TOPIC 2.
THE STRUCTURE OF ATOMS

What distinguishes atoms of different elements?
An atom of any element is the smallest particle that still retains the properties of that element. The mass of a single atom of an element is in the range $10^{-24} \text{ g}$ to $10^{-21} \text{ g}$, depending upon the element chosen. As noted in Topic 1, each element has its own distinct type of atom, and what causes each type of atom to have its unique properties will now be considered.

Sub-atomic particles.
While the atom is the smallest unit of any element, atoms themselves consist of smaller particles. All atoms are found to contain three basic particles, viz. PROTONS, ELECTRONS and NEUTRONS. Protons have a positive electrical charge, electrons have a negative charge of identical magnitude to the proton, and neutrons carry no charge. The mass of a proton is slightly less than that of the neutron, while the mass of the electron is negligible compared with the proton mass. The following table gives the relative charge and mass for each particle.

<table>
<thead>
<tr>
<th>Particle Name</th>
<th>Mass</th>
<th>Relative mass</th>
<th>Relative Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>proton</td>
<td>$1.673 \times 10^{-24} \text{ g}$</td>
<td>1</td>
<td>+1</td>
</tr>
<tr>
<td>electron</td>
<td>$1/1836 \times$ proton mass = $0.0009 \times 10^{-24} \text{ g}$</td>
<td>1/1836</td>
<td>-1</td>
</tr>
<tr>
<td>neutron</td>
<td>$1.675 \times 10^{-24} \text{ g}$</td>
<td>1</td>
<td>0</td>
</tr>
</tbody>
</table>

Structure of atoms.
All atoms have the same structure consisting of a very small NUCLEUS of radius about $10^{-14} \text{ m}$ which contains all the protons and neutrons. The number of electrons is numerically equal to the number of protons and the electrons are in constant motion in the region of space outside the nucleus. The radius of an electron is less than $10^{-14} \text{ m}$ and, depending on the element involved, the average atom's radius is about $10^{-8} \text{ m}$. From these figures, it can be seen that the volume of the atom is mostly made up of empty space.

The number of protons in the nucleus of an atom of a given element is called the ATOMIC NUMBER of that element. Each element has its own unique atomic number. For example, the smallest atom is that of hydrogen which contains just 1 proton in its nucleus and 1 electron outside the nucleus. The next largest atom is that
of helium, atomic number = 2, which has two protons in its nucleus and 2 electrons outside the nucleus.

**It is the number of protons in the nucleus (the atomic number) that determines which element the atom represents.**

The following table lists the first 20 elements in atomic number order. (There is no value in memorising this table.)

<table>
<thead>
<tr>
<th>ATOMIC NUMBER</th>
<th>NAME</th>
<th>ATOMIC NUMBER</th>
<th>NAME</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>hydrogen</td>
<td>11</td>
<td>sodium</td>
</tr>
<tr>
<td>2</td>
<td>helium</td>
<td>12</td>
<td>magnesium</td>
</tr>
<tr>
<td>3</td>
<td>lithium</td>
<td>13</td>
<td>aluminium</td>
</tr>
<tr>
<td>4</td>
<td>beryllium</td>
<td>14</td>
<td>silicon</td>
</tr>
<tr>
<td>5</td>
<td>boron</td>
<td>15</td>
<td>phosphorus</td>
</tr>
<tr>
<td>6</td>
<td>carbon</td>
<td>16</td>
<td>sulfur</td>
</tr>
<tr>
<td>7</td>
<td>nitrogen</td>
<td>17</td>
<td>chlorine</td>
</tr>
<tr>
<td>8</td>
<td>oxygen</td>
<td>18</td>
<td>argon</td>
</tr>
<tr>
<td>9</td>
<td>fluorine</td>
<td>19</td>
<td>potassium</td>
</tr>
<tr>
<td>10</td>
<td>neon</td>
<td>20</td>
<td>calcium</td>
</tr>
</tbody>
</table>

**The nucleus.**

The nuclei of all atoms except those of hydrogen contain neutrons as well as protons, in about equal numbers. The role of neutrons is to provide stability to the nucleus of the atom. Most elements have atoms with varying numbers of neutrons present and each variation represents a particular **ISOTOPE** of that element. For example, carbon atoms always have 6 protons and most also have 6 neutrons in their nuclei, but about 1% of all carbon atoms occur as another isotope which has the required 6 protons (the characteristic which determines that they are atoms of carbon) and 7 neutrons.

The number of (neutrons + protons) in the nucleus of a given atom is called its **MASS NUMBER**. Sometimes it is useful to indicate the nuclear composition of a given atom by using the following notation:

$$^{a}^{z}X$$  
where $$a = \text{mass number} = (\text{number of protons} + \text{neutrons})$$  
and $$z = \text{atomic number} = \text{number of protons}$$.

From this notation, the number of neutrons present in the nucleus of a given atom = ($$a$$ - $$z$$). By this method, any specific nucleus type or **NUCLIDE** can be represented.

The number of neutrons present in the nucleus does not influence the chemical properties of the atoms, and the various isotopes of a given element behave identically
in all chemical reactions. When an isotope has significantly more or less neutrons than
the number of protons in its nucleus, it is usually unstable and undergoes a
NUCLEAR DECAY spontaneously over a period of time. Unstable nuclei are called
 RADIOACTIVE species.

Hydrogen has 3 isotopes which contain (1 proton + 0 neutrons), (1 proton + 1 neutron)
or (1 proton + 2 neutrons) respectively. Almost all hydrogen atoms are of the first
isotopic form. The second isotope, often called "deuterium" and given the special
symbol D, is present to the extent of 0.015% of all naturally occurring hydrogen nuclei,
while the third isotope ("tritium", T) is unstable and does not occur naturally, being
formed in nuclear reactions. These three isotopes of hydrogen can be represented as
$^1_1\text{H}$, $^2_1\text{H}$ and $^3_1\text{H}$ respectively or alternatively as $^1_1\text{H}$, $^2_1\text{D}$ and $^3_1\text{T}$.

It is a fundamental law of electrostatics that like electrical charges repel each other
while opposite charges attract. Further, these forces are proportional to $1/(d^2)$, where d
is the distance separating the charges. As the protons are extremely close to each other
in the nucleus and they all carry a +1 electrical charge, the electrostatic repulsion
between protons must be very strong and so there must be some other force operating
to overcome this repulsion in order to hold the nucleus together. This NUCLEAR
FORCE is one which operates only at very close distances and is much stronger than
the more familiar electrostatic force. The nuclear force operating between protons,
between neutrons and between protons and neutrons is the reason why the nucleus is
stable provided there is an appropriate ratio of protons to neutrons. It is a nuclear force
and is not experienced by electrons.

Electrons.
As atoms are electrically neutral and the charge on the electron is equal but opposite to
that on the proton, there must be identical numbers of electrons and protons in any
atom. The electrons are envisaged as being in rapid motion distributed around the
nucleus, but never actually being within the nucleus. Now the normal laws of
electrostatics would require the electrons to collapse into the nucleus due to the
attraction of the protons. Because this does not happen, there must be other laws
which govern the behaviour of electrons in atoms. Models to explain this will be
presented later in the year as part of all first year chemistry courses, but the results of
certain experimental evidence presented here is independent of those models.

Experiments show that the electrons occupy only certain ORBITS around the nucleus,
each orbit being characterised by its own associated energy and average distance from
the nucleus. This model of an atom was proposed by Neils Bohr in 1919. In this
model the orbits are grouped into ENERGY LEVELS or SHELLS and numbered 1, 2, 3,... outwards from the nucleus. The number of orbits available is strictly limited.
Electrons occupying orbits closest to the nucleus have the lowest energy while
electrons in orbits further out from the nucleus have higher energy. This is because, in order to overcome electrostatic attraction to the nucleus, energy must be supplied to an electron to move it from an orbit closer to the nucleus to an orbit further out. Electrons normally occupy the available orbits from the lowest energy upwards. Consequently, the larger the atomic number of an atom, the more electrons it will have and the larger will be its ATOMIC RADIUS. When all electrons in a given atom occupy the lowest possible energy orbits as just described, the atom is said to be in the GROUND STATE. If just the right amount of energy were supplied to an atom, one or more electrons can jump to occupy a higher energy orbit for a brief period and the atom is then in an EXCITED STATE. Excited electrons quickly collapse back to lower energy orbits in one or more steps and release the additional energy which was absorbed when they were promoted to the higher orbit. The energy released by excited electrons returning to lower energy orbits is in the form of ELECTROMAGNETIC RADIATION having wavelengths specific to the energy changes accompanying the transitions. This energy release often corresponds to the visible region of the electromagnetic spectrum in which case it can be detected when viewed through a prism as one or more coloured lines corresponding to light of specific wavelengths. The pattern of lines observed is the ATOMIC EMISSION SPECTRUM of that element and it can be used to identify it. The following diagram illustrates the various energy levels available to electrons in atoms, each orbit being shown as a circle around the nucleus.

The web site [http://chemistry.bd.psu.edu/jircitano/periodic4.html](http://chemistry.bd.psu.edu/jircitano/periodic4.html) presents illustrations of the atomic spectra of all the elements.
Check your understanding of this section.
Give the relative mass and charge properties of the sub-atomic particles: electrons, protons and neutrons.
What determines which element an atom represents?
What is the role of the neutron in the nucleus?
Identify the element represented by the symbol $^{14}_{6}X$ and give as much information as possible about the structure from the nuclide’s symbol.
How can an electron move from the ground state to an excited state?
What happens when an excited state electron falls back to a lower energy orbit?

Ground state electron arrangements.
Hydrogen has only 1 electron, and this will be in the energy level closest to the nucleus. The next element, helium with atomic number = 2 has 2 electrons, both of which can still occupy the lowest energy level.

![Diagram](image)
The arrangement of electrons in
a) hydrogen and b) helium

However, experiments show that no more than 2 electrons can use the lowest energy level orbit. Therefore the third electron in the next element, lithium (atomic number = 3), must occupy an orbit in the next highest energy level, located further away from the nucleus. This second shell can accommodate a maximum of 8 electrons, and for each of the subsequent elements through to neon, the additional electrons will be located in the n = 2 orbit. The following diagram represents the electron arrangements for the elements having atomic numbers 1 through to 10 - i.e. for hydrogen to neon.

![Diagram](image)
Ground state electron configurations of the first ten elements.
The element having an atomic number that is one more than neon’s is sodium which has atomic number = 11. The outermost electron in the sodium atom must occupy an orbit in the third shell. This is illustrated in the diagram below.

Similarly, the **GROUND STATE ELECTRONIC STRUCTURE** of each of the elements can be built up in this manner using orbits with the lowest available energy levels in each case. The diagram on the page II-8 illustrates the electron structures for the first 20 elements.

**Excited states and emission spectra**

For the simplest atom, hydrogen, there is just one electron which in the ground state occupies the lowest energy orbit, designated as n = 1. Various excited state orbits are available to the electron by its occupying the n = 2, n = 3, n = 4 etc orbits provided that the exact amount of energy required for the transition is supplied. Having attained an excited state which is typically a very short-lived condition, the electron can return back to the ground state either by a single jump whereby all the extra energy gained is lost in a single emission or alternatively, by a series of steps, each of which takes the electron to a lower energy level until the ground state is reached.

Depending upon the energy differences of the orbits occupied, energy will be emitted as infrared light, visible light or ultraviolet light. For example, electrons returning to the n = 2 orbit from orbits of higher energy than that of the n = 2 orbit emit visible light, giving rise to the red, blue, indigo or violet lines seen in the atom’s atomic spectrum. Likewise, all transitions to the n = 1 orbit (ground state) produce ultraviolet light at specific frequencies while transitions from higher energy orbits to the n = 3 orbit emit infrared light of specific frequencies. Infrared and ultraviolet light cannot be seen by eye but are easily detected by instruments designed specifically for the frequencies involved. This scheme is illustrated in the following diagram.
Formation of ions.
The elements whose outer electron shell is filled are helium, neon, argon, krypton, xenon and radon. Note that these are members of the eighth group of elements listed in Table 2 of the notes on Topic 1. The elements in this group are referred to as the INERT GASES or the NOBLE GASES. All the members of this group are characterised by being almost inert - i.e. they do not normally enter into any chemical reactions. For a chemical reaction to occur, electrons must be redistributed in some way between atoms of the reacting elements to form a CHEMICAL BOND. The reason for the lack of reactivity in this group of elements lies in the filled outer level which results in the outer electrons being subjected to the maximum attractive force from the nucleus. [See Topic XII page 9 for more details.] Thus atoms of noble gases require too much energy to be involved in any redistribution of electrons with other atoms. Sodium atoms have one more electron than the noble gas neon, and that electron is screened from the full attractive force of the protons by all the electrons between it and the nucleus. Thus only a relatively small amount of energy is required to cause that electron to be transferred to atoms of certain other elements. The loss of that outer electron from a sodium atom would result in there then being an excess +1 charge on the sodium atom.
The arrangement of electrons in the atoms of elements with atomic number 1 - 20.
When an atom acquires a charge and is no longer electrically neutral, it is called an **ION** and ions which have a positive charge are called **CATIONS**. The charge on the ion is shown by a superscript on the symbol for the corresponding atom. For example, in the case of sodium atoms forming sodium ions of charge +1, the process could be represented as

\[
\text{Na} \quad \text{→} \quad \text{Na}^+ \quad + \quad \text{electron}
\]

sodium atom \quad \text{sodium ion}

The Na atom and the Na\(^+\) ion could be diagrammatically represented as follows:

![Diagram of sodium atom and sodium ion](image)

The arrangement of electrons in a) a sodium atom, Na, b) a sodium ion, Na\(^+\), and c) a neon atom, Ne.

Note how the electron arrangement in the Na\(^+\) ion is the same as for the Ne atom which has atomic number 10, one smaller than the Na atom. All the other members of the first group of elements in Table 2 (the alkali metals) also have just one more electron than a noble gas atom, and they all behave as does sodium in that relatively little energy is needed to form their +1 cations Li\(^+\), Na\(^+\), K\(^+\), Rb\(^+\) and Cs\(^+\). Each of these cations has the same electron arrangement as the atom of the noble gas whose atomic number is 1 less than that of the first group element. Cations take the same name as the neutral element.

Some other elements have atoms which only require **one more** electron in order to obtain the noble gas arrangement. These atoms are F, Cl, Br and I, all of which are just one electron short of having a filled outer level. In chemical reactions these atoms frequently gain one electron from an atom of another element to form an ion which has an excess of one electron over the number of protons in the nucleus. For example, an atom of fluorine could be converted to an ion with a −1 charge. When an ion has a negative charge, it is called an **ANION**, the charge being written as a superscript as for cations. The members of the seventh group in Table 2 (called the "halogens") all form the −1 charged ions F\(^−\), Cl\(^−\), Br\(^−\) and I\(^−\) respectively. To show that they are in the form of the anion in a compound as distinct from the uncombined element, the ending "ide" is attached to a stem derived from the name of the element thus:

\[
\text{F}^− \quad \text{fluoride ion (instead of fluorine)} \quad \text{Cl}^− \quad \text{chloride ion (instead of chlorine)} \\
\text{Br}^− \quad \text{bromide ion (instead of bromine)} \quad \text{I}^− \quad \text{iodide ion (instead of iodine)}
\]

The formation of a chloride ion from a chlorine atom can be represented as

\[
\text{Cl} \quad + \quad \text{electron} \quad \rightarrow \quad \text{Cl}^−
\]

chlorine atom \quad \text{chloride ion}
The formation of cations from atoms is not limited to those elements which have only 1 electron more than a noble gas. The atoms of the elements in the second group of Table 2 all have two more electrons in their outer orbits than a noble gas. These atoms require only a relatively small amount of energy to enter into chemical reactions where the outer 2 electrons are redistributed to atoms of another element, forming cations with a 2+ charge such as Mg$^{2+}$ from Mg, Ca$^{2+}$ from Ca, etc. Likewise those elements whose atoms contain 3 more electrons in their outer orbits than a noble gas are often able to form 3+ ions in reactions. The third group in Table 2 all have 3 outer electrons and they may form ions such as Al$^{3+}$ from Al.

(The elements at the top of the second and third groups, beryllium and boron, do not form cations but can form compounds by using another method of bonding.)

Formation of cations with a 4+ charge requires so much energy that it rarely occurs. Similarly, the formation of anions from atoms of elements is not limited to only the halogens forming 1− ions. Atoms of the sixth group of elements in Table 2 starting with oxygen are all just two electrons short of the same electron arrangement as a noble gas, and these elements frequently form 2− anions in chemical reactions by gaining two electrons from atoms of other elements. Thus ions such as oxide (O$^{2−}$) from oxygen, sulfide (S$^{2−}$) from sulfur, selenide (Se$^{2−}$) from selenium and telluride (Te$^{2−}$) from tellurium can form in chemical reactions.

The gaining of three electrons to attain the noble gas electron arrangement is not common, but atoms from the fifth group of elements in Table 2 starting with nitrogen which are all 3 electrons short of the noble gas structure, can in some instances gain
the required 3 electrons to form \( 3^- \) anions. The nitride ion, \( N^3^- \) for example is formed in some reactions. Formation of anions with a 4– charge requires so much energy that it does not occur. When atoms of different elements attain the same electron arrangement by forming cations or anions, they are said to be ISOELECTRONIC.

From an examination of those elements which form cations and those which form anions in ionic compounds, using the data in Table 2 from Topic 1, it can be seen that **metals form cations and non-metals form anions**. This is a chemical property associated with metals and non-metals which will be further examined in later Topics.

<table>
<thead>
<tr>
<th>noble gas</th>
<th>noble gas</th>
<th>noble gas</th>
<th>noble gas</th>
<th>noble gas</th>
<th>noble gas</th>
<th>noble gas</th>
</tr>
</thead>
<tbody>
<tr>
<td>-3 e</td>
<td>-2 e</td>
<td>-1 e</td>
<td>+1 e</td>
<td>+2 e</td>
<td>+3 e</td>
<td></td>
</tr>
<tr>
<td>He</td>
<td>Li</td>
<td>Be</td>
<td>B</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>N</td>
<td>O</td>
<td>F</td>
<td>Ne</td>
<td>Na</td>
<td>Mg</td>
<td>Al</td>
</tr>
<tr>
<td>P</td>
<td>S</td>
<td>Cl</td>
<td>Ar</td>
<td>K</td>
<td>Ca</td>
<td>Ga</td>
</tr>
<tr>
<td>As</td>
<td>Se</td>
<td>Br</td>
<td>Kr</td>
<td>Rb</td>
<td>Sr</td>
<td>In</td>
</tr>
<tr>
<td>Sb</td>
<td>Te</td>
<td>I</td>
<td>Xe</td>
<td>Cs</td>
<td>Ba</td>
<td>Tl</td>
</tr>
</tbody>
</table>

**Test your understanding of this section.**

What is the maximum number of electrons that can occupy the \( n = 1 \) orbit?

How many can occupy the \( n = 2 \) orbit?

Why are the noble gases so unreactive?

In terms of their atomic structures, what is the difference between Na and Na⁺?

What does the term “isoelectronic” imply about two species?

Why is it unlikely that cations with a 4⁺ charge or anions with a 4⁻ charge would form?

From the name, how can one recognise that a non-metal such as chlorine is present in a compound rather than being present as the free element?

**Objectives of this Topic**

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Know the names and characteristics of the sub-atomic particles: electrons, protons and neutrons.
2. Know the general structural features of atoms.

3. Understand that the chemical properties of an element are determined by its atomic number which is the number of protons in the nucleus and this is also equal to the number of electrons outside the nucleus.

4. Understand the terms: atomic number; mass number; isotope; atomic radius; ground state; excited state.

5. Understand the concept of electron orbits and their associated energy levels.

6. Be able to draw simple electron orbit diagrams for the first 20 elements.

7. Recognise that the underlying reason why simple cations and anions can form is that in so doing, they attain the same electron arrangement as a noble gas and that this arrangement is particularly stable.

8. Be able to write formulas for cations and anions formed by elements of the first, second, third, sixth and seventh groups listed in Table 2 from Topic 1.

**SUMMARY**

Atoms consist of an extremely small, dense nucleus containing almost all of the mass in the form of protons carrying a positive electrical charge and neutrons which carry no charge. In constant motion around the nucleus are electrons, each of which bears an electrical charge of identical magnitude but opposite sign to that of the proton. The mass of an electron is negligible compared with that of a proton while protons and neutrons have about the same mass. Orbits used by electrons have very large radii compared with the diameter of the nucleus and consequently most of the volume of an atom is empty space.

Each element has its own characteristic number of protons in the nuclei of its atoms, called the atomic number. Atoms of elements are electrically neutral, so the atomic number of a given element is also the number of electrons in the atoms of that element. While the number of protons in the nucleus of a given element’s atoms is fixed, the number of neutrons usually varies, giving rise to different isotopes of that element. The various isotopes of a given element have identical chemical properties while their physical properties only differ in those associated with their differing masses.

The laws of electrostatics explain how attraction between opposite charges holds electrons in the atom through attraction to the nucleus, but those same laws require that
there would be extremely strong electrostatic repulsive forces between the protons in the nucleus. These repulsions are more than overcome by the existence of a much stronger attractive force called the nuclear force which only operates when particles are extremely close together. This force operates between protons and between neutrons and also between protons and neutrons, but is not experienced by electrons. The nuclear force allows stable nuclei to exist provided there are about the same number of protons and neutrons present. If this balance is not present, the nucleus can undergo a process of radioactive or nuclear decay, sometimes in a number of successive steps, until ultimately a stable nucleus is formed.

While electrostatic attraction explains how an electron is held to the nucleus of an atom, it does not explain why the electron does not spiral into the nucleus and self-destruct. Apparently the familiar laws of electrostatics observed in the macroscopic world do not rigidly apply in the sub-microscopic environment of electrons in atoms. Instead, experimental observations show that electrons can only occupy a relatively small number of specific orbits around the nucleus and that a unique energy is associated with an electron in any given orbit. The closer the orbit is to the nucleus, the lower is the energy of the electron which occupies that orbit. Electrons can only move between orbits by absorbing or emitting an amount of energy which corresponds exactly to the energy difference of the orbits occupied. Each orbit can hold only a fixed number of electrons. An atom is said to be in its ground state if all its electrons are in the lowest available energy orbits. An atom can briefly enter the short-lived excited state if one or more of its electrons absorb the required energy to move to a higher energy orbit but such electrons quickly emit some energy to return to the lower energy orbits once more. The emitted energy is in the form of electromagnetic radiation and is sometimes in the visible part of the spectrum so it can be detected by eye. Each element displays its own unique emission spectrum when excited electrons fall back to lower energy levels.

Elements of one group known as the noble gases generally do not enter into chemical reactions, a process which involves the redistribution of electrons in various ways between the reacting atoms. This particular stability of noble gases results from having a filled outer level of electrons in their atoms which endows them with a large effective nuclear charge. The effective nuclear charge of an atom determines how strongly electrons are held by the atom. Atoms of metals typically have only 1, 2 or 3 more electrons than a noble gas in their outer level and if supplied with a relatively small amount of energy, these outer electrons can be removed forming a positively charged cation from the previously neutral metal atom. These cations have the same number of electrons as a noble gas (but different number of protons) and are said to be isoelectronic with the noble gas. Cations are represented by the same symbol as their parent atom but with the charge shown as a superscript. Similarly, non-metal atoms that have 1, 2 or 3 less electrons in their outer level than a noble gas may gain the
required additional electrons in a reaction to form anions which are isoelectronic with that noble gas. To identify the presence of simple anions formed from elements in compounds, the ending “ide” is attached to a stem from the name of the element e.g. oxide from the element oxygen.

Energy requirements restrict the gain or loss of electrons to form cations or anions isoelectronic with the nearest noble gas to a maximum of three in most cases. A chemical property of a metal is the ability to form cations in reactions while a chemical property of a non-metal is the ability to form an anion in reactions.

**TUTORIAL QUESTIONS - TOPIC 2.**

The copy of the Periodic Table which is printed as a tear-out page at the end of these notes may be consulted as required when answering the questions in this Tutorial. The Periodic Table lists elements in order of increasing atomic number.

1. How many electrons would be in an atom of each of the following elements?
   
<table>
<thead>
<tr>
<th>(i)</th>
<th>(vii)</th>
<th>(xiii)</th>
</tr>
</thead>
<tbody>
<tr>
<td>nitrogen</td>
<td>B</td>
<td>copper</td>
</tr>
<tr>
<td>beryllium</td>
<td>Ne</td>
<td>potassium</td>
</tr>
<tr>
<td>sodium</td>
<td>Li</td>
<td>sulfur</td>
</tr>
<tr>
<td>silicon</td>
<td>Mg</td>
<td>oxygen</td>
</tr>
<tr>
<td>helium</td>
<td>Cl</td>
<td>fluorine</td>
</tr>
<tr>
<td>argon</td>
<td>As</td>
<td>bromine</td>
</tr>
</tbody>
</table>

2. How many protons would be in the nucleus of an atom of each of the following elements?

<table>
<thead>
<tr>
<th>(i)</th>
<th>(vii)</th>
<th>(xiii)</th>
</tr>
</thead>
<tbody>
<tr>
<td>carbon</td>
<td>Pb</td>
<td>oxygen</td>
</tr>
<tr>
<td>aluminium</td>
<td>He</td>
<td>helium</td>
</tr>
<tr>
<td>chlorine</td>
<td>Ag</td>
<td>iron</td>
</tr>
<tr>
<td>calcium</td>
<td>Sn</td>
<td>strontium</td>
</tr>
<tr>
<td>neon</td>
<td>N</td>
<td>argon</td>
</tr>
<tr>
<td>boron</td>
<td>Be</td>
<td>fluorine</td>
</tr>
</tbody>
</table>

3. Give the symbols for all those elements whose atoms

<table>
<thead>
<tr>
<th>(i)</th>
<th>(ii)</th>
<th>(iii)</th>
<th>(iv)</th>
<th>(v)</th>
<th>(vi)</th>
</tr>
</thead>
<tbody>
<tr>
<td>have the noble gas structure</td>
<td>have outer orbits which contain 1 electron more than a noble gas</td>
<td>have outer orbits which contain 2 electrons more than a noble gas</td>
<td>have outer orbits which are 1 electron short of the noble gas structure</td>
<td>have outer orbits which are 2 electrons short of the noble gas structure</td>
<td>have outer orbits which are 3 electrons short of the noble gas structure</td>
</tr>
</tbody>
</table>
4. Define the terms "excited state", "ground state", "isotope", "atomic number", "cation", "anion", "mass number".

5. Give the formulas for the following:
   (i) oxide ion
   (x) telluride ion
   (ii) magnesium ion
   (xii) potassium ion
   (iii) aluminium ion
   (xiii) barium ion
   (iv) iodide ion
   (xiv) selenide ion
   (v) rubidium ion
   (xv) lithium ion
   (vi) calcium ion
   (xvi) sodium ion
   (vii) sulfide ion
   (xvii) phosphide ion
   (viii) chloride ion
   (xviii) strontium ion
   (ix) nitride ion
   (xix) fluoride ion
   (x) caesium ion

6. Give the names of each of the following ions and the number of electrons present in each ion. The Periodic Table at the back of the book can be used to do this.
   (i) Ca$^{2+}$
   (ii) F$^-$
   (iii) S$^{2-}$
   (iv) Li$^+$
   (v) N$^{3-}$
   (vi) Cl$^-$
   (vii) O$^{2-}$
   (viii) Al$^{3+}$
   (ix) As$^{3-}$
   (x) Na$^+$
   (xi) Br$^-$
   (xii) I$^-$
   (xiii) Sr$^{2+}$
   (xiv) P$^{3-}$
   (xv) Se$^{2-}$
   (xvi) Ba$^{2+}$
   (xvii) K$^+$
   (xviii) Cs$^+$
   (xix) Rb$^-$
   (xx) Mg$^{2+}$
7. If necessary, consult Table 2 from Topic 1 to complete the following.

(a) Write the symbols of the elements and their cations for all of the elements in the first (alkali metals) and second (alkaline earth metals) groups.

(b) Write the symbols of the elements and their anions for all of the elements in the sixth and seventh groups.

8. From the information given, identify each of the following atoms or ions. Write the symbol for each species including the atomic number and mass number using the notation shown on page II-2.

(i) 8 protons + 8 neutrons + 8 electrons
(ii) 8 protons + 8 neutrons + 10 electrons
(iii) 6 protons + 8 neutrons + 6 electrons
(iv) 17 protons + 18 neutrons + 18 electrons
(v) 1 proton + 1 neutron + 1 electron
(vi) 37 protons + 48 neutrons + 36 electrons
(vii) 20 protons + 20 neutrons + 18 electrons
(viii) 26 protons + 30 neutrons + 23 electrons
(ix) 36 protons + 48 neutrons + 36 electrons
(x) 35 protons + 44 neutrons + 35 electrons
(xi) 35 protons + 46 neutrons + 36 electrons
(xii) 7 protons + 7 neutrons + 10 electrons
(xiii) 15 protons + 16 neutrons + 18 electrons
(xiv) 13 protons + 14 neutrons + 10 electrons
(xv) 26 protons + 30 neutrons + 24 electrons
(xvi) 82 protons + 126 neutrons + 82 electrons
(xvii) 4 protons + 5 neutrons + 4 electrons
(xviii) 38 protons + 50 neutrons + 36 electrons
(xix) 19 protons + 20 neutrons + 19 electrons
(xx) 16 protons + 16 neutrons + 18 electrons

9. Write the formulas for six ions which are isoelectronic with the neon atom.

10. Explain the chemical distinction between "chlorine" and "chloride". What does the statement that a certain compound contains chlorine really mean?

11. What is the origin of the yellow light emitted from sodium vapour street lamps?
ANSWERS TO TUTORIAL TOPIC 2

1. In an atom (as distinct from an ion) the number of electrons is identical to the number of protons in the nucleus, the atomic number of that element. To answer this question, locate the element in the Periodic Table given at the end of this book and obtain its atomic number.

   (i) 7  (vii) 5  (xiii) 29
   (ii) 4  (viii) 10  (xiv) 19
   (iii) 11  (ix) 3  (xv) 16
   (iv) 14  (x) 12  (xvi) 8
   (v) 2  (xi) 17  (xvii) 9
   (vi) 18  (xii) 33  (xviii) 35

2. The number of protons in the nucleus of an atom is equal to the atomic number of that element which can be obtained by consulting the Periodic Table provided.

   (i) 6  (vii) 82  (xiii) 8
   (ii) 13  (viii) 2  (xiv) 2
   (iii) 17  (ix) 47  (xv) 26
   (iv) 20  (x) 50  (xvi) 38
   (v) 10  (xi) 7  (xvii) 18
   (vi) 5  (xii) 4  (xviii) 9

3. The easiest way to answer this question is to identify in which Periodic Table Group each element is located. Then for each position after a noble gas, the atoms of that Group will have an additional electron compared with the preceding noble gas atoms. Alternatively, for each position before a noble gas, the atoms of that element will have one fewer electron than the noble gas atoms. Thus all members of the first Group will have one electron more than the preceding noble gas in their atoms, the atoms of the second Group will have two more electrons than the preceding noble gas, etc. The atoms of all members of the halogens (and hydrogen) will be one electron short of the number in the following noble gas atoms, those of the sixth Group will be two electrons short, and so on.

   (i) He, Ne, Ar, Kr, Xe  (ii) Li, Na, K, Rb, Cs
   (iii) Be, Mg, Ca, Sr, Ba  (iv) H, F, Cl, Br, I
   (v) O, S, Se, Te  (vi) N, P, As, Sb, Bi
4. **Excited state:** when one or more electrons in an atom have received exactly the energy require to be promoted to a higher energy orbit further from the nucleus; only a short-lived condition.  
**Ground state:** when all the electrons in an atom occupy the lowest energy orbits available.  
**Isotope:** a specification of both the number of protons and neutrons present in a given atom of an element. The various isotopes of a given element differ in the number of neutrons in their nuclei.  
**Atomic number:** the number of protons in the nucleus of an atom. Atomic number is unique to the atoms of each particular element.  
**Cation:** an atom (or group of atoms) that bear a positive electrical charge.  
**Anion:** an atom (or group of atoms) that bear a negative electrical charge.  
**Mass number:** the number of (protons plus neutrons) present in the nucleus of an atom. Different isotopes of a given element must have different mass numbers but all have the same atomic number.

5. Simple cations take the same name as the element from which they are derived - thus “magnesium” is the name for the element and also the name for the cation, Mg\(^{2+}\), present in compounds. Recognition that it is a cation is made by the fact that it will be followed by the name of whichever anion accompanies the magnesium cation in the compound - e.g. the name “magnesium oxide” indicates a compound of magnesium and oxygen. For anions, it is necessary to attach an ending to a stem from the element. If the anion is just that from a single element (as distinct from the polyatomic anions dealt with in a subsequent Topic), the ending “ide” is attached to the stem. Thus “oxide” is the anion derived from the oxygen atom, “chloride” from the chlorine atom, etc, and hence the name “magnesium oxide” for the compound of formula MgO rather than “magnesium oxygen” which would be ambiguous.

To deduce the charge on a given cation or anion, it is essential to know to which group the element belongs. Using Table 2 of Topic 1, for elements in the first group, the alkali metals, the cation will always have a \( +1 \) charge, in the second group, the alkali earth metals, it will have a \( +2 \) charge, etc. If there are two or more possible charges that an element can take in forming a cation, Roman numerals are used in conjunction with the name - e.g. “copper(II)” for Cu\(^{2+}\) to distinguish it from the possible Cu\(^{+}\) cation. For anions of elements that need 1, 2 or 3 more electrons to achieve the noble gas configuration, i.e. the seventh, sixth and fifth groups in Table 2 of Topic 1, the anion charge will be \(-1\), \(-2\) and \(-3\) respectively.
(i) \(O^{2-}\)  
(ii) \(Mg^{2+}\)  
(iii) \(Al^{3+}\)  
(iv) \(I^-\)  
(v) \(Rb^+\)  
(vi) \(Ca^{2+}\)  
(vii) \(S^{2-}\)  
(viii) \(Cl^-\)  
(ix) \(N^{3-}\)  
(x) \(Cs^+\)  
(xi) \(Te^{2-}\)  
(xii) \(K^+\)  
(xiii) \(Ba^{2+}\)  
(xiv) \(Se^{2-}\)  
(xv) \(Li^+\)  
(xvi) \(Na^+\)  
(xvii) \(P^{3-}\)  
(xviii) \(Sr^{2+}\)  
(xix) \(F^-\)  
(xx) \(Br^-\)

6. To deduce the number of electrons present in a cation, **subtract** the magnitude of the charge from the number of electrons in the neutral atom. For example, a calcium atom has 20 electrons. In the \(Ca^{2+}\) cation, the atom has lost 2 electrons but still has the 20 protons in its nucleus (and hence the resulting 2+ charge on the cation), so there must be \(20 - 2 = 18\) electrons left in the cation.

For anions, **add** the excess negative charge to the number of electrons in the neutral atom. For example, the \(F^-\) ion has the 9 electrons of a neutral F atom plus an additional electron (= 10 electrons in total) that bestows the \(-1\) charge on the anion.

(i) calcium ion; 18 electrons  
(ii) fluoride ion; 10 electrons  
(iii) sulfide ion; 18 electrons  
(iv) lithium ion; 2 electrons  
(v) nitride ion; 10 electrons  
(vi) chloride ion; 18 electrons  
(vii) oxide ion; 10 electrons  
(viii) aluminium ion; 10 electrons  
(ix) arsenide ion; 36 electrons  
(x) sodium ion; 10 electrons  
(xi) bromide ion; 36 electrons  
(xii) iodide ion; 54 electrons  
(xiii) strontium ion; 36 electrons  
(xiv) phosphide ion; 18 electrons  
(xv) selenide ion; 36 electrons  
(xvi) barium ion; 54 electrons  
(xvii) potassium ion; 18 electrons  
(xviii) caesium ion; 54 electrons  
(xix) rubidium ion; 36 electrons  
(xx) magnesium ion; 10 electrons
8. First identify the element from the number of protons as this is also its atomic number. Then obtain the mass number of the particular isotope by adding the number of protons and neutrons. Write the atomic number as a subscript and the mass number as a superscript, both ahead of the symbol for the element. Finally, to deduce if the species is a cation or anion, check that the number of electrons equals the number of protons. If there are more electrons than protons, the species is an anion and the excess negative charge shown as a superscript after the symbol. If there are more protons than electrons, the species is a cation and again, the excess charge is shown as a superscript. For example, in (ii), the number of protons = 8 so the element is oxygen, symbol O. The mass number = 8 protons + 8 neutrons = 16. Finally, there are 10 electrons but only 8 protons in the nucleus so the species is in fact an ion with a 2 – charge and the full description would be $^{16}_{8}\text{O}^{2-}$.

(i) $^{16}_{8}\text{O}$  (ii) $^{16}_{8}\text{O}^{2-}$  (iii) $^{14}_{6}\text{C}$  (iv) $^{35}_{17}\text{Cl}^{-}$  
(v) $^{2}_{1}\text{H}$  (vi) $^{85}_{37}\text{Rb}^{+}$  (vii) $^{40}_{20}\text{Ca}^{2+}$  (viii) $^{56}_{26}\text{Fe}^{3+}$  
(ix) $^{84}_{36}\text{Kr}$  (x) $^{79}_{35}\text{Br}^{-}$  (xi) $^{81}_{35}\text{Br}^{-}$  (xii) $^{14}_{7}\text{N}^{-}$  
(xiii) $^{31}_{15}\text{P}^{3-}$  (xiv) $^{27}_{13}\text{Al}^{3+}$  (xv) $^{56}_{26}\text{Fe}^{2+}$  (xvi) $^{208}_{82}\text{Pb}$  
(xvii) $^{9}_{4}\text{Be}$  (xviii) $^{88}_{38}\text{Sr}^{2+}$  (xix) $^{39}_{19}\text{K}$  (xx) $^{32}_{16}\text{S}^{2-}$

9. Isoelectronic species have the same electron arrangement but differ in the number of protons in their nuclei - i.e. they represent different elements. Here the following ions all have the same number of electrons as the Ne atom (10 electrons) located as 2 in the first orbit and 8 in the second, but they all have different numbers of protons.

F$^{-}$, O$^{2-}$, N$^{3-}$, Na$^{+}$, Mg$^{2+}$, Al$^{3+}$

10. "Chlorine" refers to the free element, molecular Cl$_2$, while "chloride" indicates that the Cl atoms are combined with other atoms in a compound, possibly as Cl$^{-}$ ions, and this is the true meaning of the statement.
11. In a sodium lamp, sodium metal is vaporized by heating to form free Na atoms as a gas and some of those atoms at any instant receive enough energy to promote electrons to higher energy level orbits, i.e. to the excited state. When the electrons return to the ground state, they emit energy which corresponds to the energy difference between the orbits used, and in some cases this energy is released as light from the visible part of the spectrum corresponding to yellow light.

See the web site [http://chemistry.bd.psu.edu/jircitano/periodic4.html](http://chemistry.bd.psu.edu/jircitano/periodic4.html) for the full atomic emission spectrum of sodium.
SOME RELATED ACTIVITIES

1. Demonstrating electrostatic attraction and repulsion.
While it is easy to show that like poles of magnets repel and unlike poles attract, the corresponding interactions between electrical charges are less familiar. The following simple experiments show that like electrical charges repel and opposite electrical charges attract.

(a) Take two lengths of sticky tape about 15 cm long and stick them firmly down on a table so that about 2 cm hangs over the end. In each hand, grasp the overhanging ends and simultaneously pull the tapes quickly upwards. Allowing the tapes to dangle down, bring the two sticky sides together and observe that they repel each other.
Next, run one of the tapes through two of your other fingers and again bring the tapes together. This time they attract each other.

Explanation: Pulling the tapes off the table gave them both a positive electrical charge because electrons were left behind on the table's surface. When brought close together, the tapes repelled each other because they both carried the same charge. Running one tape through your fingers transferred some electrons from your hand to that tape - in fact excess electrons, causing that tape to acquire a negative charge. Bringing the tapes together again caused them to attract as they now bear opposite electrical charges. This is the principle that Gladwrap uses, an electrostatic charge being obtained by the action of pulling it off the roll. As metals are good conductors of electricity, Gladwrap does not work on metal containers as the charge on the surface of the Gladwrap is immediately dispersed through the metal.

(b) Take a piece of paper and tear a few very small pieces off it. Rub a plastic ruler vigorously on a cloth and then hold the ruler a little distance above the pieces of paper. Note that the paper spontaneously "jumps" up to stick to the ruler.

Explanation: Depending on what type of cloth is used, rubbing the ruler with the cloth can either remove electrons from the ruler, leaving it with an excess positive charge, or alternatively, electrons may be rubbed off the cloth onto the ruler, thereby leaving the ruler with an excess negative charge. If the ruler carries an excess positive charge, when it is held above the paper, electrons are attracted from the ground and onto the paper which thereby acquires an excess negative charge. As the paper is not anchored to the ground and it has an electrical charge which is opposite to that of the ruler, it is attracted strongly enough to the ruler to overcome gravitational attraction and "fly" up to the ruler where it is held by electrostatic attraction.
Alternatively, if the ruler has an excess negative charge as a result of being rubbed on the cloth, the effect of bringing it near the paper is to drive electrons away from the paper to the ground, thereby leaving the paper with an excess positive charge. Again, the paper will be attracted to the ruler because it and the ruler have opposite electrical charges on their surfaces.
2. How do fireworks work?

Residents of Sydney have become accustomed to spectacular and expensive displays of fireworks. Having completed Topic 2, you can now appreciate some of the chemistry involved.

Fireworks usually contain gunpowder which is ignited for example by a fuse or by electrically heating it. The rapidly burning gunpowder evolves a lot of heat energy which causes particles of metal elements such as magnesium and iron that have been mixed in with the gunpowder to burn and release more heat, increasing the temperature even further. When these metal particles become very hot, they give off white light - this phenomenon is called *incandescence*, and is the basis for the operation of tungsten filament light globes. This is the source of the bright white light seen in the exploding fireworks.

To obtain the range of colours in fireworks, use is made of the *atomic spectrum* emitted by atoms of various elements such as the red colour from strontium atoms, yellow from sodium atoms, green from barium atoms and blue from copper atoms. These atoms are present as cations in compounds (usually their chlorides) but at the elevated temperature in the fireworks they are converted to the metal atoms by gaining electrons from the anion in the salt. The compounds are mixed in with the gunpowder and the metal particles. The high temperature causes some of the electrons in the metal atoms formed to become excited and jump to higher energy orbits by *absorbing energy*. However, these excited electrons quickly lose the extra energy gained and fall back to lower energy orbits, *emitting energy* as light of a specific colour which depends on the energy difference between the orbits which the electron uses.

Note the sequence of events: to become excited, the electrons must gain energy from some source (the burning gunpowder and metal particles); the excited state is short lived and the electron quickly gives out energy in the form of coloured light. The process continues for whatever period of time that there is sufficient heat supplied to excite the electrons in the metal atoms formed by the combustion reaction.

When you next see a display of fireworks, contemplate the chemistry of what is happening.

**Recommended follow up chemical module:**
Section: Properties of Atoms
Module: Atomic and Nuclear Structure
Topics covered: *Fundamental concepts; nuclear equations; electromagnetic radiation; wave-particle duality; Bohr model.*
FOR THOSE WHO WANT TO KNOW MORE.

Sub-shell electronic structure of atoms

The information given in Topic 2 presents the Bohr model of the atom. It describes the orbits which are available for an atom’s electrons to occupy. Evidence for the Bohr model comes from the atomic emission spectra of the elements and there is good agreement between experimental data and the model for the hydrogen atom. However for all other elements the emission spectra are much more complicated than can be explained by the Bohr model. In addition the shape of the Periodic Table suggests a more complex structure of electron energy levels than the 2, 8, 18 proposed by Bohr. This additional information is intended to extend the Bohr model so as to be a more accurate representation of the electron structure of atoms. A more extensive treatment is to be found in Appendix 3 at the end of this book.

Electron orbits and sub-shells.

The Bohr model proposes that electrons are moving in circular orbits at a fixed distance from the nucleus. The orbits are labeled \( n = 1, n = 2, n = 3 \) etc where the \( n = 1 \) orbit is the closest to the nucleus and the most stable (lowest energy). In a more accurate description, the concept of major energy levels is retained. Here we call them electron shells, and also label them \( n = 1, n = 2, n = 3 \) etc. However the new model allows for sub-shells within each of these shells. The larger the value of \( n \), the more sub-shells are available within the shell. For example the second shell \((n = 2)\) contains two sub-shells; the third shell \((n = 3)\) contains three sub-shells and so on. The sub-shells are labeled \( s, p, d \) and \( f \).
In a hydrogen atom the sub-shells of a given \( n \) all have the same energy. In all atoms other than hydrogen, the sub-shells have slightly different energy to one another, increasing in the order \( s < p < d < f \).

The \( s \) sub-shell can hold a maximum of 2 electrons; the \( p \) sub-shell a maximum of 6 electrons and the \( d \) sub-shell a maximum of 10 electrons. This is reflected in the shape of the modern periodic table. The periodic table contains blocks of elements (the ‘\( s \)-block’, the ‘\( p \)-block’, the ‘\( d \)-block’) representing the type of sub-shell the outermost electrons of the element occupy. The width of the block mirrors the maximum number of electrons that can be accommodated in the corresponding sub-shell. So the \( s \)-block is 2 elements wide, the \( p \)-block is 6 elements wide and the \( d \)-block is 10 elements wide.

The slightly different energy of the sub-shells, coupled with the observation that the shells (\( n = 1, n = 2, n = 3 \) etc) become closer together in energy as they become further from the nucleus, results in an unexpected order of filling of the third and fourth shell. The \( s \) sub-shell of the fourth shell is filled before the \( d \) sub-shell of the third shell. This can be illustrated by looking at the elements of atomic numbers 17-21.