text{M}^+ + \text{e}^-$
- Reactivity increases down the group with the decrease in ionization energies.
- Their ability to conduct electricity is also due to the mobility of their outer electron.



- The alkali metals react with water to produce hydrogen and the metal hydroxide. The resulting solution is alkaline owing to the presence of hydroxide ions: $2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$

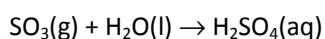
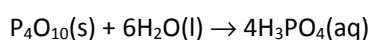
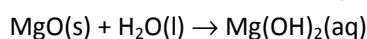
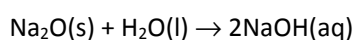
Group 17

- The halogens are diatomic non-polar molecules.
- They generally accept electrons and act as **oxidizing agents**: $X_2 + 2e^- \rightarrow 2X^-$
- Reactivity decreases down the group, as the atom gets larger and attraction for extra electrons decreases. The more reactive halogens, X_2 , **displace** the less reactive halogens, Y, from their compounds:
 $X_2 + 2Y^- \rightarrow Y_2 + 2X^-$
 e.g. the more reactive Cl displaces Br: $Cl_2 + 2Br^- \rightarrow Br_2 + 2Cl^-$
- The halogens react with the Group 1 metals to form ionic **halides**: $2M + X_2 \rightarrow 2MX$
 The most vigorous reaction occurs between the elements which are furthest apart in the Periodic Table.

Period 3 oxides

- Oxides change from basic through amphoteric to acidic across a period.

Formula and state at room temperature	Na ₂ O(s)	MgO(s)	Al ₂ O ₃ (s)	SiO ₂ (s)	P ₄ O ₁₀ (s)/ P ₄ O ₆ (s)	SO ₃ (l)/ SO ₂ (g)	Cl ₂ O ₇ (l)/ Cl ₂ O(g)
Structure	Ionic			Giant covalent	Molecular covalent		
Acid/base character	Basic		Amphoteric	Acidic			



As SiO₂(s) is insoluble it does not change the pH if added to H₂O.



13.1 First-row d-block elements

The transition elements have characteristic properties; these properties are related to their all having incomplete d sub-levels.

- Transition metals have partially filled d orbitals in their atoms or ions.
- Zn is not a transition element because it has a full d sub-level in its atoms and ions.
- Characteristic properties include: variable oxidation number, complex ion formation with ligands, existence of coloured compounds, have catalytic and magnetic properties.
- Multiple oxidation states arise because the 3d and 4s sub-levels are close in energy and both are involved in bonding.
- All d-block elements except Sc show an oxidation state of +2.
- All d-block elements except Zn show an oxidation state of +3.
- A ligand is a molecule or negative ion that donates a pair of electrons to a central metal ion to form a covalent (coordinate) bond. They are Lewis bases.
- Complex **ions** are formed when a central metal ion is bonded to a ligand with a coordinate bond. Examples include $[\text{Fe}(\text{H}_2\text{O})_6]^{3+}$, $[\text{Fe}(\text{CN})_6]^{3-}$, $[\text{CuCl}_4]^{2-}$, and $[\text{Ag}(\text{NH}_3)_2]^+$.
- The charge on a complex ion is the sum of the charges of the metal ion and the ligands.
- Transition metals act as heterogeneous catalysts as they can provide a surface for reaction: they use the 3d and 4s electrons to form weak bonds to reactant molecules.
- Magnetic properties are a result of unpaired electrons in the transition metal atom or ion.

13.2 Coloured complexes

d orbitals have the same energy in an isolated atom, but split into two sub-levels in a complex ion. The electric field of ligands cause the d orbitals in complex ions to split so that the energy of an electron transition between them corresponds to a photon of visible light.

Transition metal ions are coloured due to **d–d electron transitions** between d orbitals which are split in the electric field due to the presence of the ligands.

- The colour observed is complementary to the colour absorbed, and can be deduced from the colour wheel.
- The colour of a complex depends on the identity of the metal ion, the oxidation state of the metal, and the identity of the ligand.
- Ions with higher charge and ligands with greater charge density produce a greater split in the d orbitals.
- The spectrochemical series arranges the ligands according to the energy separation between the two sets of d orbitals.
- Polydentate ligands form more than one coordinate bond with the metal ion.