Matter and material: The periodic table

Objectives

- Describe the periodic table as displaying elements in order of increasing atomic number and showing how periodicity of physical and chemical properties of elements relates to atomic structure.
- Define the group number and the period number.
- Relate the period number to the outermost energy level in which electrons exist.
- Relate the group number to the electronic structure in terms of valence electrons and valency.
- Describe periodicity in terms of atomic radius and ionisation energy.

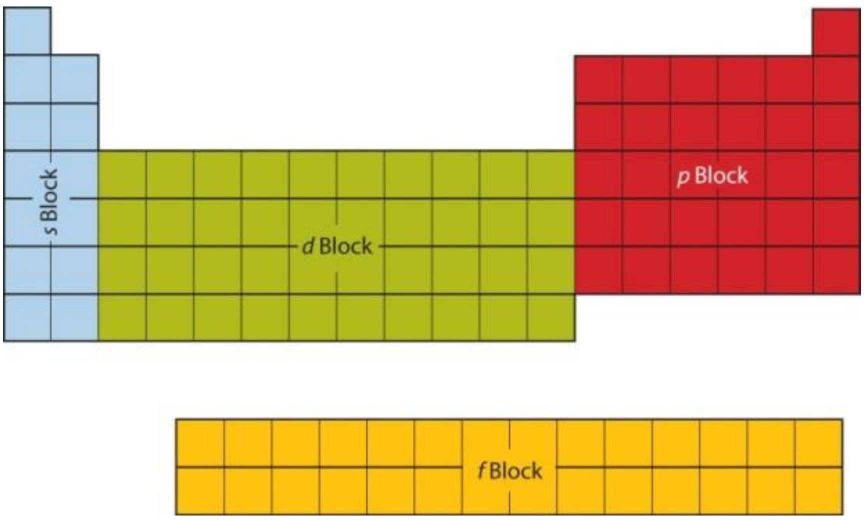
Systematic arrangement of the elements

The periodic table is a systematic way to arrange elements. The periodic table displays the elements in increasing atomic number. The atomic number represents the number of protons in the nucleus. The properties of the elements are also based on the electronic structure of the atoms since valence electrons are responsible for chemical bonding. Recall that valence electrons are the electrons in the outermost energy level of an atom. The elements in the periodic table are **grouped** based on the number of valence electrons.

Each **period** in the periodic table starts with atoms with electrons in a higher energy level. The maximum number of electrons in the outermost energy level of each period corresponds to the number of elements in a period in the periodic table. For the 7 periods, the numbers are:

- In the first energy level (n=1), there is a maximum of 2 electrons.
- In the second energy level (n=2), there is a maximum of 8 electrons.
- In the third energy level (n=3), there is a maximum of 8 electrons.
- In the fourth energy level (n=4), there is a maximum of 18 electrons.
- In the fifth energy level (n=5), there is a maximum of 18 electrons.
- In the sixth energy level (n=6), there is a maximum of 32 electrons.
- In the seventh energy level (n=7), there is a maximum of 32 electrons.

Groups 1 and 2 elements make up the s block meaning that all their outermost valence electrons are in the s shell. Group 13-18 elements make up the p block meaning that all their outermost electrons are in the p shell. Groups 3-12 elements (except elements 57-71 and 89-103) make up the d block meaning that all their outermost electrons are in the d shell. Metals in the d block are known as **transition metals**. Elements 57-71 and 89-103 at the bottom of the periodic table make up the f block meaning that all their outermost electrons are in the f shell. See the diagram below for the block groupings. We will mostly consider the s and the p blocks with a few elements that we need to be aware of in the d block.

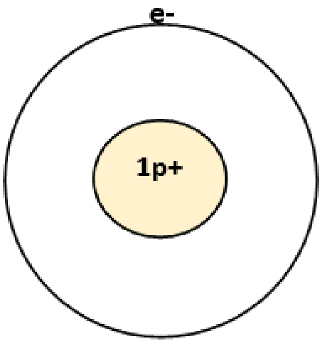


All of this may sound confusing now, but it will make sense by the end of this section.

Group 1 elements

Group 1 elements are known as **alkali metals** (the name comes from the fact that when these metals or their oxides are dissolved in water, the result is a basic or alkaline solution. Alkali metals are highly reactive and are always found as ionic compounds in nature). Let's consider **group 1** elements hydrogen (H), lithium (Li), sodium (Na) and potassium (K).

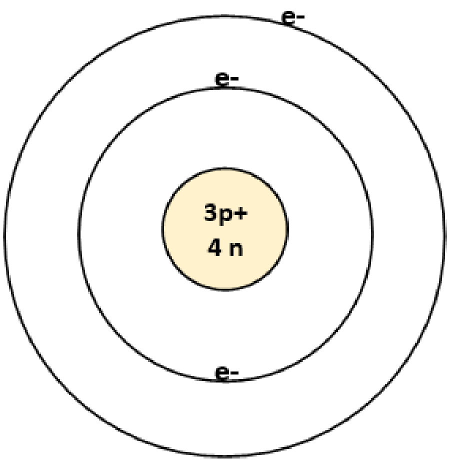
A hydrogen atom has one electron:



This means that:

- Hydrogen has one valence electron in the first energy level n=1.
- The first energy level can fit a maximum of two electrons. Recall that all atoms would like to have a full outer energy level.
- Therefore, to have a full outer energy level, hydrogen may lose one electron to become $\mathbf{H^+}$ or gain one electron to become $\mathbf{H^-}$.

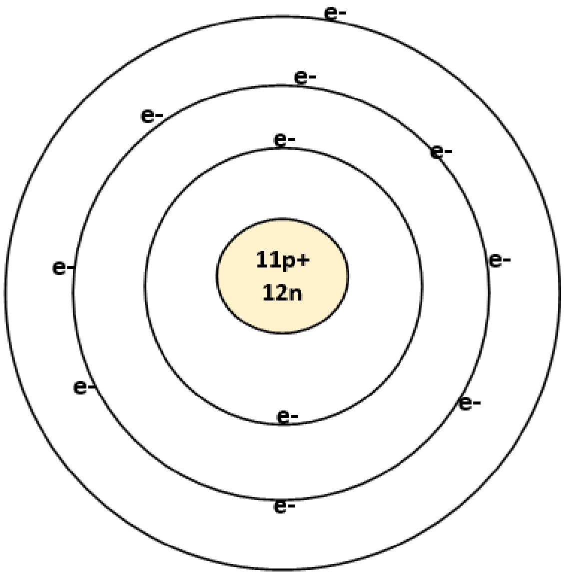
A lithium atom has three electrons.



This means that:

- Lithium has one valence electron in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, lithium may lose one electron to become $\mathbf{Li^+}$.

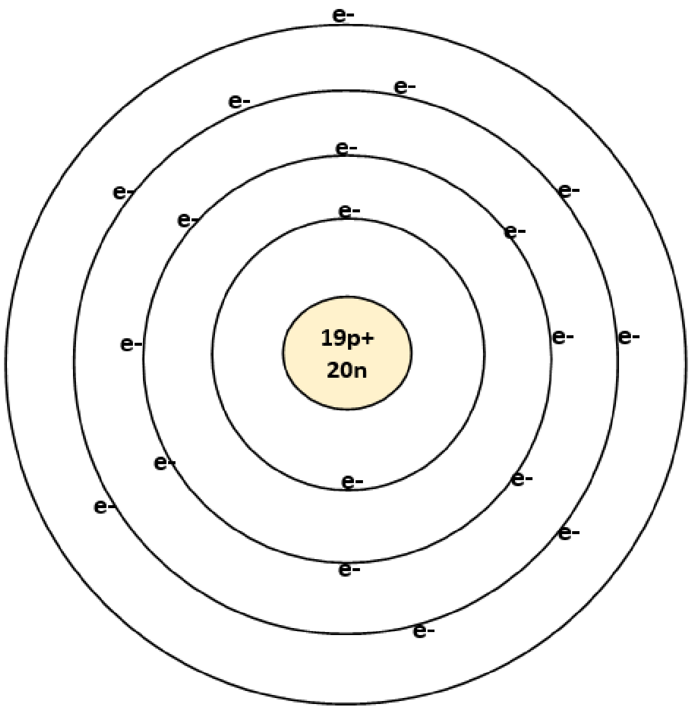
A sodium atom has 11 electrons.



This means that:

- Sodium has one valence electron in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, sodium may lose one electron to become $\mathbf{Na^+}$.

A potassium atom has 19 electrons.



This means that:

- Potassium has one valence electron in the fourth energy level n=4.
- The fourth energy level can fit a maximum of 18 electrons.
- Therefore, to have a full outer energy level, potassium may lose one electron to become $\mathbf{K^+}$.

Now let's consider the electron configuration of group 1 elements:

- Hydrogen's electron configuration is $1s^1$. This shows us that hydrogen's outermost valence electron is in the s shell.
  If hydrogen loses its valence electron, all that is left is a proton. If hydrogen gains an electron, it will form the H^- ion and its electron configuration will be $1s^2$. This is the same electron configuration as helium, a noble gas.

- Lithium's electron configuration is $1s^2 2s^1$. This shows us that lithium's outermost valence electron is in the s shell.
  If lithium loses its valence electron, it will form the Li^+ ion and its electron configuration will be $1s^2$. This is the same electron configuration as helium, a noble gas.

- Sodium's electron configuration is $1s^2 2s^2 2p^6 3s^1$. This shows us that sodium's outermost valence electron is in the s shell.
  If sodium loses its valence electron, it will form the Na^+ ion and its electron configuration will be $1s^2 2s^2 2p^6$. This is the same electron configuration as neon, a noble gas.

- Potassium's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. This shows us that potassium's outermost valence electron is in the s shell.
  If potassium loses its valence electron, it will form the K^+ ion and its electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^6$. This is the same electron configuration as argon, a noble gas.

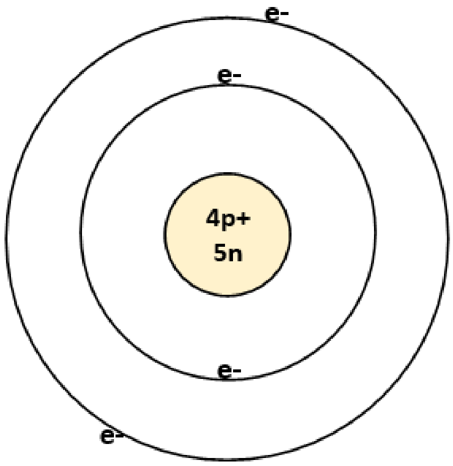
Here are three very important things that we can note from this:
1) All group 1 elements have **one valence electron**.
2) A group 1 element will tend to lose its one valence electron to have a full outer energy level and form a 1+ ion. This means that group 1 elements have a **valency of 1+**. (Exception: hydrogen may gain an electron to form a 1– ion).
3) All group 1 elements have their valence electron in the **s shell** of their outermost energy level.

Why do group 1 elements tend to lose one electron rather than gain seven electrons? It takes less energy to lose one electron than to gain seven. Because these atoms only need to lose one electron, they are highly reactive. Another very important thing to note is that an ion of a group 1 element will have the same electron configuration as a noble gas. In fact, this applies to ALL ions. **All ions have the electron configuration of a noble gas.** This is because a noble gas has a full outer energy level, and all ions become ions to have a full outer energy level.

Group 2 elements

Group 2 elements are known as **alkaline earth metals** (the name comes from the fact that when these metals or their oxides are dissolved in water, a basic/alkaline solution results; and the fact that these metal oxides are commonly found in the earth's crust). Group 2 elements are very reactive. Let's consider **group 2** elements beryllium (Be), magnesium (Mg) and calcium (Ca).

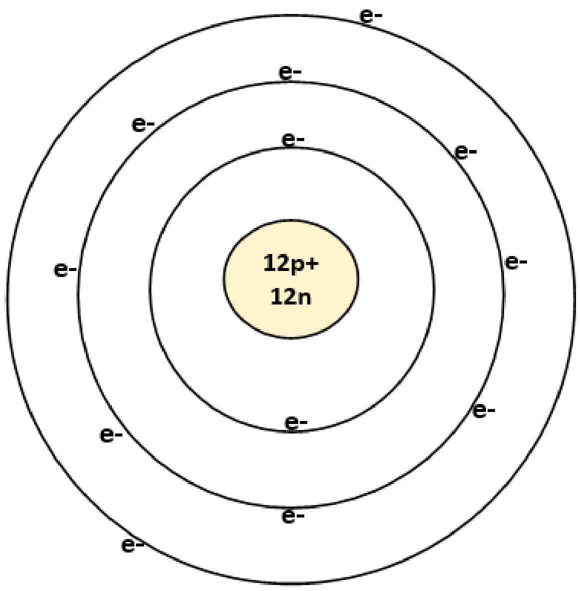
A beryllium atom has four electrons.



This means that:

- Beryllium has two valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, beryllium may lose two electrons to become $\mathbf{Be^{2+}}$.

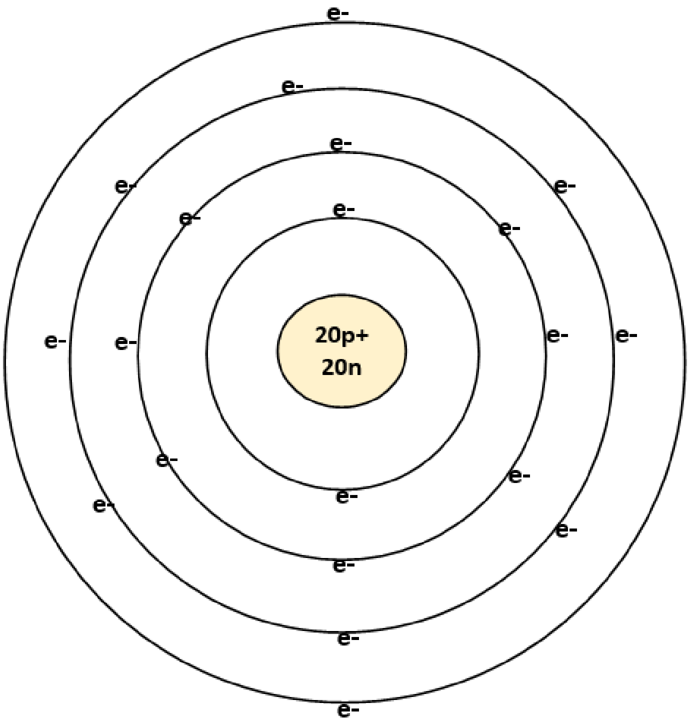
A magnesium atom has 12 electrons.



This means that:

- Magnesium has two valence electrons in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, magnesium may lose two electrons to become $\mathbf{Mg^{2+}}$.

A calcium atom has 20 electrons.



This means that:

- Calcium has two valence electrons in the fourth energy level n=4.
- The fourth energy level can fit a maximum of 18 electrons.
- Therefore, to have a full outer energy level, calcium may lose two electrons to become $\mathbf{Ca^{2+}}$.

Now let's consider the electron configuration of group 2 elements:

- Beryllium's electron configuration is $1s^2 2s^2$. This shows us that beryllium's outermost valence electrons are in the s shell.
  If beryllium loses its two valence electrons, it will form the Be^{2+} ion and its electron configuration will be $1s^2$. This is the same electron configuration as helium, a noble gas.

- Magnesium's electron configuration is $1s^2 2s^2 2p^6 3s^2$. This shows us that magnesium's outermost valence electrons are in the s shell.
  If magnesium loses its two valence electrons, it will form the Mg^{2+} ion and its electron configuration will be $1s^2 2s^2 2p^6$. This is the same electron configuration as neon, a noble gas.

- Calcium's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$. This shows us that calcium's outermost valence electrons are in the s shell.
  If calcium loses its two valence electrons, it will form the Ca^{2+} ion and its electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^6$. This is the same electron configuration as argon, a noble gas.

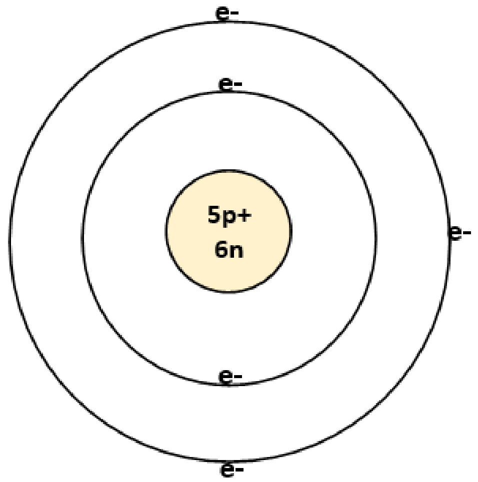
Here are three very important things that we can note from this:
1) All group 2 elements have **two valence electrons**.
2) A group 2 element will tend to lose its two valence electrons to have a full outer energy level and form a 2+ ion. This means that group 2 elements have a **valency of 2+**.
3) All group 2 elements have their valence electrons in the **s shell** of their outermost energy level.

Because these atoms only need to lose two electrons, they are very reactive. Again, note that an ion of a group 2 element will have the same electron configuration as a noble gas. This applies to all ions.

Group 13 elements

We will now skip groups 3-12 (the d block) and consider **group 13** elements boron (B) and aluminium (Al).

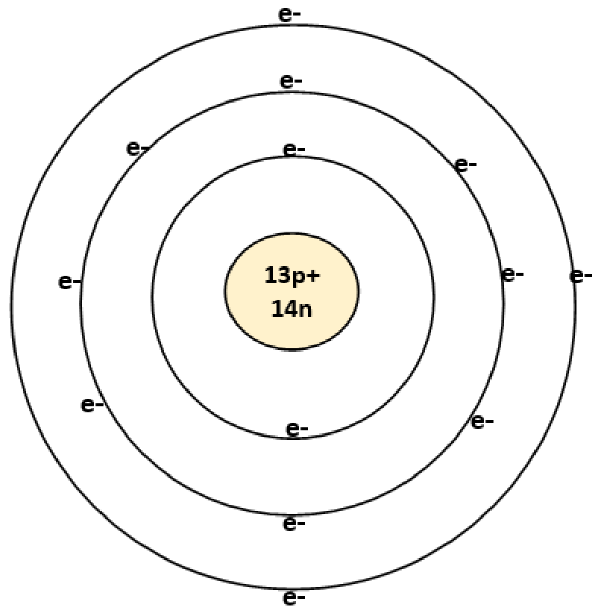
A boron atom has five electrons.



This means that:

- Boron has three valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, boron may lose two electrons to become $\mathbf{B^{3+}}$.

An aluminium atom has 13 electrons.



This means that:

- Aluminium has three valence electrons in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, aluminium may lose two electrons to become $\mathbf{Al^{3+}}$.

Now let's consider the electron configuration of group 13 elements:

- Boron's electron configuration is $1s^2 2s^2 2p^1$. This shows us that boron's outermost valence electrons are in the p shell.
  If boron loses its three valence electrons, it will form the B^{3+} ion and its electron configuration will be $1s^2$. This is the same electron configuration as helium, a noble gas.

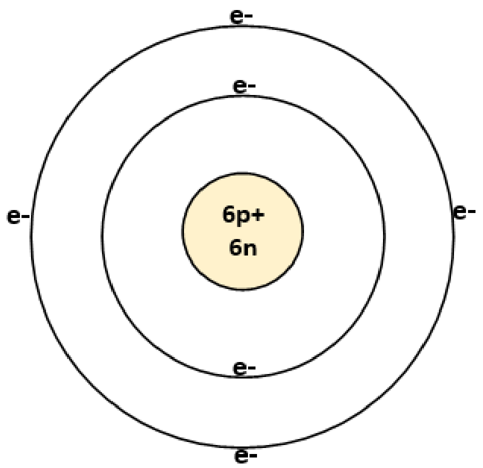
- Aluminium's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^1$. This shows us that aluminium's outermost valence electrons are in the p shell.
  If aluminium loses its three valence electrons, it will form the Al^{3+} ion and its electron configuration will be $1s^2 2s^2 2p^6$. This is the same electron configuration as neon, a noble gas.

Here are three very important things that we can note from this:
1) All group 13 elements have **three valence electrons**.
2) A group 13 element will tend to lose its three valence electrons to have a full outer energy level and form a 3+ ion. This means that group 13 elements have a **valency of 3+**.
3) All group 13 elements have their valence electron in the **p shell** of their outermost energy level.

Group 14 elements

Let's consider **group 14** elements carbon (C) and silicon (Si).

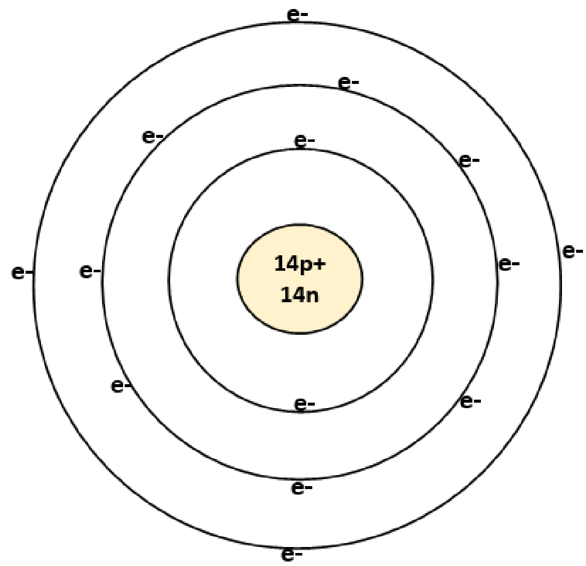
A carbon atom has six electrons.



This means that:

- Carbon has four valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, carbon may lose OR gain four electrons.

A silicon atom has 14 electrons.



This means that:

- Silicon has four valence electrons in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, silicon may lose OR gain four electrons.

Now let's consider the electron configuration of group 14 elements:

- Carbon's electron configuration is $1s^2 2s^2 2p^2$. This shows us that carbon's outermost valence electrons are in the p shell. In an ionic form, carbon will have the same electron configuration of a noble gas. However, carbon rarely forms an ion.

- Silicon's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^2$. This shows us that silicon's outermost valence electrons are in the p shell. In an ionic form, silicon will have the same electron configuration of a noble gas. However, silicon rarely forms an ion.

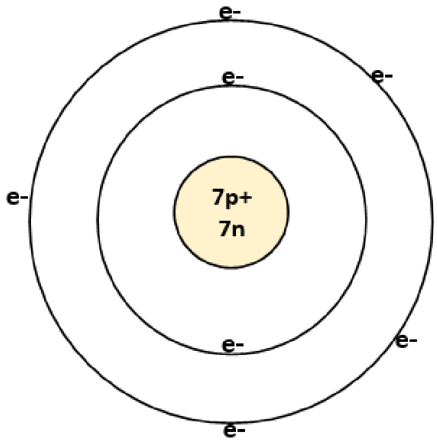
Here are three very important things that we can note from this:
1) All group 14 elements have **four valence electrons**.
2) A group 14 element may lose its four valence electrons or gain four electrons to have a full outer energy level. This means that group 14 elements have a **valency of 4+ or 4–**. (In group 14, only lead and tin tend to form ions).
3) All group 14 elements have their valence electron in the **p shell** of their outermost energy level.

In group 14, only lead and tin tend to form ions, the other elements do not generally form ions. Ions are formed when electrons are gained or lost. This results in the formation of an ion. Ions then chemically bond to other ions to form an **ionic** chemical bond. Elements like carbon and silicon tend to get involved in **molecular / covalent** chemical bonds, which does not involve the formation of an ion. This will be clarified in the next section on chemical bonding.

Group 15 elements

Let's consider **group 15** elements nitrogen (N) and phosphorus (P).

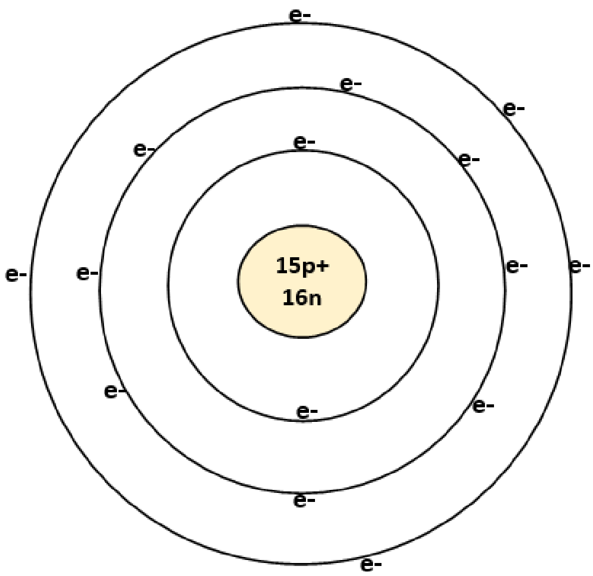
A nitrogen atom has seven electrons.



This means that:

- Nitrogen has five valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, nitrogen may gain three electrons to become $\mathbf{N^{3-}}$.

A phosphorus atom has 15 electrons.



This means that:

- Phosphorus has five valence electrons in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, phosphorus may gain three electrons to become $\mathbf{P^{3-}}$.

Now let's consider the electron configuration of group 15 elements:

- Nitrogen's electron configuration is $1s^2 2s^2 2p^3$. This shows us that nitrogen's outermost valence electrons are in the p shell.
  If nitrogen gains three valence electrons, it will form the N^{3-} ion and its electron configuration will be $1s^2 2s^2 2p^6$. This is the same electron configuration as neon, a noble gas.

- Phosphorus's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^3$. This shows us that phosphorus's outermost valence electrons are in the p shell.
  If Phosphorus gains three valence electrons, it will form the P^{3-} ion and its electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^6$. This is the same electron configuration as argon, a noble gas.

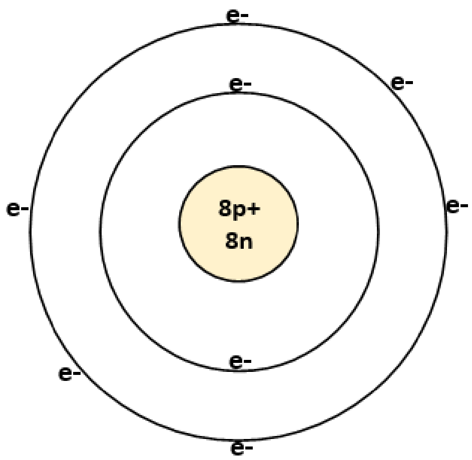
Here are three very important things that we can note from this:
1) All group 15 elements have **five valence electrons**.
2) A group 15 element will tend to gain three electrons to have a full outer energy level and form a 3– ion. This means that group 15 elements have a **valency of 3–**.
3) All group 15 elements have their valence electrons in the **p shell** of their outermost energy level.

Why do group 15 elements tend to gain three electrons rather than lose five electrons? It takes less energy to gain three electrons than to lose five. The ions of nonmetals have a name that ends in -ide. For example, the nitrogen ion is called nitride; the phosphorus ion is called phosphide.

Group 16 elements

Let's consider **group 16** elements oxygen (O) and sulfur (S).

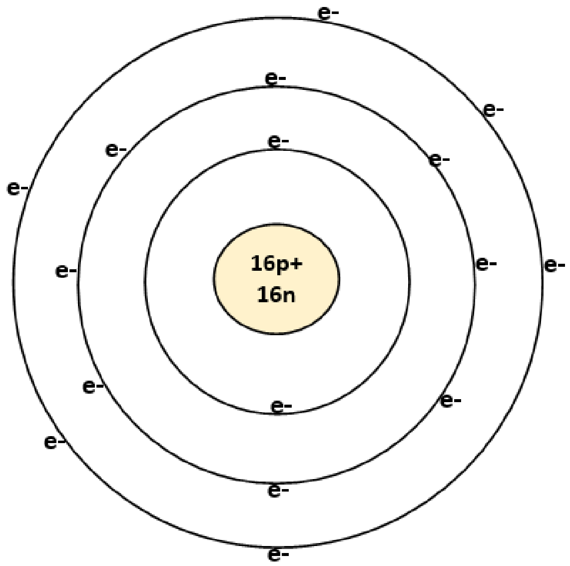
An oxygen atom has eight electrons.



This means that:

- Oxygen has six valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, oxygen may gain two electrons to become $\mathbf{O^{2-}}$.

A sulfur atom has 16 electrons.



This means that:

- Sulfur has six valence electrons in the third energy level n=3.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, sulfur may gain two electrons to become $\mathbf{S^{2-}}$.

Now let's consider the electron configuration of group 16 elements:

- Oxygen's electron configuration is $1s^22s^22p^4$. This shows us that oxygen's outermost valence electrons are in the p shell.
  If oxygen gains two valence electrons, it will form the O^{2-} ion and its electron configuration will be $1s^22s^22p^6$. This is the same electron configuration as neon, a noble gas.

- Sulfur's electron configuration is $1s^22s^22p^63s^23p^4$. This shows us that sulfur's outermost valence electrons are in the p shell.
  If sulfur gains two valence electrons, it will form the S^{2-} ion and its electron configuration will be $1s^22s^22p^63s^23p^6$. This is the same electron configuration as argon, a noble gas.

Here are three very important things that we can note from this:
1) All group 16 elements have **six valence electrons**.
2) A group 16 element will tend to gain two electrons to have a full outer energy level and form a 2– ion. This means that group 16 elements have a **valency of 2–**.
3) All group 16 elements have their valence electrons in the **p shell** of their outermost energy level.

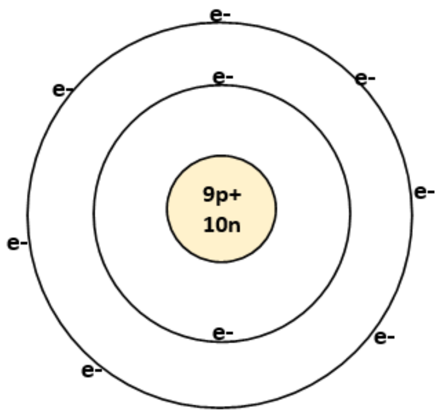
Recall that the ions of nonmetals have a name that ends in -ide. For example, the oxygen ion is called oxide; the sulfur ion is called sulfide. Note that a compound formed with oxygen is known as an oxide *[D], e.g. iron oxide and carbon dioxide are both examples of oxide compounds.

*[D] represents a definition. Definitions need to memorised word for word.

Group 17 elements

Group 17 elements are known as **halogens**. Compounds formed with halogens are known as halides [D]. For example, chlorine is a halogen and silver chloride is a halide. Let's consider **group 17** elements fluorine (F) and chlorine (Cl).

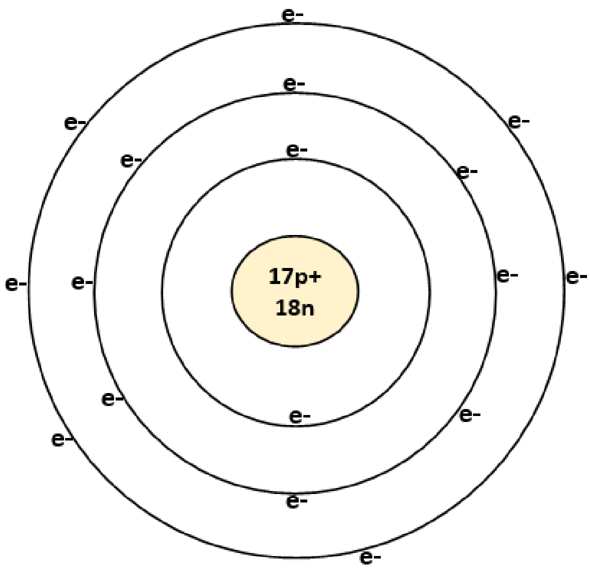
A fluorine atom has nine electrons.



This means that:

- Fluorine has seven valence electrons in the second energy level n=2.
- The second energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, fluorine may gain an electron to become **F^-**.

A chlorine atom has 17 electrons.



This means that:

- Chlorine has seven valence electrons in the third energy level n=2.
- The third energy level can fit a maximum of eight electrons.
- Therefore, to have a full outer energy level, chlorine may gain an electron to become **Cl^-**.

Now let's consider the electron configuration of group 17 elements:

- Fluorine's electron configuration is $1s^2 2s^2 2p^5$. This shows us that fluorine's outermost valence electrons are in the p shell.
  If fluorine gains one electron, it will form the $F^–$ ion and its electron configuration will be $1s^2 2s^2 2p^6$. This is the same electron configuration as neon, a noble gas.

- Chlorine's electron configuration is $1s^2 2s^2 2p^6 3s^2 3p^5$. This shows us that chlorine's outermost valence electrons are in the p shell.
  If chlorine gains one electron, it will form the $Cl^–$ ion and its electron configuration will be $1s^2 2s^2 2p^6 3s^2 3p^6$. This is the same electron configuration as argon, a noble gas.

Here are three very important things that we can note from this:
1) All group 17 elements have **seven valence electrons**.
2) A group 17 element will tend to gain an electron to form a 1– ion. This means that group 17 elements have a **valency of 1–**.
3) All group 17 elements have their valence electrons in the **p shell** of their outermost energy level.

Recall that the ions of nonmetals have a name that ends in -ide. For example, the fluorine ion is called fluoride; the chlorine ion is chloride.

Group 18 elements

Group 18 contains the unreactive gases such as helium, neon and argon. They are called **noble gases** or **inert gases** and they do not form molecules with other elements, including with each other, because each atom already has a full outer energy level and is in a low energy state. This makes these atoms very stable and unreactive. Helium has two electrons in its outermost energy level and neon and argon have eight. As we saw above, all ions have the electron configuration of a noble gas since this means a stable electron arrangement.

Now let's go back to the periodic table and note a few things (see the periodic table on the next page):

(1) Period number: the period number corresponds to the energy level in which the valence electrons exist. Each period in the periodic table starts an element whose electrons are in a higher energy level.
- Period 1 elements, H and He, have their outermost electrons in the first energy level n=1.
- Period 2 elements Li to Ne have their outermost electrons in the second energy level n=2.
- Period 3 element Na to Ar have their outermost electrons in the third energy level n=3 and so on.
- (Elements 58-71 are part of n=6 and elements 90-103 are part of n=7).

(2) Group number: the group number corresponds to the number of valence electrons and valency.
- Group 1 elements have 1 valence electron and a valency of 1+.
- Group 2 elements have 2 valence electrons and a valency of 2+.
- Group 13 elements have 3 valence electrons and a valency of 3+.
- Group 14 elements have 4 valence electrons and a valency of 4+/–.
- Group 15 elements have 5 valence electrons and a valency of 3–.
- Group 16 elements have 6 valence electrons and a valency of 2–.
- Group 17 elements have 7 valence electrons and a valency of 1–.
- Group 18 elements are noble gases and have a full outer energy level. He has two valence electrons; Ne and Ar have eight valence electrons each.

(3) Shells:
- Groups 1 and 2 elements make up the s block meaning that their valence electrons are in an s shell.
- Group 13-18 elements make up the p block meaning that their valence electrons are in a p shell.
- Elements in groups 3-12 make up the d block, meaning that their valence electrons are in a d shell. Metals in the d block are known as **transition metals**.
- Elements 57-71 and 89-103 at the bottom of the periodic table make up the f block, meaning that their valence electrons are in an f shell.

	Alkali metals	Alkaline earth metals															Halogens	Noble/inert gases
Valence e-	1	2											3	4	5	6	7	8
Valency	1+	2+											3+	4+/–	3–	2–	1–	0
n=1	1 2,1 H 1																	2 He 4
n=2	3 1,0 Li 7	4 1,5 Be 9											5 2,0 B 11	6 2,5 C 12	7 3,0 N 14	8 3,5 O 16	9 4,0 F 19	10 Ne 20
n=3	11 0,9 Na 23	12 1,2 Mg 24											13 1,5 Aℓ 27	14 1,8 Si 28	15 2,1 P 31	16 2,5 S 32	17 3,0 Cℓ 35,5	18 Ar 40
n=4	19 0,8 K 39	20 1,0 Ca 40	21 1,3 Sc 45	22 1,5 Ti 48	23 1,6 V 51	24 1,6 Cr 52	25 1,5 Mn 55	26 1,8 Fe 56	27 1,8 Co 59	28 1,8 Ni 59	29 1,9 Cu 63,5	30 1,6 Zn 65	31 1,6 Ga 70	32 1,8 Ge 73	33 2,0 As 75	34 2,4 Se 79	35 2,8 Br 80	36 Kr 84
n=5	37 0,8 Rb 86	38 1,0 Sr 88	39 1,2 Y 89	40 1,4 Zr 91	41 Nb 92	42 1,8 Mo 96	43 1,9 Tc	44 2,2 Ru 101	45 2,2 Rh 103	46 2,2 Pd 106	47 1,9 Ag 108	48 1,7 Cd 112	49 1,7 In 115	50 1,8 Sn 119	51 1,9 Sb 122	52 2,1 Te 128	53 2,5 I 127	54 Xe 131
n=6	55 0,7 Cs 133	56 0,9 Ba 137	57 La 139	72 1,6 Hf 179	73 Ta 181	74 W 184	75 Re 186	76 Os 190	77 Ir 192	78 Pt 195	79 Au 197	80 Hg 201	81 1,8 Tℓ 204	82 1,8 Pb 207	83 1,9 Bi 209	84 2,0 Po	85 2,5 At	86 Rn
n=7	87 0,7 Fr	88 0,9 Ra 226	89 Ac															

58 Ce 140	59 Pr 141	60 Nd 144	61 Pm	62 Sm 150	63 Eu 152	64 Gd 157	65 Tb 159	66 Dy 163	67 Ho 165	68 Er 167	69 Tm 169	70 Yb 173	71 Lu 175
90 Th 232	91 Pa	92 U 238	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

s block

p block

Periodicity of the elements

Periodicity refers to a property that varies uniformly with the arrangement of the elements in the periodic table. There are many physical and chemical properties of the elements that follow a sequence in the periodic table across a period and down a group.

Valence electrons and valency

We have already noted the periodicity of elements regarding the number of valence electrons and valency. How does this correspond to the way that atoms react? Let's consider the formation compounds that are held together by a type of bond called an **ionic bond.** This type of chemical bond forms when a positively charged ion and a negatively charged ion attract due to their electrostatic force of attraction. The **metal** always forms a **cation** by transferring an electron/s to the nonmetal, making the **nonmetal** an **anion**. We need to understand that all the compounds that will result are neutral, so we need to balance the charge of the cation and anion. Here are some examples:

Lithium and fluorine make lithium fluoride: $Li^+ + F^- \rightarrow LiF$
Lithium can only become Li^+ since it wants to give up its valence electron to have a full outer energy level; fluorine can only become F^- since it wants to gain an electron to have a full outer energy level. The positive charge of Li and the negative charge of F balance out, forming LiF.

Sodium and oxygen make sodium oxide: $Na^+ + O^{2-} \rightarrow Na_2O$
Sodium can only become Na^+ since it wants to give up its valence electron to have a full outer energy level; oxygen can only become O^{2-} since it wants to gain two electrons to have a full outer energy level. The O has a negative charge of 2–, and therefore 2 Na^+ need to balance out with one O^{2-} to form a neutral compound, Na_2O.

Aluminium and chlorine make aluminium chloride: $Al^{3+} + Cl^- \rightarrow AlCl_3$
Aluminium can only become Al^{3+} since it wants to give up its three valence electrons to have a full outer energy level; chlorine can only become Cl^- since it wants to gain an electron to have a full outer energy level. The Al has a positive charge of 3+, and therefore 3 Cl^- need to balance out with one Al^{3+} to form a neutral compound, $AlCl_3$

We may also use the crossing method. Lithium has a valency of 1+ so the 1 transfers to the subscript of F. Fluoride has a valency of 1– so the 1 transfers to the subscript of Li. This results in Li_1F_1, which is the same as LiF.

$Li^+ + F^-$

$Li^+ + F^-$

$\rightarrow LiF$

Sodium has a valency of 1+ so the 1 transfers to the subscript of O. Oxygen has a valency of 2– so the 2 transfers to the subscript of Na. This results in Na_2O_1, which is the same as Na_2O.

$Na^+ + O^{2-}$

$Na^+ + O^{2-}$

$\rightarrow Na_2O$

Aluminium has a valency of 3+ so the 3 transfers to the subscript of Cl. Chlorine has a valency of 1– so the 1 transfers to the subscript of Al. This results in Al_1Cl_3, which is the same as $AlCl_3$.

$Al^{3+} + Cl^{-}$

$Al^{3+} + Cl^{-}$

$\rightarrow AlCl_3$

This method can help us solve a slightly more confusing example: aluminium and oxygen make aluminium oxide: $Al^{3+} + O^{2-} \rightarrow Al_2O_3$

Aluminium can only become Al^{3+} since it wants to give up its three valence electrons to have a full outer energy level; oxygen can only become O^{2-} since it wants to gain two electrons to have a full outer energy level.

The Al has a positive charge of 3+, and O has a negative charge of 2–. The number 6 is the lowest common multiple of 2 and 3. Therefore, two Al^{3+} will balance out with three O^{2-} to form a neutral compound, Al_2O_3.

$Al^{3+} + O^{2-}$

$Al^{3+} + O^{2-}$

$\rightarrow Al_2O_3$

If you are going to use the crossing method, be careful in a situation like this: aluminium and nitrogen make aluminium nitride: $Al^{3+} + N^{3-} \rightarrow AlN$

Aluminium can only become Al^{3+} since it wants to give up its three valence electrons to have a full outer energy level; nitride can only become N^{3-} since it wants to gain three electrons to have a full outer energy level. The three positive charges of Al and the three negative charges of N balance out, forming AlN.

$Al^{3+} + N^{3-}$

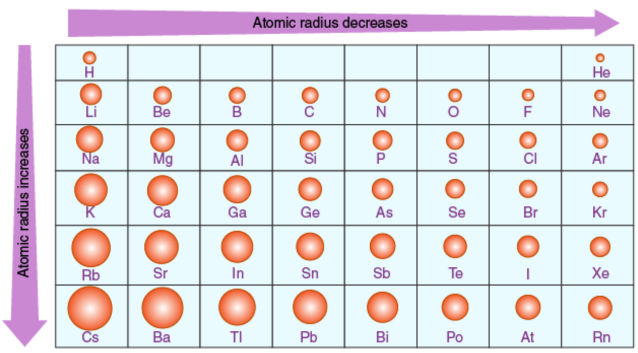
$Al^{3+} + N^{3-}$

This does not form Al_3N_3! In a situation like this, find simplest ratio of the subscript numbers. We may simplify 3:3 to 1:1. Therefore, aluminium nitride is

$\rightarrow AlN$

Atomic radius

Atomic radius [D] is a measure of the size of an atom. The size of the atom is determined by the position of its outermost valence electrons. See the trends in atomic radius in the diagram below:



Why does atomic radius increase as we move down a group?

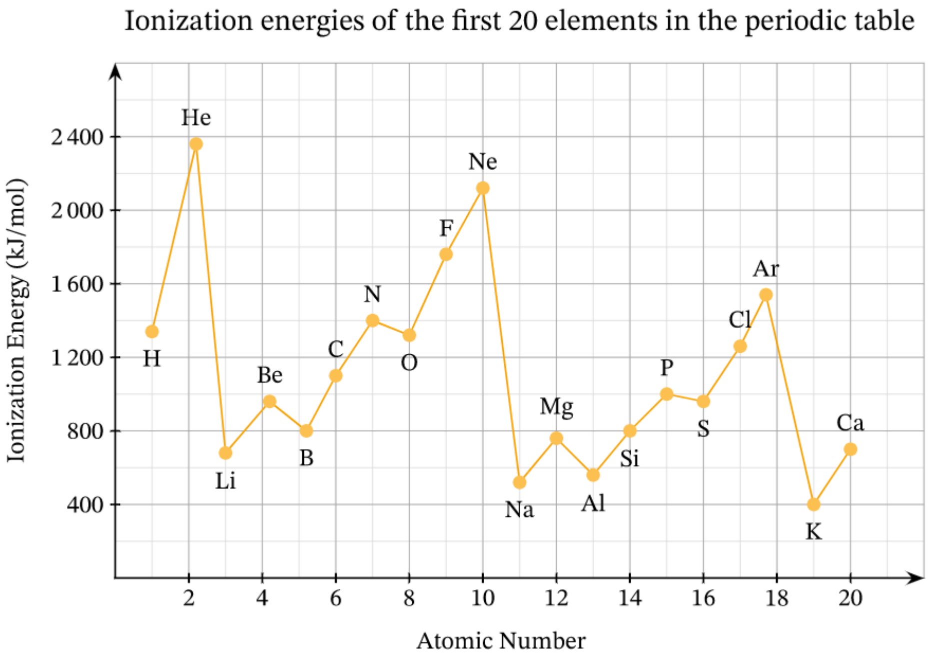
The outermost electron of H is in n=1; the outermost electron of Li is in n=2; the outermost electron of Na is in n=3 and so on. Each energy level is further away from the nucleus. For example, n=3 is further away from the nucleus than n=1, making Na a larger atom than H.

Why does atomic radius decrease as we move across a period?

Consider period 3: all the elements in period 3 have their outermost electrons in n=3. However, sodium has 11 protons in its nucleus while phosphorus has 15 protons in its nucleus. In the case of phosphorus, each electron experiences a stronger electrostatic pull towards the nucleus relative to the electrons of sodium, since there are more protons in the nucleus of phosphorus. This stronger electrostatic pull on the electrons makes the atomic radius of the atom smaller.

Ionisation energy

Ionisation energy [D] is the energy needed per mole to remove an electron from a neutral atom in the gas phase. First ionisation energy refers to the energy required to remove the outermost electron of an atom – the easiest one to remove. Second ionisation energy refers to the energy required to remove the next outermost electron of an atom and so on. See the trends in the first ionisation energy in the graph below:



Why does ionisation energy decrease as we move down a group?

The ionisation energy of H is approximately 1300 kJ/mol; the ionisation energy of Li is approximately 700 kJ/mol; the ionisation energy of Na is approximately 500 kJ/mol; and the ionisation energy of K is approximately 400 kJ/mol.

The outermost electron of H is in n=1; the outermost electron of Li is in n=2; the outermost electron of Na is in n=3 and so on. Each energy level is further away from the nucleus. Being further away from the nucleus means a weaker electrostatic force of attraction on the electrons from the nucleus, making electrons further from the nucleus easier to remove, meaning that less energy is needed to remove them.

Why does ionisation energy increase as we move across a period?

Consider period 3 elements sodium and phosphorus. The ionisation energy of Na is approximately 500 kJ/mol while the ionisation energy of P is approximately 1000 kJ/mol.

All the elements in period 3 have their outermost electrons in n=3. However, sodium has 11 protons in its nucleus while phosphorus has 15 protons in its nucleus. In the case of phosphorus, each electron experiences a stronger electrostatic pull towards the nucleus relative to the electrons of sodium, since there are more protons in the nucleus of phosphorus. This stronger electrostatic pull on the electrons makes the electrons harder to remove, meaning that more energy is needed to remove them.

Practice questions and answers

Practice questions

1. Consider the formation of the following ionic compounds based on their valency. Write the ionic equation of each:
   1.1 Potassium and nitrogen form potassium nitride.
   1.2 Beryllium and iodine form beryllium iodide.
   1.3 Magnesium and sulfur form magnesium sulfide. :
   1.4 Barium and phosphorus form barium phosphide.
   1.5 Calcium and oxygen form calcium oxide.

2. Identify a halide and an oxide from the list above.

3. Which element has a larger atomic radius – silicon or tin? Explain.

4. Which element has a smaller atomic radius – calcium or zinc? Explain.

5. Which element has a higher first ionisation energy – silicon or tin? Explain.

6. Which element has a higher first ionisation energy – calcium or zinc? Explain.

Practice question answers

1.
   1.1 $K^+ + N^{3-} \rightarrow K_3N$
   1.2 $Be^{2+} + I^- \rightarrow BeI_2$
   1.3 $Mg^{2+} + S^{2-} \rightarrow MgS$
   1.4 $Ba^{2+} + P^{3-} \rightarrow Ba_3P_2$
   1.5 $Ca^{2+} + O^{2-} \rightarrow CaO$

2. Halide: beryllium iodide; oxide: calcium oxide

3. Tin has a larger atomic radius.
   Tin's outermost electrons are in in n=5. Silicon's outermost electrons are in in n=3.
   The fifth energy level is further from the nucleus than the third energy level, making tin a larger atom.

4. Zinc has a smaller atomic radius.
   Both calcium and zinc's outermost electrons are in in n=4. However, calcium has 20 protons and zinc has 30 protons in its nucleus.
   Therefore, the electrostatic pull on zinc's electrons is higher, and this pulls them close to the nucleus, making the atomic radius smaller.

5. Silicon has a higher first ionisation energy.
   Tin's outermost electrons are in in n=5. Silicon's outermost electrons are in in n=3.
   The third energy level is closer to the nucleus than the fifth energy level, making it harder to remove an electron.

6. Zinc has a higher first ionisation energy.
   Both calcium and zinc's outermost electrons are in in n=4. However, calcium has 20 protons and zinc has 30 protons in its nucleus.
   Therefore, the electrostatic pull on zinc's electrons is higher, making it harder to remove an electron.