# TOPIC 3: PERIODICITY

**3. 1. THE PERIODIC TABLE**

E.I.: The arrangement of elements in the periodic table helps to predict their electron configuration.

**Nature of science** 1.9: Obtain evidence for scientific theories by making and testing predictions based on them – scientists organize subjects based on structure and function; the periodic table is a key example of this. Early models of the periodic table from Mendeleev, and later Moseley, allowed for the prediction of properties of elements that had not yet been discovered.

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| Understandings* 3.1 U1 The periodic table is arranged into four blocks associated with the four sub- levels—s, p, d, and f.
* 3.1 U2 The periodic table consists of groups (vertical columns) and periods (horizontal rows).
* 3.1.U3 The period number (n) is the outer energy level that is occupied by electrons.
* 3.1 U4 The number of the principal energy level and the number of the valence electrons in an atom can be deduced from its position on the periodic table.
* 3.1 U5 The periodic table shows the positions of metals, non-metals and metalloids.
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| Applications and skills3.1. AS1 Deduction of the electron configuration of an atom from the element’s position on the periodic table, and vice versa. |

1. Modern periodic table: elements are in order of increasing atomic number.
2. Periodic table is arranged into four blocks associated with the four sub-levels—s, p, d, and f.
3. **group** = vertical column of elements; within a group there are distinct chemical similarities but also gradual changes in properties..
4. **period** = horizontal row of elements from an alkali metal to noble gas.
5. From its position in the periodic table, the following information about an atom can be deduced:
6. The period number = the number of the principal energy level that is being filled (or equally the number of principal energy levels that have electrons in them)
7. Valence electrons = the electrons that are available for bonding;
8. in the *s* block these are the electrons in the *s* sublevel which is the group number of the group the element is in.
9. in the *p* block they are the electrons in the highest *s* and *p* sublevels; in the case of bromine that is 7 (this is = to the group number – 10)
10. in the *d* block they are the *s* and *d* sublevel electrons (this is the same as the group number)
11. The block an atom is in indicates the sublevel that is being filled, for example, sodium is on the *s*-block because its 3s sublevel is being filled.
12. The group an atom is in a block indicates the number of electrons there are in that sublevel, e.g:
13. atoms in the 3rd group in block *p* have 3 electrons in the highest *p* sublevel; 3 valence electrons
14. atoms in the 3rd group in block *d* have 3 valence electrons in the highest *d* sublevel.
15. The number of valence electrons and the principal energy level determine the chemical and physical properties of an atom.
16. Periodic table shows the positions of metals, non-metals and metalloids.
17. You need to know the names of some groups and blocks in the periodic table:
18. Group 1= alkali metals
19. Group 17 = halogens
20. Transition metals are across groups 3 to 12.
21. Lanthanoids are the 15 metallic elements from atomic number 57 to 71 in the first row of the *f* block (only 14 elements). However, lanthanum is in the d-block as its highest occupied sub-level is a d sub-level (5d).
22. Actinoids or actinides are the 15 metallic elements from atomic number 89 to 103 in the second row of the *f* block – only 14 elements. Also actinium is a d-block element as its 6d sub-level is being filled.

**3. 2. PHYSICAL PROPERTIES**

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

**Nature of science** 3.1 Looking for patterns - the position of an element in the periodic table allows scientists to make accurate predictions of its physical and chemical properties. This gives scientists the ability to synthesize new substances based on the expected reactivity of elements.

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| Understandings* 3.2 U1 Vertical and horizontal trends in the periodic table exist for atomic radius, ionic radius, ionization energy, electron affinity and electronegativity.
* 3.2.U2 Trends in metallic and non-metallic behaviour are due to the trends above.
* 3.2.U3 Oxides change from basic through amphoteric to acidic across a period.
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| Applications and skills* 3.2.AS1 Prediction and explanation of the metallic and non-metallic behaviour of an element based on its position in the periodic table.
* 3.2.AS2 Discussion of the similarities and differences in the properties of elements in the same group, with reference to alkali metals (group 1) and halogens (group 17).
* 3.2.AS3 Construction of equations to explain the pH changes for reactions of Na2O, MgO, P4O10, and the oxides of nitrogen and sulfur with water.
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**Atomic radius**

How do we measure the atomic radius?

1. in non-metallic elements: the atomic radius is half the distance between the centres of two covalently bonded atoms (= covalent radius).
2. in metallic elements: the atomic radius is half the internuclear distance in the metallic crystal.

These radii can only be measured as distances between two nucleii as only the positions of the nucleii can be established accurately enough. The atomic radius is limited to the point at which the nuclear attraction is cancelled out by the electron repulsion. In a simplified way, the atomic radius is similar to the distance between the nucleus and the most outer electron.

***Horizontal trend***

Atomic size/radius decreases as you move across periods because:

* the nuclear charge increases;
* electrons are added to same main energy level /similar shielding effect as the number of complete inner energy levels is the same.
* As a result of the above two factors the effective nuclear charge increases and the outer electrons are attracted more strongly. The effective nuclear charge is the atomic number – number of electrons on the filled inner principal energy levels.

The increased repulsion between the electrons which would cause the radius to increase is cancelled out by the increased nuclear charge.

***Vertical trend***

Atomic radius increases as you go down groups because:

* the number of main energy levels increases (more electrons) which cancels out the increase in nuclear charge; outermost electron is placed in a higher energy level which is further away from the nucleus.
* the shielding effect increases as you go down as there are more filled energy levels.

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| **shielding effect**The screening or shielding effect from the nuclear charge by the inner electrons; as a result the valence electrons do not feel the full effect of that nuclear charge e.g. the 1s and 2s electrons shield the 2p electrons from the nuclear charge. The greater the shielding effect, the less an electron is attracted by the nucleus and the lower the ionisation energy. The shielding effect depends on:1. the number of inner energy levels; shielding effect is always greater when an inner energy level is full;
2. the closer the inner electrons are to the nucleus, the greater their shielding effect;
3. a full inner main energy level has a greater shielding effect than an incompletely filled energy level; electrons on the same energy level do NOT shield each other.
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**Ionization energy**

Ionization energy refers to the minimum amount of energy required to remove an electron from one mole of gaseous atoms (or ions); it is measured in kilojoules per 1 mole and is defined by the following equation:

 atom in ground state (g) + ionisation energy ⎯→ atom +  (g) + e -

Ionization energy has a positive value as it is an endothermic process – energy is needed.

The ionization of an atom depends on:

1. the distance of the electron from the nucleus (for the outer most electron this will be the atomic radius) that will be removed from the nucleus/the energy level it is on/size of attraction;
2. the size of the positive nuclear charge (=number of protons);
3. the screening or shielding effect from the nuclear charge by the inner electrons.

The greater the ionisation energies in an atom/ion, the more stable the electron configuration of that atom/ion.

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| 1. first ionisation energy:
2. minimum energy needed to remove first electron from 1 mole of gaseous or free atom to make a gaseous ion;

 X (g) ⎯→ X+ (g) + e- 1. the more strongly the electron is attracted to the nucleus, the greater the amount of energy needed (measured in joules per mole);
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| 1. successive ionisation energy:
* energy needed to remove second, third,... electrons from 1 mole of gaseous ions to form gaseous ions; e.g.

 second ionization energy X+ (g) ⎯→ X2+ (g) + e- 1. the lower the energy level an electron is on, the more energy is needed;
2. successive ionization energies increase as electrons are removed by a positive charge from increasingly more positive ions;
3. successive ionization energies also increase because the electron-electron repulsion decreases (this repulsion puts electrons onto higher energy levels)
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***Horizontal trend***

Overall trend = ionisation energy increases when moving across because more energy is needed because:

1. the increased nuclear charge;
2. smaller atomic radius (outermost electron closer to nucleus);
3. electrons go in the same energy level (similar shielding effect).

The result is a stronger attraction that pulls the valence electrons closer to the nucleus/stronger attraction.

***Vertical trend***

Ionisation energy decreases as you go down groups because:

* Atomic radius increases/outer electron to be removed is further away from nucleus, reducing the attraction between the valence electrons and the nucleus (this offsets the increased nuclear charge).
* Increased shielding as there are more energy levels so there is less effective nuclear charge; increased shielding also offsets increased nuclear charge.

**Electronegativity**

Electronegativity is the ability for an atom to attract a bonding pair of electrons (a shared pair in a diatomic covalent bond). The more commonly used scale is the Pauling scale in which all values are measured relative to fluorine that has the maximum electronegativity of 4.0.

***Horizontal trend***

Increases when moving to the across the periods because:

* the positive nuclear charge increases;
* the atomic radius decreases so the nucleus can attract other electrons better;
* the difference in shielding effect is minimal as all have the same number of filled inner shells.

***Vertical trend***

Decreases as you go down groups:

* the number of main energy levels increases, increasing the shielding effect on free electrons (=electrons part of another atom),
* increased the atomic radius;
* the two above cancel out the increased nuclear charge.

**Ionic radius: cations**

Cations are always smaller than the parent atom from which they are made. This is because:

* when all electrons in an outer main energy level are removed, that energy level is lost;
* when an electron is removed the other electrons in the same energy level experience less repulsion between them and they contract towards the nucleus.

***Horizontal trend***

Ionic radius decreases as the cations have the same number of electrons (= isoelectronic) but the positive charge increases pulling more on the electrons.

***Vertical trend***

Increases as you go down the group as the number of main energy levels increases which offsets the increase in nuclear charge.

**Ionic radius: anions**

Anions are always larger than the atoms they are derived from as the added electrons increase the repulsion between the valence electrons pushing them further away from the nucleus. The electrons are held less tightly.

***Horizontal trend***

Ionic radius decreases as the anions have the same number of electrons (= isoelectronic) but the positive charge increases pulling more on the electrons.

***Vertical trend***

Increases as you go down the group as the number of main energy levels increases which offsets the increase in nuclear charge.

**For all isoelectronic atoms/ions: the greater the atomic number (nuclear charge), the smaller the radius.**

**Electron affinity**

Electron affinity of an element is the enthalpy change that occurs when one electron is gained by each atom in a mole of gaseous atoms of the element to give one mole of gaseous ions, each with a single negative charge (at standard temperature and pressure). Electron affinity can be negative (usually e.g. halogens) which means it is an exothermic process as energy is released or it can be positive (like some group 2 elements or the noble gases in group 18) indicating an endothermic process.

The equation defines first electron affinity: X(g) + e- ⎯→ X- (g)

The more negative the electron affinity value the greater the ability of the atoms of the element to accept electrons.

***Horizontal trend***

Overall electron affinities become more negative as you go across a period because when moving to the across the periods:

* the positive nuclear charge increases;
* the atomic radius decreases so the nucleus can attract other electrons better;
* the difference in shielding effect is minimal as all have the same number of filled inner shells.

***Vertical trend***

Trends differ in different groups but generally electron affinity values become less negative as you go down a group because ...

* the number of main energy levels increases, increasing the shielding effect on free electrons (=electrons part of another atom),
* increased nuclear charge but this is cancelled out by the increased shielding – similar effective nuclear charge;
* increased atomic radius;

**Summary of all physical properties**

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| --- | --- | --- |
| **physical property** | **horizontal trend** | **vertical trend** |
| atomic radius | decreases | increases |
| ionization energy | increases | decreases |
| electronegativity | increases | decreases |
| cation radius | decreases | increases |
| anion radius | decreases | increases |
| electron affinity | increases | decreases |

**Trends in metallic and non-metallic character**

Metals have low ionization energies and are therefore more likely to oxidize and form cations.

Metallic character decreases as you go across a period because:

* the atomic radius decreases
* ionization energy increases
* less likely to form cations

Metallic character increases as you go down a group because:

* atomic radius increases
* ionization energy decreases
* more likely to form cations.

Non-metallic character increases as you go across a period because:

* atomic radius decreases
* electron affinity becomes more negative
* more likely to form anions

Non-metallic character decreases as you go down a group because:

* atomic radius increases
* electron affinity becomes less negative
* less likely to form anions

**Chemical properties: differences and similarities of elements in the same group**

**Alkali metals (Li to Cs)**

*Similarities:*

1. they are all in the same group because they all have 1 valence electron.
2. all alkali metals are very reactive (low ionisation energy); because of their low ionisation energies (their only valence electron is well shielded and is further away from the nucleus ) they are easily oxidised, they are very good reducing agents:

 M (g) ⎯→ M+ (g) + e-

1. they have low first ionization energies but high second ionization energies because of similar electron configuration.
2. they reduce water to hydrogen and hydroxide forming a solution of a **strong** alkali which is why they are called alkali metals: (observations: the metals move around and fizz). Also a gas is produced.

 2Na (s) + 2H2O (l) ⎯→ 2Na+(aq) + 2OH- (aq) + H2 (g)

1. they readily react with air producing distinctly coloured flames as they absorb and emit visible light;
2. they have low melting and boiling points (compared to other metals), low density (they all float on water) and are soft because of weaker - relative - metallic bonds;
3. they react easily with halogens to form white ionic compounds called halides; they form compounds with formula *XY* with halogens. These reactions are highly exothermic.

 2Na (s) + Cl2 (g ⎯→ 2NaCl (s)

*Differences: e.g. observations: Li just floats on water, Na might catch fire, K always catches fire!!!*

More metallic as you go down/more reactive/more ionic:

* more reactive; as you go down the first ionisation energy decreases as outer electrons are further from the nucleus. The outer electrons are also more shielded. Fr is the most reactive with the lowest first ionisation energy, Fr is also the best reducing agent.
* lower m.p/b.p;
* higher density: Li, Na and K float on water whilst the Rb, Cs and Fr sink.

**Halogens (Cl, Br, I)**

*Similarities*

1. 7 electrons in outer shell;
2. all reactive/good oxidising agents (\*remove electrons)/become easily reduced;

 Hal + e- ⎯→ Hal-

1. form halide ions: Hal- /oxidation state is mostly -1 (only fluorine has -1 as its only oxidation state; the other have variable oxidation states)
2. diatomic molecules/single covalent bonds;
3. more soluble in organic solvents (e.g. hexane) than in water;
4. react with each other; more reactive halogen displaces less reactive one (you need to be able to explain the experimental procedure of these displacement reactions and all the observations e.g. colour changes that occur which are summarised below) (See Practical on Halogens).

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|  | **Solution (aq)** | **Chlorine water****(Cl2 (aq))** | **Bromine water****(Br2(aq))** | **Iodine water****(I2(aq))** |
| **Initial colour** | **colourless** | **orange** | **brown** |
| **Potassium chloride (Cl- (aq) = ions)** | Colour after shaking |  | No change | No change |
| Products |  | No reaction | No reaction |
| Equation |  |  |  |
| **Potassium bromide (Br- (aq) = ions)** | Colour after shaking | From colourless to yellow/red |  | No change |
| Products | Bromine + potassium chloride |  | No reaction |
| Equation | Cl2 +2Br-→ Br2 + 2Cl- |  |  |
| **Potassium iodide****(I- (aq) = ions)** | Colour after shaking | From colourless to red/brown | From orange to brown colour |  |
| Products | Iodine + potassium chloride | Iodine + potassium bromide |  |
| Equation | Cl2 +2I-→ I2 + 2Cl- | Br2 +2I- → I2 + 2Br- |  |

The element not included in this table is fluorine because it is far too reactive. Flourine displaces all other halide ions according to the equation:

 F2  + 2Cl-/Br-/I-/At- → Cl/Br/I/At2 + 2F-

* react with water: fluorine and chlorine to form acidic solutions; e.g.

 Cl2 (g) + H2O (l) ⎯→ HCl (aq) + HClO (aq) and 2F2 (g) + 2H2O (l) ⎯→ 4HF (aq) + O2 (aq)

* react with silver ions to form solid silver halides some of which decompose in sunlight as shown by the table below.

|  |  |  |  |
| --- | --- | --- | --- |
| **test** | **chloride (Cl- (aq) = ions)** | **bromide (Br- (aq) = ions)** | **iodide (I- (aq) = ions)** |
| add AgNO3 (aq) | white ppt | pale yellow/cream ppt | yellow ppt |
| effect of sunlight | ppt turns purple grey | ppt turns grey to yellow | no change to ppt |

The above table shows:

* darkening of silver halide as you go down;
* that silver iodide is not sensitive to light but the other two silver halides are (which is why they are used in photography).

As you go down:

* reactivity decreases; fluorine is the most reactive halogen and best oxidising agent (why does the oxidising ability decrease as you go down?) (fluorine has the highest electronegativity because of its smallest radius and little shielding);
	+ **F2  oxidises Cl-, Br- and I- to Cl2, Br2 and I2  because it is the best oxidizing agent; it**

 **has the greatest attraction for electrons .**

* + **Cl2  oxidises Br- and I- to Br2 and I2  because it is a better oxidizing agent.**
	+ **Br2  oxidises I- (but not Cl- or F-) to I2  as it is a better oxidizing agent than I2.**
	+ **I2will not react with F-,Cl-, and Br- as it is the weakest oxidizing agent of the four.**
* astatine ions are oxidized the easiest of all halide ions because of lowest electronegativity of all halogens;
* less soluble in water;
* higher melting and boiling points increase;
* colours get darker.

**Trends in oxides across Period 3**

As you go across Period 3 in the Periodic Table there is a change from metallic to non-metallic character; this is shown by:

* increase in ionisation energy as you go across; this means the elements are less likely to give up electrons and form positive ions and ionic bonds which is typical metallic behaviour;
* properties of the oxides also show a change in properties from metallic to non-metallic:
* metals form oxides which are ionic and can act like bases (in the case of sodium oxide it dissolves easily in water and forms an alkali) and can neutralise acids;

 Na2O (s) + H2O (l) ⎯→ 2 NaOH (aq)

* Al2O3 is an amphoteric oxide which means it can neutralise both acids and bases;
* non-metals form oxides which dissolve in water – except silicon dioxide – to form acids; they are called acidic oxides and can neutralise bases or alkalis.

Equation acidic oxide dissolving: SO3 + H2O ⎯→ H2SO4

* high electronegativity towards the right as the atoms have greater ability to attract shared electrons in a covalent bond.
* argon does not form any compounds as it is a noble gas with a full outer principal energy level.

**Summary of some properties of the period 3 oxides**

|  |  |
| --- | --- |
| Group | 1 2 13 14 15 16 17 18 |
| Element  |  Na Mg Al Si P S Cl Ar  |
| Structure of element |  giant←⎯ giant metallic ⎯→ covalent ←⎯ simple molecular ⎯→ |
| Formula of oxide |  Na2O MgO Al2O3 SiO2 P4O10 SO3 Cl2O7  P4O6 SO2 Cl2O |
| Structure of oxide | ←⎯ giant ionic ⎯→ giant ←⎯ simple molecular ⎯→ covalent |
| Acid-base character of oxide | ←⎯ basic ⎯→ amphoteric ←⎯⎯⎯⎯⎯ acidic ⎯⎯⎯⎯⎯→ no oxide |

Equations to show the behavior of some of the oxides when reacting with water:

The basic behaviour of sodium and magnesium oxide:

Na2O (s) + H2O (l) ⎯→ 2Na+(aq) + 2OH- (aq) /2NaOH (aq)

MgO (s) + H2O (l) ⎯→ Mg2+(aq) + 2OH- (aq) /Mg(OH)2 (aq)

The acidic behavior of phosphorus oxide and nitrogen and sulphur oxides:

P4 O10 (s) + 6H2O (l) ⎯→ 4H3PO4 (aq)

NO2 (g) + H2O (l) ⎯→ HNO3 (aq) + HNO2 (aq)

SO2 (g) + H2O (l) ⎯→ H2SO3 (aq)

SO3 (g) + H2O (l) ⎯→ H2SO4 (aq)

The amphoteric behavior of phosphorus oxide and nitrogen and sulphur oxides:

Al2O3 (s) + 6HCl (l) ⎯→ AlCl3 (aq) + 3H2O(aq)

Al2O3 (s) + 2NaOH (l) ⎯→ 2NaAlO2 (aq) + H2O(aq)