Matter and material: The atom

Objectives

- Understand the concept of relative atomic mass.
- Describe the structure of an atom and understand atomic number and atomic mass.
- Determine the number of protons, electrons and neutrons in elements.
- Determine the number of electrons and charge on an ion after removing/adding electrons to an atom.
- Define isotopes.
- Calculate the relative atomic mass of naturally occurring elements from the percentage of each isotope in a sample.
- Represent atoms and ions using the notation E, Z and A.
- Describe an atomic energy level, shells and orbitals.
- Use Aufbau diagrams and spectroscopic notation to represent the electron configuration of elements 1-20.
- Understand the difference between core electrons, valence electrons and valency.

Atomic diameter, mass and relative atomic mass

Atoms have extremely small diameters. The diameter of an atom ranges from 1×10^{-10} m (0.0000000001 m) to 4×10^{-10} m (0.0000000004 m). To put 1×10^{-10} m in perspective: divide a one-meter line into ten billion equal parts – one of these parts is equivalent to the diameter of the smallest atom, which is helium. Atoms also have extremely small masses. The table below shows the mass of hydrogen, carbon, oxygen and uranium in grams (g).

Element	Average* atomic mass (g)
Hydrogen	1.673×10^{-24}
Carbon	1.99×10^{-23}
Oxygen	2.66×10^{-23}
Uranium	3.95×10^{-22}

* You will understand later why we refer to the AVERAGE atomic mass when we learn about isotopes.

This means that the atomic mass of hydrogen, the lightest element, is:
0.0000000000000000000000001673 g.

Uranium is the heaviest natural element with an atomic mass of:
0.000000000000000000000395 g.

Video on the small size of the atom: **https://www.youtube.com/watch?v=yQP4UJhNn0I&t=30s**

These are extremely small numbers that would be difficult to work with. We therefore use a relative atomic mass. What does that mean?

Imagine creating a baby scale. A 3 kg baby receives a value of 1 unit on the baby scale.

- If someone has a mass of 6 kg, their mass on the baby scale will be 6kg/3kg = 2 units on the baby scale
- If someone has a mass of 52 kg, their mass on the baby scale will be 57kg/3kg = 19 units on the baby scale
- If someone has a mass of 1.5 kg, their mass on the baby scale will be 1.5kg/3kg = 0.5 units on the baby scale

In the same way, scientists have decided that a carbon-12 atom (that is a carbon atom that contains 6 protons and 6 neutrons), which has a mass of 1.99×10^{-23} g, has a value of 12 units on the atomic scale.

Therefore, **1 unit on the atomic scale** = 1.99×10^{-23} g / 12 = $\mathbf{1.658 \times 10^{-24}}$ **g**.

This is approximately the mass of one proton, or the mass of one neutron, or the mass of one hydrogen atom that is made up of one proton and one electron.

This means that 1.658 x 10^{-24} g of an atom receives a value of 1 unit on the atomic scale.

- If hydrogen has a mass of 1.673 x 10^{-24} g, its mass on the atomic scale will be: 1.673 x 10^{-24}g/1.658 x 10^{-24}g = 1.01 units on the atomic scale
- If oxygen has a mass of 2.66 x 10^{-23} g, its mass on the atomic scale will be: 2.66 x 10^{-23}g/1.658 x 10^{-24}g = 16.04 units on the atomic scale
- If uranium has a mass of 3.95 x 10^{-22} g, its mass on the atomic scale will be: 3.95 x 10^{-22}g/1.658 x 10^{-24}g = 238.24 units on the atomic scale

Do these numbers look familiar?
Often, these numbers are rounded off in a periodic table as the atomic mass of an atom (represented under the symbol of the atom in a periodic table). We say that a hydrogen atom has an atomic mass of 1; an oxygen atom has an atomic mass of 16; and a uranium atom has an atomic mass of 238. All that we need to understand from this is that when we refer to the mass of an atom, we are not referring to the real mass of the atom in grams, but rather we to a relative mass on a very small atomic scale, which makes the numbers easier to work with.

Structure of the atom

Protons, neutrons and electrons

Particle physicists have identified subatomic particles that make up atoms, which include protons, neutrons and electrons. The table below shows the masses (in grams) and charges (in coulombs) of these particles:

Subatomic particle	Mass (g)	Charge (C)
Electron	9.109 x 10^{-28}	-1.6 x 10^{-19}
Proton	1.673 x 10^{-24}	+1.6 x 10^{-19}
Neutron	1.673 x 10^{-24}	No charge

Let’s note a few important points about the masses of subatomic particles:

- The mass of a proton and the mass of a neutron is the same.
- The mass of an electron is much smaller than the mass of a proton or neutron. (To be precise, the mass of an electron is 1836 times smaller than the mass of a proton or neutron. This is like comparing the mass of a 50 kg human to the mass of a 0.027g earthworm. If you stood on a scale and then stood on a scale holding an earthworm, this would not make a significant difference to your mass). Therefore, we say that the mass of an electron is negligible.
- The mass of a proton is the same as the mass of a hydrogen atom. This is because a hydrogen atom is made up of 1 proton (as well as one electron, but the mass of an electron is negligible).
- On the atomic scale, the mass of a proton = 1 unit (equivalent to a hydrogen atom); the mass of a neutron = 1 unit and the mass of an electron = 0 units.

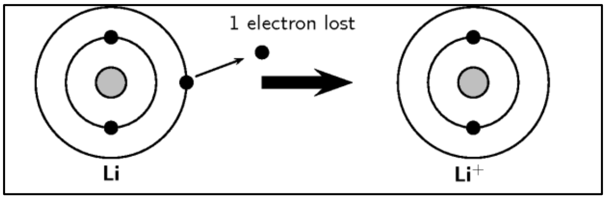
Let’s note a few important points about the charges of subatomic particles:

- Neutrons are neutral, meaning that they have no charge. They only contribute to the mass of an atom.
- Protons have a positive charge.
- Electrons have a negative charge.
- The value of the charge of a proton equals the value of the charge of an electron.
- An atom is neutral if the number of positively charged protons equals the number of negatively charged electrons.

Protons and neutrons are both found in the centre of the atom inside a very small and dense nucleus. Around the nucleus, occupying a relatively large region, are the electrons. It’s very important to note that it’s the **proton** that gives an atom its identity. For example, any atom that contains three protons, regardless of the number of neutrons or electrons it has, is a lithium atom. Any atom that contains nine protons, regardless of the number of neutrons or electrons it has, is a fluorine atom.

It’s also very important to note that the nucleus remains intact, and the number of protons (and neutrons) in each atom will never change. On the other hand, electrons may be added or removed from an atom.

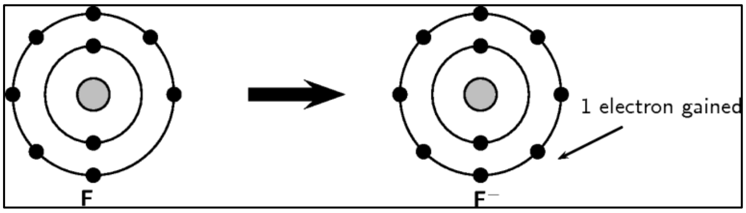
Consider the diagram below representing a lithium atom:



A neutral lithium atom has 3 protons and 3 electrons.

- If an electron is **removed** from a lithium atom, it will still have 3 protons, but only 2 electrons.
- Therefore the positive charge will be higher than the negative charge.
- The atom will have a net charge of **+1**.
- It is an **ion** because it is an atom with a charge.
- A positively charged ion is called a **cation**.

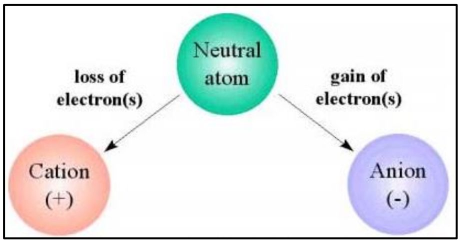
Consider the diagram below representing a fluorine atom:



A neutral fluorine atom has 9 protons and 9 electrons.

- If an electron is **added** to the fluorine atom, it will still have 9 protons, but it will now have 10 electrons.
- Therefore the positive charge will be lower than the negative charge.
- The atom will have a net charge of **-1**.
- It is an **ion** because it is an atom with a charge.
- A negatively charged ion is called an **anion**.

In summary, any charged atom is called an ion. An atom that loses electron/s becomes positively charged and is called a cation. An atom that gains electron/s becomes negatively charged and is called an anion:

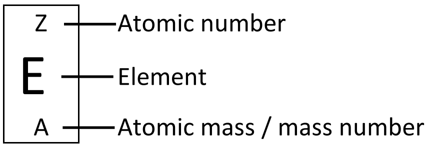


Atomic number and atomic mass

Each element has a particular number of protons. As mentioned above, the number of protons in the nucleus of an element is what differentiates it from another element. The number of protons in the nucleus can never change. In the periodic table:

- The number above the element is known as the **atomic number**. It may also be represented by the letter **Z**. The atomic number represents the number of **protons**.
- The number below the element is known as the **atomic mass**. It may also be represented by the letter **A**. The atomic mass represents the **mass of the entire atom**.
- **A** may also be referred to as the **mass number**, which represents **the total number of protons plus neutrons** in an atom. This is because a proton and a neutron each have an atomic mass of 1.

In the periodic table, each block is represented as follows:



Where E represents the element, Z is the number above the element which represents the atomic number, and A is the number below the element which represents the atomic mass.

We may calculate the number of neutrons in an atom as follows:
= Atomic mass (A) – atomic number (Z)

For example, in the Al atom represented below, there are 13 protons. How many neutrons exist in an atom of aluminium?

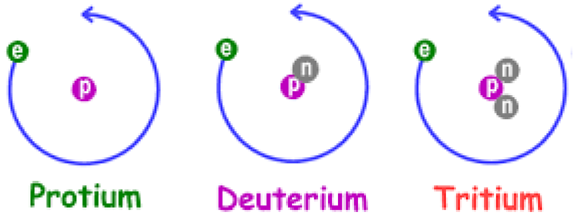
13
Al
27

Number of neutrons = Atomic mass (A) – atomic number (Z) = 27 – 13 = 14 neutrons

Isotopes

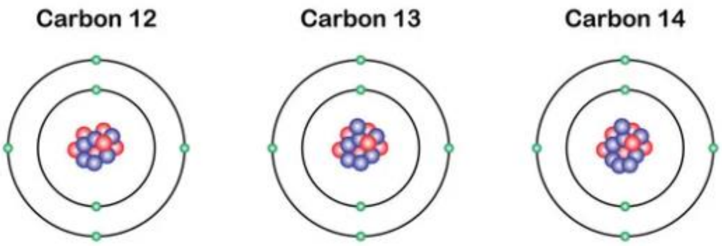
Atoms of the same element always have the same number of protons. However, atoms of the same element may have a different number of **neutrons**. This means that they differ in their **atomic mass**. These are known as different isotopes of the element. Isotopes *D are atoms of the same element with the same number of protons but a different number of neutrons.

All elements exist in nature as a mixture of isotopes, ranging from 2-36 isotopes. This is just how they have existed since the beginning of time, and how they will continue to exist. An isotope cannot change to form a different isotope. For example, these are the three different isotopes of hydrogen found in nature:



All three of the above isotopes represent hydrogen atoms since they each contain one proton. However:
- Protium has no neutrons in its nucleus and therefore has an atomic mass of 1 unit.
- Deuterium has one proton and one neutron in its nucleus and therefore has an atomic mass of 2 units.
- Tritium has one proton and two neutrons in its nucleus and therefore it has an atomic mass of 3 units.

Carbon also occurs naturally in three isotopes:



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

All three of the above isotopes represent an atom of carbon because they each contain six protons. However:

- Carbon 12 has six protons and six neutrons in its nucleus and therefore has an atomic mass of 12 units.
- Carbon 13 has six protons and seven neutrons in its nucleus and therefore has an atomic mass of 13 units.
- Carbon 14 has six protons and eight neutrons in its nucleus and therefore has an atomic mass of 14 units.

Hydrogen is the only element whose isotopes have unique names. Isotopes of other elements are simply named according to their **atomic mass**. For example:

- Carbon 12 or C-12 has an atomic mass of 12 units. Since all carbon atoms have six protons, this isotope of carbon has six neutrons (since 12 – 6 = 6).
- Carbon 14 or C-14 has an atomic mass of 14 units. Since all carbon atoms have six protons, this isotope of carbon has eight neutrons (since 14 – 6 = 8).
- Lithium 7 or Li-7 has an atomic mass of 7 units. Since all lithium atoms have three protons, this isotope of lithium has four neutrons (since 7 – 3 = 4).
- Aluminium-26 or Al-26 has an atomic mass of 26 units. Since all aluminium atoms have 13 protons, this isotope of lithium has 13 neutrons (since 26 – 13 = 13).

Isotopes are represented by isotopic symbols as follows:

$$^{A}_{Z}\mathrm{E}$$

Therefore, C-12 would be written as:

$$^{12}_{6}\mathrm{C}$$

C-14 would be written as:

$$^{12}_{6}\mathrm{C}$$

Li-7 would be written as:

$$^{7}_{3}\mathrm{Li}$$

Al-26 would be written as:

$$^{26}_{13}\mathrm{Al}$$

Symbols for isotopes are often simply written as ^{12}C, ^{14}C, ^{7}Li or ^{26}Al because for a particular element, the atomic number is always the same, and we may always find the atomic number in the periodic table.

In some periodic tables, the atomic mass is a number with decimal places rather than a whole number. This is because the atomic mass averages out the isotopes of each element. Scientists gather naturally occurring samples of an element and identify the percentage occurrence of each isotope of that element. They can then calculate the relative atomic mass of that element. For example, naturally occurring bromine consists of ^{79}Br and ^{81}Br, having relative atomic masses of 50.686% and 49.314%, respectively. Calculate the relative atomic mass (Ar) of bromine.

$$\begin{aligned} Ar &= \left(\frac{50{,}686}{100} \times 79\right) + \left(\frac{49{,}314}{100} \times 81\right) \\ &= 40{,}04194 + 39{,}94434 \\ &= 79{,}99 \end{aligned}$$

Therefore, the relative atomic mass of Bromine is 79.99 units. This represents the relative average of all the different isotopes of bromine atoms present in nature.

Here are some important things to note about the calculation above:

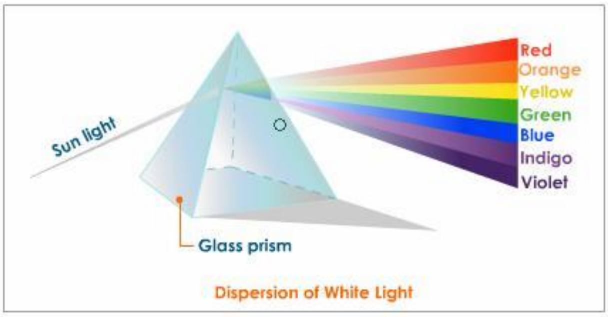
- When you provide your final Ar answer, leave out a unit of measurement.
- Make sure you represent the percentage as $\frac{number}{100}$.
- Multiply the percentage by the **atomic mass** of the element.

Electron configuration

Now that we have discussed the protons and neutrons that make up the nucleus of an atom, we will consider the arrangement of the electrons in an atom.

<u>Light spectrum and visible light</u>

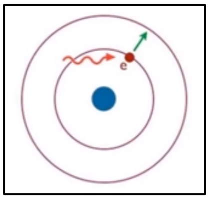
When a beam of white light is refracted as it passes through a triangular prism, it results in a continuous spectrum of all the colours of the rainbow (white light is a combination of all the colours of the rainbow). It is called a continuous spectrum because it consists of all the colours of the rainbow in their order.



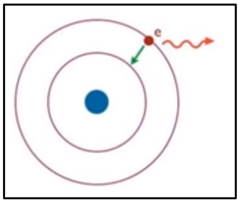
The light spectrum occurs because different frequencies of light travel slower than others through glass. Violet light travels the fastest while red light travels the slowest. Upon entry of white light into the prism, there will be a slight separation of the white light into the colours of the rainbow. Upon exiting the triangular prism, the separation becomes even greater. The different coloured lights bend differently if the prism surfaces on which light is going in and coming out are not parallel.

Video on light prism: **https://www.youtube.com/watch?v=--b1F6jUx44**

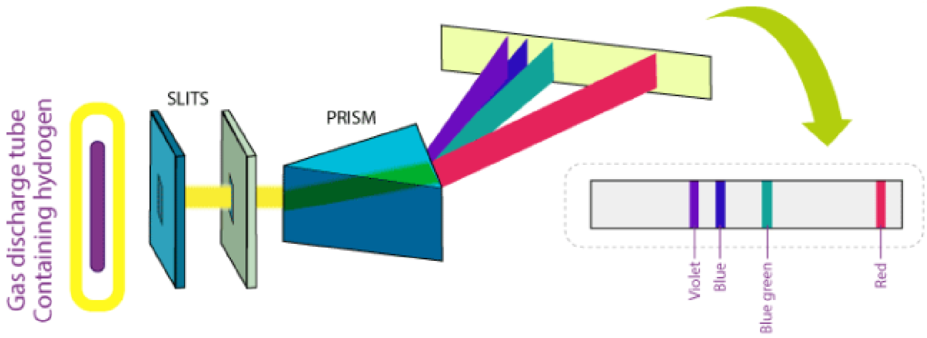
In 1913, Danish scientist, Niels Bohr, produced a model that suggests that electrons around an atom moved only in particular energy levels. Each energy level has a fixed amount of energy. When electrons absorb energy, they move to a higher energy level.



When electrons lose energy, they move back down the energy levels and release the energy in the form of <u>light</u> energy.



Because each element has a different set of electrons in a different number of energy levels, specific colours of light are given out by each element. For example, when an electric current is passed through hydrogen gas, <u>pink light</u> is emitted. When this light is passed through a triangular prism, it will separate into specific wavelengths of light that are part of hydrogen's line emission spectrum (non-continuous spectrum), including violet, blue, blue-green and red. This corresponds to the fixed amount of energy lost as the electrons go down the energy levels.

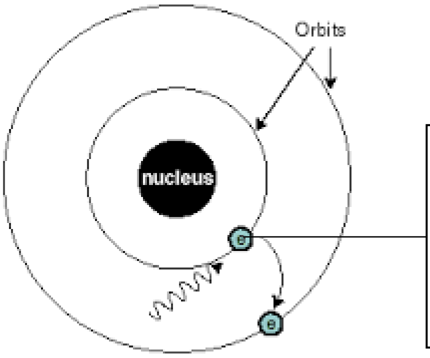


A flame test is a procedure used in chemistry to identify the presence of certain elements based on the characteristic flame colour and flame emission spectrum. For example, when group 1 metals react with oxygen, they burn with brightly coloured flames to form metal oxides. Lithium burns in oxygen to create a red flame; sodium burns in oxygen to create an orange flame; and potassium burns in oxygen to create a lilac flame. These colours help us to identify which elements are present.

The key thing to understand here is that when electrons gain energy, they get ‘excited’ and jump onto a higher energy level that is further away from the nucleus. When that electron loses that energy and moves back down to its own energy level, it releases that energy in the form of light energy. Each element has its own number of electrons and energy levels so each element will emit its own characteristic colour of light.

Energy quantisation

When Bohr conducted his experiments on hydrogen, he suggested an explanation for the line spectrum of hydrogen. He proposed that the energy of electrons in hydrogen atoms is **quantised**. This means that an electron needs to gain a **fixed amount of energy** to jump onto the next energy level. He suggested that the electron in a hydrogen atom can only move from one energy level to the next by absorbing or emitting fixed amounts of **quanta** of energy. Quanta are tiny but distinct quantities or pockets of energy (the singular of quanta is quantum).



Take this electron as an example. It needs 13.1 eV (electron volts, a small unit of energy) to jump onto the next energy level. If it only absorbs 13 eV, it won’t be able jump to the next energy level. The **quantum** of energy that this electron requires is 13.1 eV.

Video on electrons moving up and down energy level: **https://www.youtube.com/watch?v=uyBwV0o4l8w** up to 2:04

Energy levels

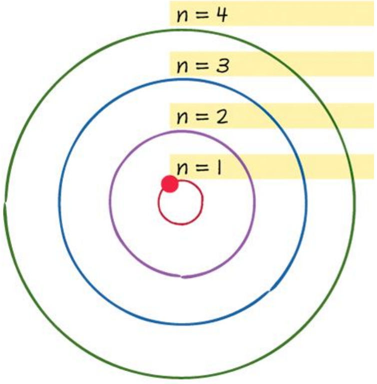
What are energy levels? The electrons surrounding an atom are found in regions around the nucleus called energy levels. An energy level represents the 3-dimensional space surrounding the nucleus where electrons are most likely to be found. Electrons in different energy levels have differing amounts of energy.

The lowest energy level, closest to the nucleus of the atom, has a quantum number of n=1, the next energy level has a quantum number of n=2 and so (the highest energy level that exists is n=7). The first energy level is closest to the nucleus. It can hold fewer electrons, and these electrons have lower energies than electrons in higher energy levels:

- n=1 can hold a maximum of 2 electrons.
- n=2 can hold a maximum of 8 electrons.
- n=3 can hold a maximum of 8 electrons.
- n=4 can hold a maximum of 18 electrons.
- n=5 can hold a maximum of 18 electrons.
- n=6 can hold a maximum of 32 electrons.
- n=7 can hold a maximum of 32 electrons.

If you add up all these numbers, what number do you derive, and what does it correspond to?
2 + 8 + 8 + 18 + 18 + 32 + 32 = 118, which corresponds to the total number of elements in the periodic table.

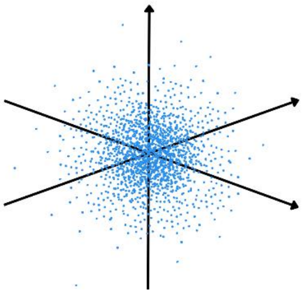
We will only consider the first four energy levels as can be seen in the diagram below, but remember that there are 7 energy levels:



Each energy level is made up of a shell or shells of different shapes. Those shapes correspond to areas of high probability of finding an electron. We will only consider the s and p shells, although d and f shells also exist.

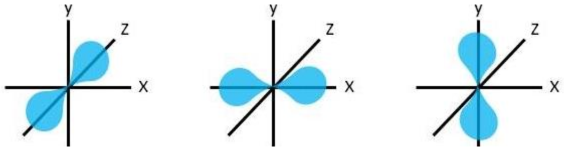
An s shell is spherical and looks like this:

What this shell and its shape mean is that an electron will have a high probability of existing within this shape, with a higher probability of its position in space depicted by the most concentrated dots, and a lower probability of its position in space being depicted by the less concentrated dots:

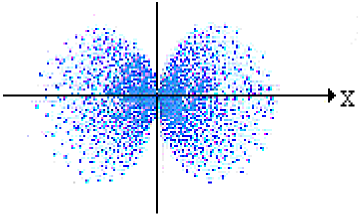


An s shell has one electron orbital and each electron orbital can hold a maximum of 2 electrons. Therefore, an s shell holds a maximum of 2 electrons.

A p shell comes in a set of three, each with its own orientation, and has the shape of an infinity sign like this:



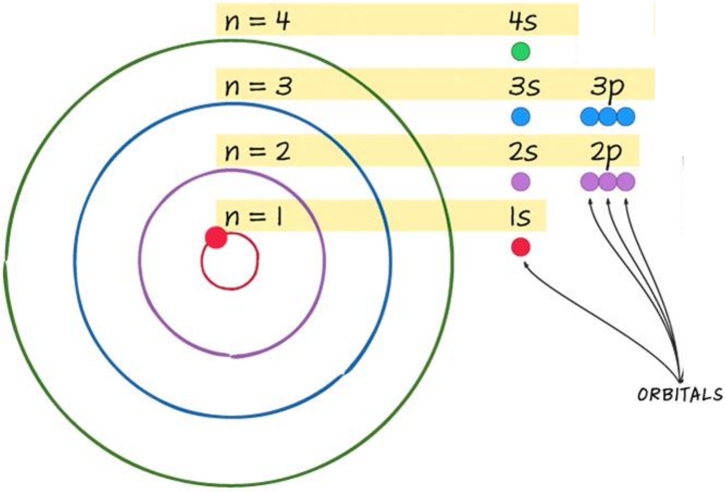
What this shell and its shape mean is that an electron will have a high probability of existing within this shape, with a higher probability of its position in space depicted by the most concentrated dots, and a lower probability of its position in space being depicted by the less concentrated dots:



Each p orientation has one electron orbital and each electron orbital can hold a maximum of 2 electrons. Therefore, a p shell holds a maximum of 6 electrons.

Each energy level has specific shells associated with it:

- n=1 has a 1s shell. The 1 refers to energy level n=1, and the s refers to the s shell. It can hold a maximum of 2 electrons. Therefore, this energy level can hold a maximum of 2 electrons in total.
- n=2 has a 2s shell and a 2p shell. The 2 refers to energy level n=2, the s refers to the s shell, and the p refers to the p shell. The s shell can hold a maximum of 2 electrons. The p shell can hold a maximum of 6 electrons. Therefore, this energy level can hold a maximum of 8 electrons in total.
- n=3 has a 3s shell and a 3p shell. The 3 refers to energy level n=3, the s refers to the s shell, and the p refers to the p shell. The s shell can hold a maximum of 2 electrons. The p shell can hold a maximum of 6 electrons. Therefore, this energy level can hold a maximum of 8 electrons in total.
- n=4 has a 4s shell. The 4 refers to energy level n=4, and the s refers to the s shell, which can hold 2 electrons. n=4 also has a 4p shell which can hold 6 electrons and a 4d shell which can hold 10 electrons. n=4 can therefore hold a maximum of 18 electrons, but we will only consider the 4s shell of n=4.

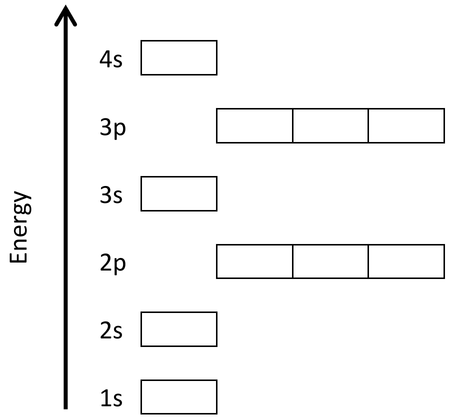


Video on shell shapes: **https://www.youtube.com/watch?v=sMt5Dcex0kg**

Electron configuration

The arrangement of electrons around the nucleus of an atom is called the electron configuration of an atom. This applies to atoms in their ground state when electrons occupy their lowest possible energy levels (as opposed to gaining energy and being in the 'excited' state). We will only consider the electron configurations of elements 1 (hydrogen) to 20 (calcium).

To represent the electron configuration of an element in its ground state we use the Aufbau principle. Aufbau means 'building up' in German. This principle states that electrons occupy atomic orbitals in order of increasing energy. We need to know the energy order of the orbitals and how the electrons fill up these orbitals. Let's start off with a generic setup of an Aufbau diagram. Always set up your Aufbau diagram exactly like in the diagram below:

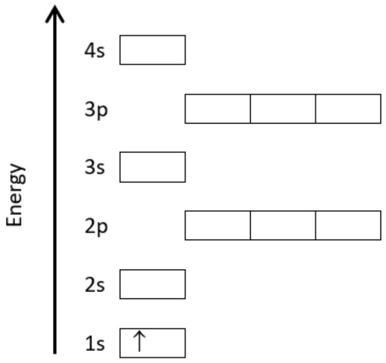


We also need to keep in mind that:

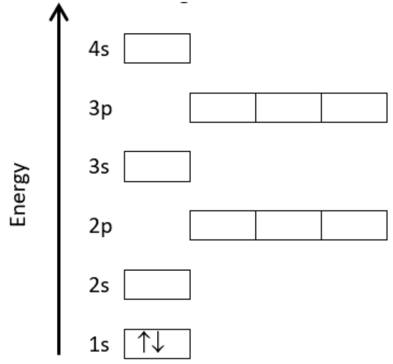
- The Pauli Exclusion Principle tells us that an orbital can hold a maximum of two electrons that have opposite spins to minimise repulsion. Recall that electrons are negatively charged, and negative charges repel other negative charges.
- Hund's Rule states that electrons fill the orbitals of a shell with their spins in the same direction before pairing to minimise repulsion.

When we draw an Aufbau diagram, we use the arrow directions to represent the opposite spin of the electrons. It doesn't matter if you start with an up facing arrow or the down facing arrow, as long as you keep it consistent in your diagram.

This is the Aufbau diagram of hydrogen:

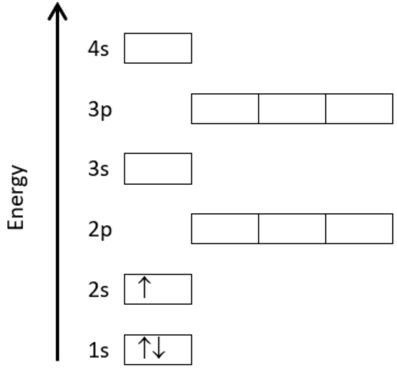


This is the Aufbau diagram of helium:

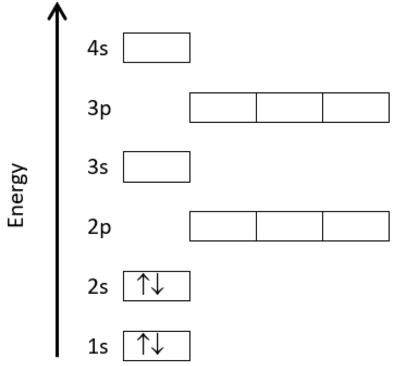


Note that we are following Pauli Exclusion Principle that tells us that an orbital can hold a maximum of two electrons that have opposite spin.

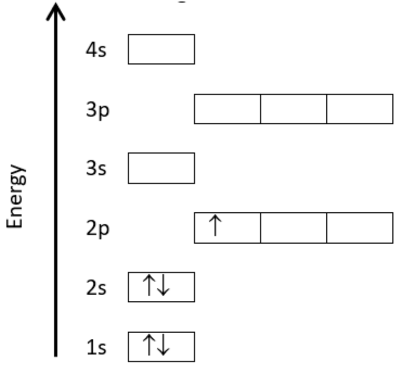
This is the Aufbau diagram of lithium:



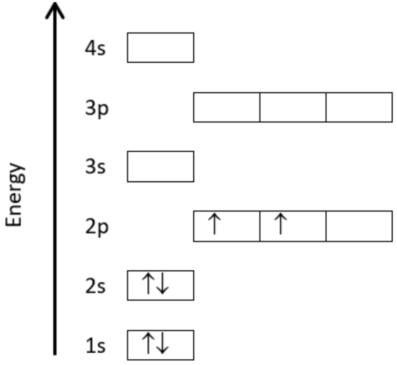
This is the Aufbau diagram of beryllium:



This is the Aufbau diagram of boron:

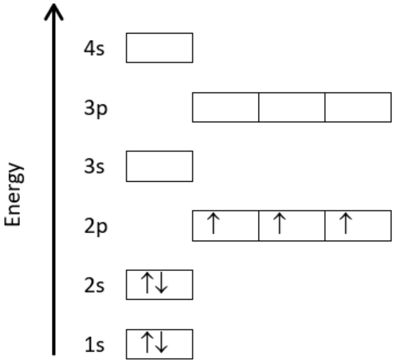


This is the Aufbau diagram of carbon:

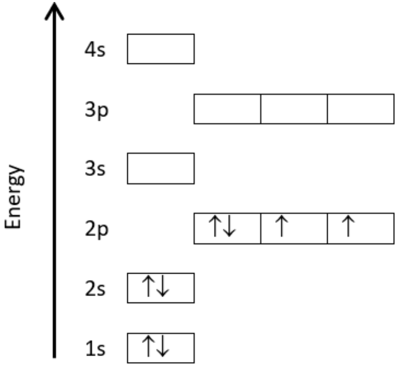


Note that we are following Hund's Rule that states that electrons fill the orbitals of a shell with their spins in the same direction before pairing

This is the Aufbau diagram of nitrogen:



This is the Aufbau diagram of oxygen:



And so on…

We may also represent the electron configuration of an atom using spectroscopic notation or sp notation as follows:

- Hydrogen has one electron in n=1 within the s shell. Therefore its sp notation is **$1s^1$**.
- Helium has two electrons in n=1 within the s shell. Therefore its sp notation is **$1s^2$**.
  Helium has a full energy level since n=1 only has an s shell, which can hold a maximum of two electrons. Helium is a noble gas.
- Lithium has two electrons in n=1 within the s shell. Lithium's third electron is in the s shell of n=2. Therefore its sp notation is **$1s^22s^1$**.
- Beryllium has two electrons in n=1 within the s shell. Beryllium's third and fourth electrons are in the s shell of n=2. Therefore its sp notation is **$1s^22s^2$**.
- Boron has two electrons in n=1 within the s shell; its third and fourth electrons are in the s shell of n=2; and its fifth electron is in the p shell of n=2. Therefore its sp notation is **$1s^22s^22p^1$**.
- Carbon has two electrons in n=1 within the s shell; its third and fourth electrons are in the s shell of n=2; and its fifth and sixth electrons are in the p shell of n=2. Therefore its sp notation is **$1s^22s^22p^2$**.
- Nitrogen has two electrons in n=1 oxygen; its third and fourth electrons are in the s shell of n=2; and its fifth, sixth and seventh electrons are in the p shell of n=2. Therefore its sp notation is **$1s^22s^22p^3$**.
- Oxygen has two electrons in n=1 oxygen; its third and fourth electrons are in the s shell of n=2; and its fifth, sixth, seventh and eighth electrons are in the p shell of n=2. Therefore its sp notation is **$1s^22s^22p^4$**.

Electrons in completely filled energy levels are called **core** electrons. Core electrons are unreactive. This means that they do not get involved in chemical bonding. Electrons in the outermost energy level of an atom are called **valence** electrons. Valence electrons are the electrons that can be gained, lost or shared in bonding with other atoms.

Noble gases are unreactive because their outermost energy level is full. Having a filled outermost energy level give the atom a lower energy state. All atoms tend towards their lowest energy state. This applies to natural systems in general. For example, a ball will roll down the hill to a lower state of potential energy. In the same way, atoms will tend towards a state of lower potential energy. Atoms that have valence electrons create chemical bonds to fill their outermost energy level and be in a lower energy state.

Note that there is a difference between **valence electrons** and **valency**:

- Valence electrons ^{D} are the number of electrons in the outermost energy level of an atom.
- Valency ^{D} is the number of electrons an atom needs to gain or lose to have a full outermost energy level. We will discuss this in more detail in the next section.

Practice questions and answers

Practice questions

1. In a laboratory, two samples of iodine exist. One sample is the stable isotope of iodine. The other sample is the radioactive isotope of iodine, I-131. Stable iodine is absorbed and used by the thyroid gland in the neck to regulate metabolism, growth, and development. I-131 is used in the treatment of thyroid cancer since this isotope of iodine will be absorbed into the thyroid and kill the cancer cells.

1.1 How many protons and neutrons does I-131 contain?
1.2 The stable isotope of iodine has 74 neutrons. Represent this isotope using isotopic notation.
1.3 If the laboratory sample has 82,3% of the stable iodine isotope and 17,7% of I-131, calculate the relative atomic mass of iodine in that sample.

2. Provide the sp notation and Aufbau diagram of the following elements:

2.1 Fluorine
2.2 Sodium
2.3 Chlorine
2.4 Calcium

3. How many core electrons and how many valence electrons are found in each of the following elements, and in which energy levels are they found?

3.1 Fluorine
3.2 Sodium
3.3 Chlorine
3.4 Calcium
3.5 Helium
3.6 Neon

Practice question answers

1.

1.1 53 protons (atomic number = number of protons)
78 neutrons (atomic mass – atomic number = 131 – 53 = 78)

1.2

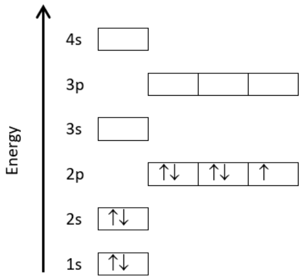
$$^{127}_{53}\mathrm{I}$$

1.3 The atomic mass of iodine with 75 neutrons: 75 + 53 = 128 ∴ 20% I-128

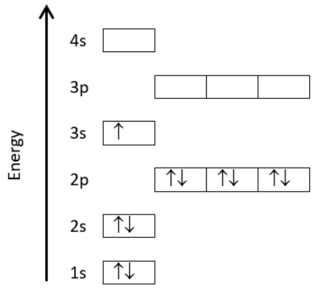
$$Ar = \left(\frac{82{,}3}{100} \times 127\right) + \left(\frac{17{,}7}{100} \times 131\right)$$
$$= 104{,}521 + 23{,}187$$
$$= 127{,}71$$

2.

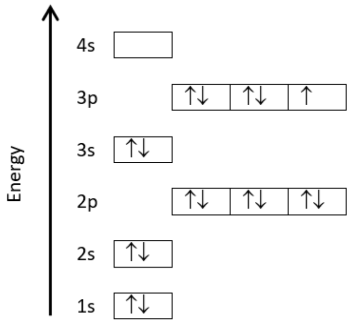
2.1 Fluorine: $1s^2 2s^2 2p^5$



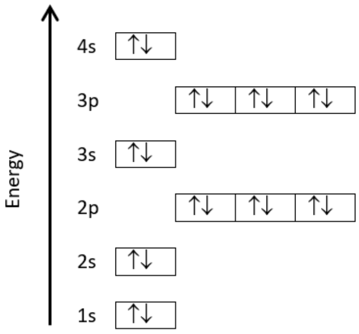
2.2 Sodium: $1s^2 2s^2 2p^6 3s^1$



2.3 Chlorine: $1s^2 2s^2 2p^6 3s^2 3p^5$



2.4 Calcium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$



3.

3.1 Fluorine: 2 core electrons in n=1; 7 valence electrons in n=2
3.2 Sodium: 2 core electrons in n=1; 8 core electrons in n=2; 1 valence electron in n=3
3.3 Chlorine: 2 core electrons in n=1; 8 core electrons in n=2; 7 valence electron in n=3
3.4 Calcium: 2 core electrons in n=1; 8 core electrons in n=2; 8 core electrons in n=3; 2 valence electron in n=4
3.5 Helium: 2 valence electrons in n=1 (considered a full valence energy level)
3.6 Neon: 2 core electrons in n=1; 8 valence electrons in n=2 (considered a full valence energy level)