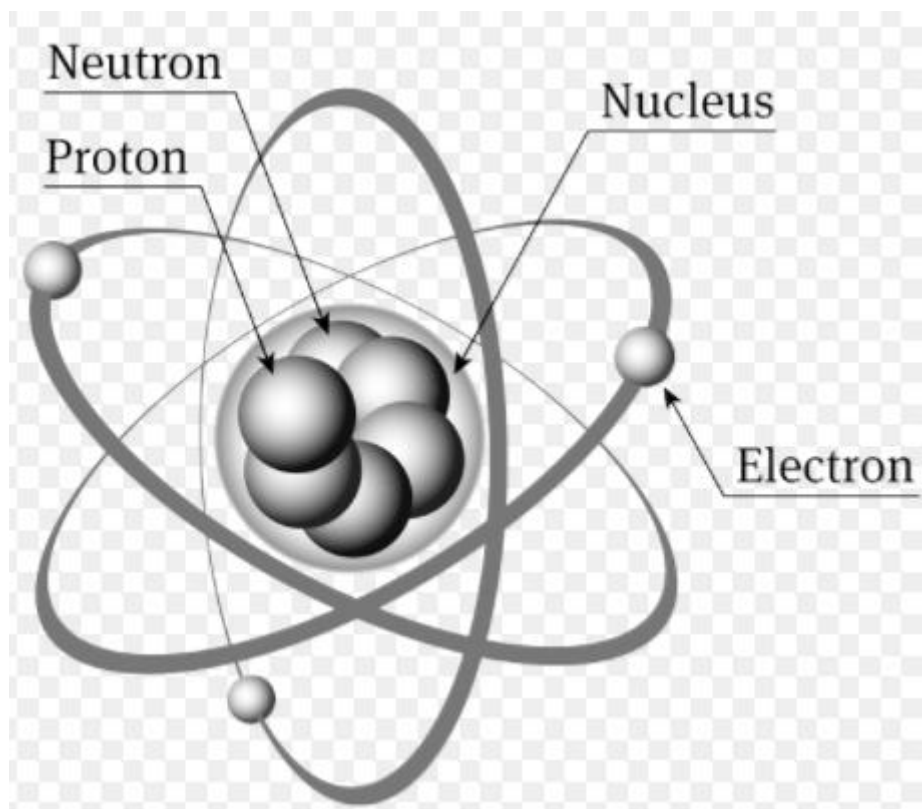


UNIT 1

ATOMS AND THE PERIODIC TABLE

Student Version



Source: www.byjus.com

Contents

- a) Atoms and Ions
- b) Electrons, Orbitals and Shells
- c) Trends in the Periodic Table
- d) The Development of the Atomic Model

Key words: atom, proton, neutron, electron, nucleon, isotope, coulomb, atomic mass unit, relative atomic mass, ion, energy level, orbital, sub-level, shell, Pauli Exclusion Principle, Aufbau Principle, Hund's Rule of Maximum Multiplicity, Electronic Configuration, Periodic Law, Ionisation Energy, Electron Affinity, Dalton, Thomson, Rutherford

Units which must be completed before this unit can be attempted:

None – this unit should be studied first

Estimated teaching time: 10 hours

The Periodic Table of the Elements

	1	2	3	4	5	6	7	8	9	10	11	12	(13)	(14)	(15)	(16)	(17)	(18)	
	6.9 Li lithium 3	9.0 Be beryllium 4		47.9 Ti titanium 22	50.9 V vanadium 23	52.0 Cr chromium 24	54.9 Mn manganese 25	55.8 Fe iron 26	58.9 Co cobalt 27	58.7 Ni nickel 28	63.5 Cu copper 29	65.4 Zn zinc 30	10.8 B boron 5	12.0 C carbon 6	14.0 N nitrogen 7	16.0 O oxygen 8	19.0 F fluorine 9	4.0 He helium 2	
	23.0 Na sodium 11	24.3 Mg magnesium 12		91.2 Zr zirconium 40	92.9 Nb niobium 41	96.0 Mo molybdenum 42	[98] Tc technetium 43	101.1 Ru ruthenium 44	102.9 Rh rhodium 45	106.4 Pd palladium 46	107.9 Ag silver 47	112.4 Cd cadmium 48	27.0 Al aluminium 13	28.1 Si silicon 14	31.0 P phosphorus 15	32.1 S sulfur 16	35.5 Cl chlorine 17	39.9 Ar argon 18	
	39.1 K potassium 19	40.1 Ca calcium 20		88.9 Y yttrium 39	88.9 Sc scandium 21	138.9 La * lanthanum 57	137.3 Ba barium 56	178.5 Hf hafnium 72	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	200.6 Hg mercury 80	114.8 In indium 49	118.7 Sn tin 50	121.8 Sb antimony 51	127.6 Te tellurium 52	126.9 I iodine 53	131.3 Xe xenon 54	
	132.9 Cs caesium 55	137.3 Ba barium 56		178.5 Hf hafnium 72	180.9 Ta tantalum 73	183.8 W tungsten 74	186.2 Re rhenium 75	190.2 Os osmium 76	192.2 Ir iridium 77	195.1 Pt platinum 78	197.0 Au gold 79	200.6 Hg mercury 80	204.4 Tl thallium 81	207.2 Pb lead 82	209.0 Bi bismuth 83	[209] Po polonium 84	[210] At astatine 85	[222] Rn radon 86	
	[223] Fr francium 87	[226] Ra radium 88		[267] Rf rutherfordium 104	[268] Db dubnium 105	[271] Sg seaborgium 106	[272] Bh bohrium 107	[270] Hs hassium 108	[276] Mt meitnerium 109	[281] Ds darmstadtium 110	[280] Rg roentgenium 111	Elements with atomic numbers 112-116 have been reported but not fully authenticated							
	140.1 Ce cerium 58	140.9 Pr praseodymium 59		144.2 Nd neodymium 60	150.4 Sm samarium 62	[145] Pm promethium 61	152.0 Eu europium 63	157.3 Gd gadolinium 64	158.9 Tb terbium 65	162.5 Dy dysprosium 66	164.9 Ho holmium 67	167.3 Er erbium 68	168.9 Tm thulium 69	173.1 Yb ytterbium 70	175.0 Lu lutetium 71				
	232.0 Th thorium 90	231.0 Pa protactinium 91		238.0 U uranium 92	[244] Pu plutonium 94	[237] Np neptunium 93	[243] Am americium 95	[247] Cm curium 96	[247] Bk berkelium 97	[251] Cf californium 98	[252] Es einsteinium 99	[257] Fm fermium 100	[258] Md mendelevium 101	[259] No nobelium 102	[262] Lr lawrencium 103				

1.0 H hydrogen 1

Key
relative atomic mass
symbol
name
atomic (proton) number

* 58 – 71 Lanthanides

† 90 – 103 Actinides

a) Atoms and Ions

(i) What are atoms?

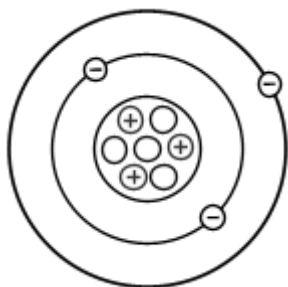
- Atoms are the building blocks of all matter; everything in the universe is made of atoms
- Atoms are very, very small, much too small to see, even under a microscope; if you line 100 billion atoms up in a row, it would only be 1 cm long
- In one small piece of chalk there are around 10,000,000,000,000,000,000 atoms
- Because atoms are too small to see, for a long time nobody knew what they looked like; different scientists in history have proposed different models showing what atoms look like, and some scientists carried out experiments to try to prove their models or disprove other people's models; our understanding of atoms has therefore changed over time, just like our understanding of many things in science

(ii) What do we know about atoms today?

- Until 150 years ago, scientists thought that an atom was the smallest particle which existed; but since then, scientists have discovered three even smaller particles which exist inside an atom
- The three particles found inside an atom are **protons**, **electrons** and **neutrons**
- Protons have a positive charge, electrons have a negative charge and neutrons have no charge
- Protons and neutrons have the almost the same mass (1 unit), but electrons have (almost) no mass
- Most of an atom is in fact empty space; in the centre of this space is a nucleus, containing the protons and neutrons; the electrons travel around the nucleus in shells (like orbits), similar to the way in which planets orbit the sun
- Protons and neutrons are collectively called **nucleons**



Test Your Knowledge 1.1: Understanding the basic structure of an atom



Copy this diagram of a simple atom and label it with the following terms: Proton, neutron, electron, nucleus, shell.

Then copy and complete the following table:

Particle	Relative Charge	Relative Mass	Location
Proton			
Neutron			
Electron			

- The nucleus is very small; if an atom were the size of a football pitch, the nucleus would be the size of a drawing pin; for this reason, most diagrams of atoms are not drawn to scale
- Atoms are held together because the negatively charged electrons are attracted to the positively charged protons in the nucleus

- Inside the nucleus, the protons repel each other, and the neutrons act as a glue to stop the protons flying apart
- The basic properties of the three fundamental particles in an atom can be summarized in the following table:

Particle	Charge	Mass	Location
Proton	+1 unit	Approx 1 unit	Nucleus
Neutron	No charge	Approx 1 unit	Nucleus
Electron	-1 unit	(almost) zero	Shells

(iii) How do we give names and symbols to atoms?

- An atom is named after the number of protons in its nucleus; if the nucleus of an atom has 1 proton, it is hydrogen; if it has two protons, it is helium; if it has 3, it is lithium etc; the number of protons in the nucleus of an atom is called the **atomic number**; it has the symbol Z

The **atomic number** is the number of protons in the nucleus of an atom

- The Periodic Table shows the name and symbol given to every atom, based on the number of protons it has
- Atoms do not have a charge, either positive or negative; in each atom, therefore, the number of protons and electrons must be the same, so the positive and negative charges cancel out



Test Your Knowledge 1.2: Understanding atomic names and symbols

Copy and complete the following table, using the Periodic Table to help you

Number of protons	Number of electrons	Name of atom	Atomic symbol
1			
3			
8			
15			
29			
		Xenon	
		Fluorine	
		Zinc	
			Al
			Na

Lesson 2 – How many neutrons do atoms have?

(iv) How many neutrons do atoms have?

- It is not possible to predict the number of neutrons in an atom from the atomic number alone; this is because atoms with the same number of protons (ie the same atomic number) can have different numbers of neutrons

The sum of the number of protons and neutrons in the nucleus of an atom is called its **mass number**; it is represented by the symbol A

The **mass number** is the sum of the number of protons and neutrons in the nucleus of an atom

- In order to describe an atom completely, it is necessary to show the name or symbol of the atom (which gives the number of protons and electrons), and the mass number of the atom (from which the number of neutrons can be deduced):
 - using symbols, the mass number of the atom is shown as a superscript just before the symbol:
eg ${}^9\text{Be}$
 - using names, the mass number of the atom is shown after the name, separated by a hyphen:
eg **beryllium-9**
 - ${}^9\text{Be}$ (or beryllium-9) has 4 protons and 5 neutrons; you know it has 4 protons because it is a Be atom; you know it has five neutrons because it has a mass number of 9, and because it has four protons each contributing one unit to the mass, there must also be five neutrons
 - ${}^{12}\text{C}$ (or carbon-12) has 6 protons (because it is a C atom) and a mass number of 12; it must therefore have 6 neutrons
 - ${}^{13}\text{C}$ (or carbon-13) has 6 protons (because it is a C atom) and a mass number of 13; it must therefore have 7 neutrons
- Atoms such as carbon-12 and carbon-13 which have the same number of protons (ie the same atomic number) but different numbers of neutrons (ie different mass numbers) are called **isotopes**

Isotopes are atoms with the same atomic number but with different mass numbers

- Magnesium, for example, has 3 naturally occurring isotopes:
 - ${}^{24}\text{Mg}$: 12 protons, 12 neutrons
 - ${}^{25}\text{Mg}$: 12 protons, 13 neutrons
 - ${}^{26}\text{Mg}$: 12 protons, 14 neutrons



Test your knowledge 2.1: Understanding atomic numbers and mass numbers

Copy and complete the following table, showing the names and symbols of different isotopes:

Number of protons	Number of neutrons	Mass number	Name of atom	Atomic symbol
6	6	12	Carbon-12	${}^{12}\text{C}$
8	10			
	20	37		
			Hydrogen-3	
				${}^{11}\text{B}$
				${}^{25}\text{Mg}$



Test Your Knowledge 2.2: Understanding atomic symbols

- (a) Deduce the number of protons, electrons and neutrons in the following atoms:
 (i) ${}^1_1\text{H}$ (ii) ${}^{17}_8\text{O}$ (iii) ${}^{132}_{54}\text{Xe}$ (iv) ${}^{235}_{92}\text{U}$ (v) ${}^{14}_6\text{C}$
- (b) Deduce the symbol of the following atoms:
 (i) 19 protons, 20 neutrons
 (ii) 8 protons, 8 neutrons
 (iii) 1 proton, 2 neutrons
 (iv) 82 protons, 126 neutrons
 (v) 53 protons, 74 neutrons

(v) Units of Atomic Mass and Charge



Summary Activity 2.3: Mass and charge in atoms

- What is the mass and charge on a proton, neutron and electron?

- The unit of charge is the **coulomb** (C); protons and neutrons both have a charge of 1.602×10^{-19} coulombs; this quantity is commonly referred to as one unit of charge and is often given the symbol e
- The mass of protons and neutrons is not fixed but varies from atom to atom; there is therefore no fundamental unit of mass; in order to define an atomic mass unit, one nucleus needs to be taken as a reference nucleus; for this purpose carbon-12 is chosen (this atom contains six protons and six neutrons); the average mass of each nucleon in this atom is 1.661×10^{-27} kg; this quantity is defined as an **atomic mass unit**; one atom of carbon-12 therefore has a mass of exactly 12.0000 atomic mass units by definition; this scale is often referred to as the ${}^{12}\text{C} = 12$ scale

An atomic mass unit is $1/12^{\text{th}}$ of the mass of one atom of carbon-12

- All other atoms are measured relative to this quantity; the exact mass of an atom compared to $1/12^{\text{th}}$ of the mass of one atom of carbon-12 is called the **relative isotopic mass** of an atom

The relative isotopic mass of an atom is the ratio of the mass of that atom to $1/12^{\text{th}}$ of the mass of one atom of carbon-12

- Some relative isotopic masses are shown below; note that the relative isotopic mass is not exactly equal to the mass number but the differences are very small

Isotope	Mass number	Relative isotopic mass
${}^1\text{H}$	1	1.007825
${}^4\text{He}$	4	4.002603
${}^9\text{Be}$	9	9.012182
${}^{27}\text{Al}$	27	26.981538
${}^{59}\text{Co}$	59	58.933200



Thinkabout 2.4: Why do protons and neutrons not have a fixed mass?

Why does the mass of protons and neutrons vary from atom to atom? Why is this fact important in the generation of nuclear power?

Lesson 3 – What is relative atomic mass?

(vi) Relative Atomic Mass

- Atoms of the same element do not all have the same mass, as they can contain different numbers of neutrons (different isotopes)
- The Periodic Table does not show all of the different possible mass numbers of every atom; it is more useful to know the **average mass** of an atom, known as the **relative atomic mass**

The relative atomic mass of an atom is the ratio of the average mass of that atom to 1/12th of the mass of one atom of carbon-12

- The **relative atomic mass** of an atom therefore depends on the mass numbers of the different isotopes and their relative abundances; it can be calculated by the formula:
$$\text{relative atomic mass} = \frac{\sum (\text{percentage abundance of each isotope} \times \text{mass of each isotope})}{100}$$
 - eg neon consists of ²⁰Ne (90%) and ²²Ne (10%)
so the relative atomic mass of Ne = (90 x 20 + 10 x 22)/100 = 20.2
- The Periodic Table shows the relative atomic mass (ie the average mass) of every atom, based on the relative abundances of their different isotopes; the Periodic Table does NOT show the masses of individual atoms



Test Your Knowledge 3.1: Calculating relative atomic masses

- (a) Deduce the relative atomic masses of the following atoms:
- (i) Silicon (²⁸Si 92.2%, ²⁹Si 4.7%, ³⁰Si 3.1%)
 - (ii) Silver (¹⁰⁷Ag 51.9%, ¹⁰⁹Ag 48.1%)
 - (iii) Copper (⁶⁵Cu 30.9%, ⁶³Cu 69.1%)
 - (iv) Gallium (⁶⁹Ga 60.2%, ⁷¹Ga 39.8%)
 - (v) Magnesium (²⁴Mg 78.6%, ²⁵Mg 10.1%, ²⁶Mg 11.3%)



Extension 3.2: Understanding relative atomic masses

- (a) Bromine has two isotopes, with mass numbers 79 and 81. Its relative atomic mass is often given as 80. What does that tell you about the relative abundance of the two isotopes?
- (b) Iridium contains a mixture of ¹⁹¹Ir and ¹⁹³Ir. If the relative atomic mass of iridium is 192.2, what are the relative abundances of the two isotopes?

- For a more accurate calculation of **relative atomic mass** of an atom, it is important to use relative isotopic masses and not mass numbers; however, in practice, mass numbers are usually used because the calculations are simpler and the answer is still usually accurate to two decimal places



Test your knowledge 3.3: Calculating relative atomic mass accurately

The relative isotopic masses of ¹H and ²H are 1.0078 and 2.0141 respectively; their percentage abundances are 99.985% and 0.0155 respectively

- (a) Which isotope of hydrogen contains heavier nucleons?
- (b) Calculate the relative atomic mass of hydrogen to four decimal places

(vii) atoms and ions

- In an atom, the number of protons and electrons are the same; atoms can gain or lose electrons, and when they do this they become ions
- Ions are species in which the proton and electron numbers are not the same, and hence have an overall positive or negative charge; the number of electrons in a species can be deduced from its charge:
Eg if a magnesium atom (Mg) loses an electron it becomes Mg^+ ; if it loses two electrons it becomes Mg^{2+} ; if it gains an electron it becomes Mg^- ; Mg^+ , Mg^{2+} and Mg^- are all **ions**
 - $^{24}\text{Mg}^{2+}$ contains 12 protons, 12 neutrons and 10 electrons
 - $^{25}\text{Mg}^+$ contains 12 protons, 12 neutrons and 11 electrons
 - $^{24}\text{Mg}^-$ contains 12 protons, 12 neutrons and 13 electrons
- Ions with a positive charge are called CATIONS; ions with a negative charge are called ANIONS
- When atoms lose electrons to form cations, the name does not change – ie a sodium atom (Na) becomes a sodium ion (Na^+); when atoms gain electrons to form anions, the final one of two syllables of the name are removed and replaced with the ending -IDE; ie an oxygen atom (O) becomes an oxide ion (O^{2-})



Test Your Knowledge 3.4: Understanding atoms and ions

- (a) Deduce the number of protons, neutrons and electrons in the following species:
(i) $^4\text{He}^{2+}$ (ii) $^{27}\text{Al}^{3+}$ (iii) $^{79}\text{Br}^-$ (iv) $^3\text{H}^+$ (v) $^{19}\text{Na}^+$
- (b) Deduce the symbols of the following ions:
(i) 8 protons, 8 neutrons, 10 electrons
(ii) 82 protons, 125 neutrons, 80 electrons
(iii) 53 protons, 74 neutrons, 54 electrons
(iv) 84 protons, 126 neutrons, 82 electrons
(v) 26 protons, 30 neutrons, 23 electrons
- (c) Which of the species in questions 1 and 2 are cations, and which are anions?

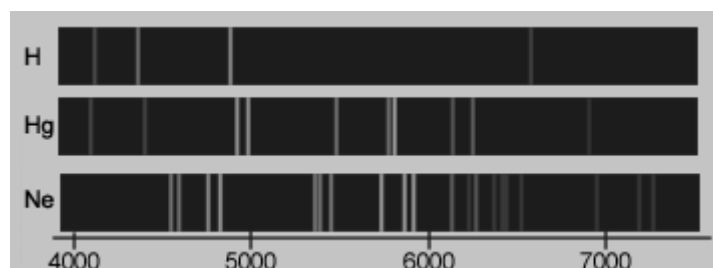
Lesson 4 – How are electrons arranged in atoms?

b) Electrons, Orbitals and Shells

(i) Energy Levels

- Electrons do not orbit the nucleus randomly; each electron must exist in one of a number of “energy levels”; these energy levels vary from atom to atom as they depend on the number protons in the nucleus as well as the other electrons; electrons in the lowest energy levels are those closest to the nucleus; electrons in higher energy levels are further away from the nucleus

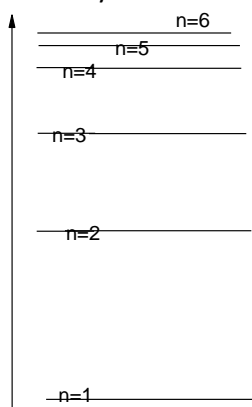
- There is clear experimental evidence that electrons can only have certain energies; this evidence comes in the form of **line spectra**; when atoms are heated or subjected to other radiation, electrons are excited from lower energy levels into higher energy levels (ie they jump into energy levels further from the nucleus); when they return to the lower energy levels (closer to the nucleus) they emit radiation with energy equal to the energy difference between the two levels; when this radiation is analysed, it is found that each atoms only emits radiation with a few different energies – this is called a line spectrum, and it shows that only certain energy levels can exist in the atom; each atom has a unique set of energy levels and so gives a distinct line spectrum which can be used to identify the atom; the analysis of emission or absorption of radiation in order to identify substances is known as **spectroscopy**



Line spectra of some different atoms

Source: www.bbc.co.uk

- Energy levels in an atom can be numbered 1,2,3 etc to infinity; level 1 (or the K level) is the lowest energy level (closest to the nucleus), followed by Level 2 (or the L level) and Level 3 (or the M level); energy level infinity corresponds to the energy of an electron which is not attracted to the nucleus at all (zero potential energy); all other energy levels have negative potential energy, and the energy levels converge as they approach infinity:



(ii) Orbitals and sub-levels

- Electrons do not in fact orbit the nucleus in an orderly way; in fact they occupy areas of space known as **orbitals**; the exact position of an electron within an orbital is impossible to determine; an orbital is simply an area of space in which there is a high probability of finding an electron
- Orbitals can have a number of different shapes, the most common of which are as follows:
 - s-orbitals: these are spherical

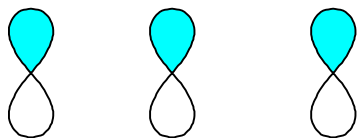


Every energy level contains one s-orbital

An s-orbital in the first energy level is a 1s orbital

An s-orbital in the second energy level is a 2s orbital, etc

- p-orbitals: these are shaped like a 3D figure of eight; they exist in groups of three:



Every energy level except the first level contains three p-orbitals

Each p-orbital in the same energy level has the same energy but different orientations: x, y and z

The p-orbitals in the second energy level are known as 2p orbitals ($2p_x$, $2p_y$, $2p_z$)

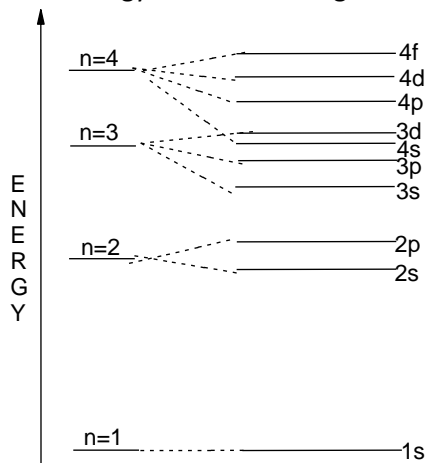
The p-orbitals in the third energy level are known as 3p orbitals ($3p_x$, $3p_y$, $3p_z$), etc

- In addition, the third and subsequent energy levels each contain five d-orbitals, the fourth and subsequent energy levels contain seven f-orbitals and so on; each type of orbital has its own characteristic shape
- s, p and d orbitals do not all have the same energy; in any given energy level, s-orbitals have the lowest energy and the energy of the other orbitals increases in the order $p < d < f$ etc; thus each energy level must be divided into a number of different **sub-levels**, each of which has a slightly different energy
- The number and type of orbitals in each energy level can thus be summarised as follows:

Energy level	Number and type of orbital		
	1 st sub-level	2 nd sub-level	3 rd sub-level
1 (K)	1 x 1s		
2 (L)	1 x 2s	3 x 2p	
3 (M)	1 x 3s	3 x 3p	5 x 3d

(iii) Shells and Sub-shells

- Since the different sub-levels have different energies, and the energies of the different levels get closer together with increasing energy level number, the high energy sub-levels of some energy levels soon overlap with the low energy sub-levels of higher energy levels, resulting in a more complex energy level diagram:



- Starting with the lowest energy, the orbitals can thus be arranged as follows:
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f
- Many of these sub-levels have similar energy and can be grouped together; a collection of sub-levels of similar energy is called a **shell**, and the sub-levels which occupy the shell can also be called **sub-shells**
Orbital: 1s | 2s 2p | 3s 3p | 4s 3d 4p | 5s 4d 5p | 6s 4f 5d 6p
Shell: 1st | 2nd | 3rd | 4th | 5th | 6th

- The arrangement of shells and orbitals/sub-levels can be summarised as follows:

Shell number	Orbitals/sub-levels in shell
1	1 x 1s
2	1 x 2s, 3 x 2p
3	1 x 3s, 3 x 3p
4	1 x 4s, 5 x 3d, 3 x 4p



Test your knowledge 4.1: Understanding energy levels, sub-shells, orbitals and shells

- What are line spectra and what do they tell you about the structure of atoms?
- How many sub-shells are there in the third energy level? What are they?
- How many orbitals of each type there in the third energy level?
- How many orbitals are there in the third shell?



Extension 4.2: Understanding energy levels, sub-shells, orbitals and shells

- Completed Test Your Knowledge 4.1? Describe all of the atomic orbitals in:
 - The fourth energy level
 - The fourth shell

Lesson 5 – How are electrons arranged in the different orbitals in an atom?

(iv) Rules for electrons occupying orbitals, sub-levels and shells

- There are three rules which determine how electrons occupy orbitals:

Rule 1: In a single orbital, it is impossible to put more than two electrons

- Electrons repel each other and this restricts the number that can be placed in one orbital
- since electrons are charged particles, and moving charges create a magnetic field, it is possible to create a small magnetic attraction between two electrons if they are spinning in opposite directions in the same orbital; this is the reason two electrons, and not one, are permitted in the same orbital
- if sharing the same orbital, the two electrons must be spinning in opposite directions
- this is known as the **Pauli Exclusion Principle - a maximum of two electrons can occupy the same orbital, and they must do so with opposite spins**
- it is thus possible to calculate the maximum possible number of electrons in each sub-level (sub-shell), and thus in each energy level or shell:

Shell	Number of electrons in each sub-level	Max. no of electrons
1	2 x 1s	2
2	2 x 2s, 6 x 2p	8
3	2 x 3s, 6 x 3p	8
4	2 x 4s, 10 x 3d, 6 x 4p	18
5	2 x 5s, 10 x 4d, 6 x 5p	18
6	2 x 6s, 14 x 4f, 10 x 5d, 6 x 6p	32

- Note that although the third energy level (Level M – 3s, 3p and 3d) can accommodate 18 electrons, the 3d sub-level is actually in the fourth shell, so the third shell only contains 8 electrons

Rule 2: Electrons always fill the lowest energy orbitals first

- this is known as the **Aufbau principle**; the lowest energy states are the most stable; so 1s is filled first, followed by 2s, 2p, 3s, 3p, 4s, 3d and then 4p

Rule 3: Electrons never pair up in the same orbital until all orbitals of the same energy are singly occupied, and that all unpaired electrons must have parallel spin

- this is known as **Hund's Rule of Maximum Multiplicity**; pairing electrons in the same orbital increases repulsion so is avoided if empty orbitals with the same energy are available; unpaired electrons in different orbitals always spin the same way as this creates a small magnetic attraction

**Test your knowledge 5.1: Filling electrons in orbitals, shells and energy levels**

- How many electrons can be placed in any one orbital?
- Identify all of the orbitals in the second energy level. How many electrons can be accommodated in this level?
- How many electrons can be accommodated in the third energy level?
- How many electrons can be accommodated in the third shell?

**Extension 5.2: Filling electrons in orbitals, shells and energy levels**

- Completed Test Your Knowledge 5.1? Now work out how many electrons can occupy the fourth energy level and the fourth shell

(v) Electronic Configuration

- The arrangement of electrons in an atom is known as its **electronic configuration**; it can be represented in two ways:
- The **arrow and box method** represents each orbital as a box and each electron as an arrow; the direction of spin is shown by the orientation of the arrow
The electronic configuration of the first 18 atoms using the arrow in box method is as follows:

	1s	2s	2p			3s	3p		
H	↑								
He	↑↓								
Li	↑↓	↑							
Be	↑↓	↑↓							
B	↑↓	↑↓	↑						
C	↑↓	↑↓	↑	↑					
N	↑↓	↑↓	↑	↑	↑				
O	↑↓	↑↓	↑↓	↑	↑				
F	↑↓	↑↓	↑↓	↑↓	↑				
Ne	↑↓	↑↓	↑↓	↑↓	↑↓				
Na	↑↓	↑↓	↑↓	↑↓	↑↓	↑			
Mg	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓			
Al	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑		
Si	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	
P	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑
S	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑
Cl	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑
Ar	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

- The **orbital method** indicates the number of electrons in each orbital with a superscript written immediately after the orbital; it is not possible to show the spin direction using this method; the electronic configurations of the first eighteen atoms can be shown with the orbital method as follows:

H:	$1s^1$
He:	$1s^2$
Li:	$1s^2 2s^1$
Be:	$1s^2 2s^2$
B:	$1s^2 2s^2 2p^1$
C:	$1s^2 2s^2 2p^2$ or $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1$
N:	$1s^2 2s^2 2p^3$ or $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$
O:	$1s^2 2s^2 2p^4$ or $1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$
F:	$1s^2 2s^2 2p^5$
Ne:	$1s^2 2s^2 2p^6$
Na:	$1s^2 2s^2 2p^6 3s^1$
Mg:	$1s^2 2s^2 2p^6 3s^2$
Al:	$1s^2 2s^2 2p^6 3s^2 3p^1$
Si:	$1s^2 2s^2 2p^6 3s^2 3p^2$ or $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1$
P:	$1s^2 2s^2 2p^6 3s^2 3p^3$ or $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$
S:	$1s^2 2s^2 2p^6 3s^2 3p^4$ or $1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$
Cl:	$1s^2 2s^2 2p^6 3s^2 3p^5$
Ar:	$1s^2 2s^2 2p^6 3s^2 3p^6$

- A shorthand form is often used for both the above methods; full shells are not written in full but represented by the symbol of the element to which they correspond, written in square brackets:

Eg $1s^2 2s^2 2p^6$ is represented as [Ne] and $1s^2 2s^2 2p^6 3s^2 3p^6$ is represented as [Ar]

- The shorthand electronic configuration of the elements with atomic numbers 18 to 36 can be written as follows:

		4s	3d					4p			
K	[Ar]	↑									
Ca	[Ar]	↑↓									
Sc	[Ar]	↑↓	↑								
Ti	[Ar]	↑↓	↑	↑							
V	[Ar]	↑↓	↑	↑	↑						
Cr	[Ar]	↑	↑	↑	↑	↑	↑	↑			
Mn	[Ar]	↑↓	↑	↑	↑	↑	↑	↑			
Fe	[Ar]	↑↓	↑↓	↑	↑	↑	↑	↑			
Co	[Ar]	↑↓	↑↓	↑↓	↑	↑	↑				
Ni	[Ar]	↑↓	↑↓	↑↓	↑↓	↑	↑				
Cu	[Ar]	↑	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓			
Zn	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓			
Ga	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑		
Ge	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	
As	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑
Se	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑
Br	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑
Kr	[Ar]	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

- Note the unusual structures of chromium ($[\text{Ar}]4s^13d^5$) and copper ($[\text{Ar}]4s^13d^{10}$); these anomalies exist because the difference in energy between the 3d and 4s electrons is very small:
 - In chromium the energy required to promote an electron from 4s to 3d is recovered in the reduced repulsion which results from the fact that they are no longer paired; thus the $4s^13d^5$ structure in Cr is preferred
 - In copper, the 3d orbitals are actually lower in energy than the 4s orbital, so the $4s^13d^{10}$ structure in Cu is preferred



Test your knowledge 5.3: Understanding electronic configurations of atoms

- Using the arrow in box method, draw the full electronic configuration of N, Ne and Na
- Using the orbital method, write the full electronic configuration of S, Ti and Br
- Using the shorthand arrow in box method, draw the electronic configuration of Fe and Se
- Using the shorthand orbital method, draw the electronic configurations of P and Cu



Extension 5.4: Understanding electronic configuration of atoms

- Completed Test Your Knowledge 5.3? Choose any three atoms in the Periodic Table with atomic numbers not greater than 36, and ask your partner to work out the electronic configuration of each one!

Lesson 6 – What is the Periodic Table?

(vi) Electronic Configuration of Ions

- The electronic configuration of ions can be deduced by simply adding or removing the appropriate number of electrons; the order in which electrons are to be removed can be deduced from the following rules:
 - remove outer shell electrons first
 - remove p-electrons first, then s-electrons and then d-electrons
 - remove paired electrons before unpaired electrons in the same sub-level



Test your knowledge 6.1: Understanding electronic configuration of ions

Write the full electronic configurations of Na^+ , O^- , Fe^{2+} , Cu^{2+} and Br^-



Extension 6.2: Understanding electronic configuration of ions

Completed Test Your Knowledge 6.2? Choose any three ions with atomic numbers not greater than 36, and ask your partner to work out the electronic configuration of each one!

(vii) chemical properties of atoms and isotopes

- The chemical properties of an atom depend on the strength of the attraction between the outer electrons and the nucleus; this in turn depends on the number of protons and on the electronic configuration, and so it follows that these two factors are instrumental in determining the chemical properties of an atom
- This is in contrast with the neutron number however, which has no effect on the chemical properties of an atom; neutrons have no charge and hence exert no attractive force on the nucleus
- Isotopes, therefore, tend to have very similar chemical properties since they have the same atomic number and the same electronic configuration; they differ only in number of neutrons, which do not influence the chemical properties of an atom; **isotopes therefore have identical chemical properties**

(viii) Size of cations and anions

- When you remove an electron from the outer shell of an atom, the repulsion between the remaining electrons in the outer shell decreases, and they move closer to the nucleus; **cations are therefore always smaller than the corresponding atoms**
 - Eg Na^+ is smaller than Na
- When you add an electron to the outer shell, the repulsion between the electrons increases, and they are pushed further away from the nucleus; **anions are therefore always larger than the corresponding atoms**
 - Eg Cl^- is larger than Cl

(ix) Introduction to the Periodic Table

- The periodic table is a list of all known atoms arranged in order of increasing atomic number, from 1 to 106; in addition to this, the atoms are arranged in such a way that atoms with the same number of shells are placed together, and atoms with the same number of electrons in the outer shell are also placed together; this is achieved by arranging the atoms in rows and columns as follows:
 - atoms with one shell are placed in the first row (ie H and He); atoms with two shells are placed in the second row (Li to Ne) and so on; a row of atoms thus arranged is called a **period**
 - atoms are aligned vertically (in columns) with other atoms in different rows, if they have the same number of electrons in their outer shell; for example, elements with one electron in their outer shell are all placed in the same column (Li, Na, K, Rb, Cs, Fr); a column of atoms thus arranged is called a **group**
- According to these principles, the periodic table can be constructed as follows:

I	II		III	IV	V	VI	VII	0										
		H								He								
Li	Be								B	C	N	O	F	Ne				
Na	Mg								Al	Si	P	S	Cl	Ar				
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
Cs	Ba	La	Ce - Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Th - Lw															

- The atoms Ce - Lu and Th - Lw belong in the periodic table as shown above. If they are placed there, however, periods 6 and 7 do not fit onto a page of A4, so they are placed below the other elements in most tables
- Hydrogen does not fit comfortably in any group; it is usually placed on its own in the Periodic Table
- From the position of an atom in the Periodic Table, it is possible to deduce how many shells it has and how many electrons it has in its outer shell
 - Eg Potassium (K) is in Group 1 and Period 4; it therefore has four shells, with one electron in its outer shell
- Using the periods and groups, the Periodic Table can be divided into four distinct blocks, based on the outer shell electronic configuration of the atoms:
 - the s-block atoms are all those with only s-electrons in the outer shell (eg H, Na)
 - the p-block atoms are all those with at least one p-electron in the outer shell (eg Ne, P)
 - the d-block atoms are all those with at least one d-electron and at least one s-electron but no f or p-electrons in the outer shell (eg Ti, Zn)
 - the f-block atoms are all those with at least one f-electron and at least one s-electron but no d or p-electrons in the outer shell (eg Ce, U)



Test your knowledge 6.3: Understanding the structure of the Periodic Table

- (a) Use the Periodic Table to predict the number of shells, and the number of electrons in the outer shell, in atoms of:
- (i) nitrogen
 - (ii) iodine
 - (iii) strontium
 - (iv) krypton
 - (v) tin
- (b) Which atom has:
- (i) four shells, with six electrons in its outer shell?
 - (ii) three shells, with five electrons in its outer shell?
- (a) In which blocks will you find the following atoms?
- (i) Ba
 - (ii) Fe
 - (iii) U
 - (iv) Br

Lesson 7 – How do the properties of atoms change with their position in the Periodic Table?

(b) Trends in the Periodic Table

- The properties of atoms in the Periodic Table show clear patterns related to the position of each atom in the Periodic Table:
 - atoms in the same group show similar properties
 - properties change gradually on crossing a period
- This means that as the atomic number increases, the atoms show trends which repeat themselves in the Periodic Table; this is known as the **Periodic Law**
- The properties of an atom depend on:
 - **the number of protons in the nucleus of the atom (the nuclear charge)**; the greater the number of protons, the stronger the attraction to the outer electrons; this causes electrons to move closer to the nucleus and makes them more difficult to remove
 - **the number of shells in the atom**; electrons in inner shells repel each other and the electrons in the outer shell; the electrons in inner shells therefore have the effect of reducing the force of attraction between the nucleus and the electrons in the outer shell; this effect is known as **shielding**; it causes electrons to move further away from the nucleus and makes them easier to remove
 - **which subshell the electrons in the outer shell are in**; p-electrons have a slightly higher energy than s-electrons and are further from the nucleus; this means that p-electrons can also be shielded by s-electrons in the same shell; this pushes them further away from the nucleus and makes them easier to remove
 - **whether or not the electrons in outer shell are sharing the same orbital**; if electrons in the outer shell have to share the same orbital, there is repulsion between them; this causes them to move further away from the nucleus and makes them easier to remove

(i) Atomic Size

- On moving across a period from left to right:
 - the number of protons increases; this causes a stronger attraction between the protons and the electrons in the outer shell, and so the outer shells are pulled closer to the nucleus
 - the number of shells, and so the shielding effect, does not change

So the size of atoms decreases across a period from left to right

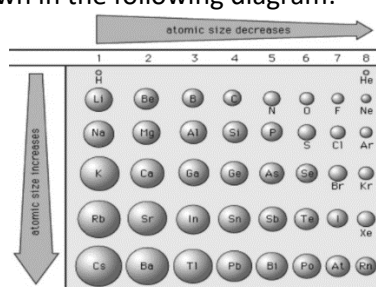
Eg Na is the largest atom in Period 3 and Ar is the smallest

- On moving down a group:
 - the number of protons increases
 - but the number of shells also increases; the extra shells increase the shielding and push the outer shell electrons further away from the nucleus; this effect outweighs the effect of more protons

So the size of the atoms increases down a group

Eg Be is the smallest atom in group 2 and Ra is the largest

- The variation in atomic size is shown in the following diagram:



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- Atoms become larger when they gain electrons, as the repulsion between the extra electrons increases, which pushes the outermost electrons further away from the nucleus; anions are always larger than their corresponding atoms
- Atoms become smaller when they lose electrons, as the repulsion between the remaining electrons decreases, and so the outermost electrons are able to move closer to the nucleus; cations are always smaller than their corresponding atoms



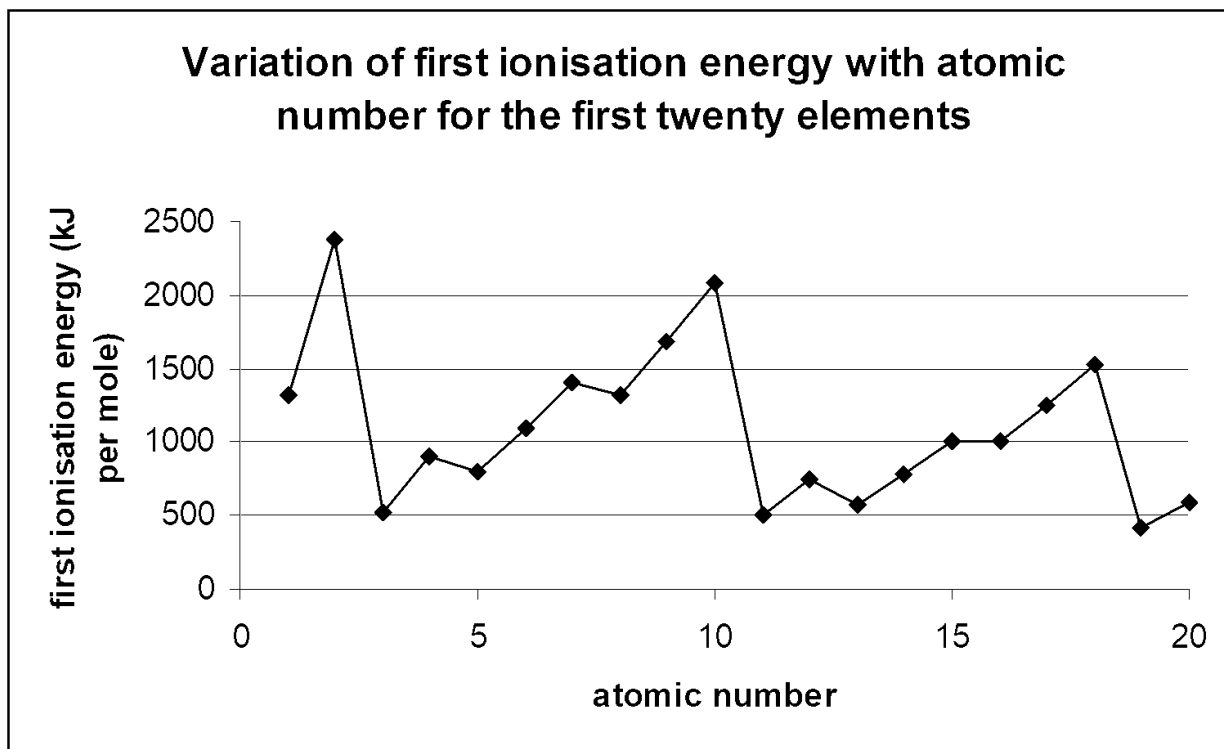
Test your knowledge 7.1: Understanding trends in atomic and Ionic Size

- Why is a magnesium atom smaller than a sodium atom?
- Why is a potassium atom larger than a sodium atom?
- Why is a chloride ion (Cl^-) larger than a chlorine atom?
- Why is a sodium ion (Na^+) smaller than a sodium atom?

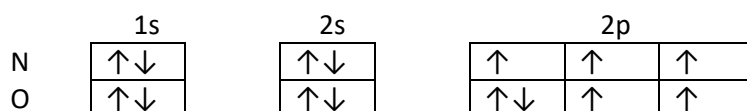
(ii) First Ionisation Energy

- The first ionisation energy of an element is the energy required to remove one electron from each of a mole of free gaseous atoms of that element; it can also be described as the energy change per mole for the process:
$$\text{M(g)} \rightarrow \text{M}^+(\text{g}) + \text{e}$$

- The first ionisation energies of the first 20 elements in the periodic table is shown below:



- From the table above, the following trends can be seen:
 - the first ionisation energy generally increases across a period (He > H, Ne > Li, Ar > Na etc), although there are some exceptions (eg Be > B, N > O, Mg > Al, P > S)
 - the first ionisation energy decreases down a group (He > Ne > Ar, Li > Na > K etc)
- The first ionisation energy increases across a period because:
 - the nuclear charge increases
 - the number of shells stays the same, so the shielding stays approximately the same
 - so the attraction between the nucleus and the outer shell electrons increases, so more energy is required to remove them
- The first ionisation energy decreases from Group 2 to Group 3, in an exception to the general trend, because:
 - the electron removed from Be is an s-electron but the electron removed from B is a p-electron
 - the p-electron in B experiences extra shielding from the 2s electrons, which reduces the attraction to the nucleus and makes the electron easier to remove; the s-electron in Be does not experience this extra shielding
 - this effect outweighs the greater nuclear charge in B
 - the decrease from Mg to Al happens for the same reason (extra shielding of 3p electron in Al)
- The first ionisation energy decreases from Group 5 to Group 6, in another exception to the general trend, because:



- in N the electron is removed from an unpaired orbital, but in O it is removed from a paired orbital
- the electrons in a paired orbital repel each other, which makes them easier to remove
- this effect outweighs the higher nuclear charge in O
- the decrease from P to S happens for the same reason

- The trend in first ionisation energies across a period can thus be summarised as follows:
 - there is a general increase across the period as the nuclear charge increases and the shielding remains the same, so the attraction between the nucleus and the outer electrons increases
 - there is a drop from the Group II atom (eg Be or Mg) to the Group III atom (eg B or Al) because in the Group 3 atom a p electron is being removed and the extra shielding from the s subshell actually causes a fall in the nuclear attraction
 - there is also a drop from the Group V atom (eg N or P) to the Group VI atoms (eg O or S) because the electron in the Group 6 atom is being removed from a paired orbital. The repulsion of the electrons in this orbital makes them less stable and easier to remove
- The first ionisation energy decreases down a group because
 - there are more shells, so more shielding of the electrons in the outer shell, which means that the attraction between the nucleus and the electrons in the outer shell is weaker, and less energy is needed to remove them
 - this effect outweighs the increase in nuclear charge on descending the group



Test your knowledge 7.2: Understanding trends in first ionisation energy

- Why is the first ionisation energy of Mg greater than that of Na?
- Why is the first ionisation energy of Mg greater than that of Al?
- Why is the first ionisation energy of P greater than that of S?
- Why is the first ionisation energy of Li greater than that of Na?



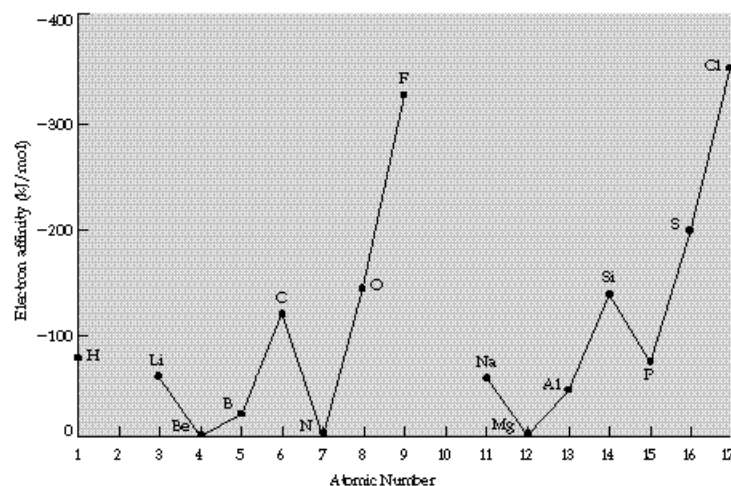
Extension 7.3: Understanding trends in first ionisation energy

- Completed Test Your Knowledge 7.2? Choose any two atoms in the s or p blocks of the Periodic Table which are either in the same Period or in the same Group, and explain which of the two atoms will have the higher first ionisation energy

Lesson 8 – What is meant by electron affinity?

(iii) First Electron Affinity

- **The first electron affinity of an element is the energy change when one electron is added to each of a mole of free gaseous atoms of that of that element;** it can also be described as the energy change per mole for the process: $M(g) + e^- \rightarrow M^-(g)$
- The first electron affinities of the first 18 atoms in the periodic table is shown below:



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- The factors affecting the trends in first electron affinity across a Period and down a Group (nuclear charge, shielding and repulsion) are the same as those affecting first ionisation energy, so the Periodic and group trends are very similar:
 - *In general, the first electron affinity increases across a period because the nuclear charge increases but the shielding remains the same, so the incoming electron is more attracted to the nucleus*
 - *From group I to group II, the electron affinity decreases because in group II the electron is added to a p-orbital, so it is shielded by the s-electrons in the outer shell; thus the incoming electron is less strongly attracted to the nucleus*
 - *The first electron affinity decreases from group IV to group V, since the any electron added to the group V atom will have to be paired, so there is more repulsion between the electrons and the incoming electron is not as easy to add*
 - *On descending a group, the nuclear charge increases but the number of shells also increases, so the shielding increases; the increase in shielding outweighs the increase in nuclear charge, and so the attraction between the nucleus and the outermost electrons decreases and there is a decrease in first electron affinity on descending a group*
- Note that the anomalies in the periodic trend in first affinity occur between Groups 1 and 2 and between Groups 4 and 5 because electrons are being added; the anomalies in the periodic trend in first ionisation energy occur between Groups 2 and 3 and between Groups 5 and 6 because electrons are being removed; otherwise the trends are the same for both ionisation energy and electron affinity



Test your knowledge 8.1: Understanding trends in first electron affinity

- Why is the first electron affinity of Al greater than that of Mg?
- Why is the first electron affinity of Na greater than that of Mg?
- Why is the first electron affinity of Si greater than that of P?
- Why is the first electron affinity of Li greater than that of Na?



Extension 8.2: Understanding trends in first electron affinity

- Completed Test Your Knowledge 8.1? Choose any two atoms in the s or p blocks of the Periodic Table which are either in the same Period or in the same Group, and explain which of the two atoms will have the higher first electron affinity

(iv) General Tendency of atoms to gain and lose electrons

- On moving across a period from left to right, the size decreases as the attraction between the protons and the electrons in the outer shell increases; it becomes harder to take away electrons, and easier to add electrons, as you move across the period:
 - **atoms on the left-hand side of the Periodic Table (especially in Groups 1 and 2) tend to lose electrons easily (and form cations) but do not gain electrons easily**
 - **atoms on the right-hand side of the Periodic Table (especially in Groups 6 and 7) tend to gain electrons easily (and form anions) but do not lose electrons easily**
- Atoms in Group 0 do not lose electrons easily because the nuclear attraction to the electrons in the outer shell is very high; but they also cannot gain electrons without creating an extra shell:
 - **atoms in Group 0 neither lose nor gain electrons easily**
- Within the same group, the attraction between the protons and the outer shell electrons is similar because the greater number of protons is balanced by the greater number of shells;
 - **atoms in the same group therefore show a similar, but not identical, tendency to gain and lose electrons**

- As the atomic size increases as you go down the group, there is a slight decrease in attraction between the nucleus and the electrons in the outer shell
 - **it therefore becomes slightly easier to remove electrons and slightly harder to add electrons as you descend a group**



Test your knowledge 8.3: Understanding the chemical properties of atoms

- Why does a chlorine atom accept an electron more readily than a sulphur atom?
- Why does a fluorine atom accept an electron more readily than a chlorine atom?
- Why is it easier to remove an electron from a sodium atom than a lithium atom?
- Why can helium, neon and argon neither gain nor lose electrons easily?

Lesson 9 – How has our understanding of the atom changed over time?

d) The Development of the Atomic Model

(i) **How scientific theories and models are developed**

- The way in which our understanding of atoms has changed over time is a good example of how scientific theories and models are developed and changed over time:
- A theory is a set of laws which explains and predicts certain properties of the universe; scientific theories need to be supported by evidence, and the theories are then tested by other scientists; if a theory is accepted, it becomes “science” until subsequent observations are made which disprove the theory, and a new theory is developed to replace it
- A typical cycle for the development of theories and models in science could be summarised as follows:
 - make an observation which contradicts an existing theory/model
 - develop a new theory/model which is consistent with all existing observations
 - communicate the new theory or model and encourage others to test it
 - the new theory/model will either be rejected or be accepted
 - continue to make observations until the theory/model is disproved



Thinkabout 9.1: How scientific theories change over time

Can you think of any other scientific theories that have changed over time? How and why did they change?

(ii) How our understanding of atomic structure has changed over time

- Since ancient times, scientists have been trying to explain what the universe is made of; the idea of small particles called atoms was first developed in ancient Greece; in the last 200 years, more detailed models of what an atom consists of have been developed, and then disproved:

When	Who	Why/How	Features of theory/model
5th century BC	Democritus and Leucippus	No evidence	The "atomos" is an indivisible building block of matter
1805	Dalton	Observed the Law of Multiple Proportions	Dalton's atomic theory: <ul style="list-style-type: none">- atoms are indivisible (from Democritus and Leucippus)- all atoms of the same element have the same mass, which is different from the mass of atoms of other elements
1897	Thomson	Discovered that metals, when heated in sealed tubes, emit negatively charged particles which are much smaller than the smallest atom	Thomson's atomic model: <ul style="list-style-type: none">- atoms are divisible- atoms contain electrons distributed inside a uniform sea of positive charge (also known as the "plum pudding" model)
1909	Rutherford	Fired alpha particles at gold foil and discovered that most passed straight through but that a small number were hugely deflected	Rutherford's atomic model: <ul style="list-style-type: none">- small positive nucleus concentrated at centre of atom- cloud of negatively electrons around nucleus

- Dalton, Thomson and Rutherford all used observations from experiments to make major contributions towards our understanding of atomic structure; later observations were made (most notably by Bohr and Schrodinger) which contradicted Rutherford's model, and the accepted model of the atom was amended again as a result



Test your knowledge 9.2: Describing how our understanding of the atom has changed over time

- Draw labelled diagrams of Dalton's Thomson's and Rutherford's atoms
- Underneath each diagram, summarise the key features of each atom
- Underneath the Dalton and Thomson summaries, write down who disproved each model and how they did it
- Then draw a timeline to show the changes in our understanding of atomic structure



Extension 9.3: Exploring further developments in understanding atomic structure

How did Niels Bohr and Edwin Schrodinger contribute to our understanding of atomic structure?

(iii) The History of the Periodic Table

- The first scientist to publish a Periodic Table was **Dimitri Mendeleev**, in 1869
- Previous scientists had discovered and named different atoms, and some scientists had already observed that the properties of the atoms tend to repeat themselves as atomic mass increases (the **Periodic Law**)

- But Mendeleev was the first to attempt to classify every atom within a specific row or column in a table:

Reihen	Gruppe I. R ⁰	Gruppe II. R ⁰	Gruppe III. R ⁰	Gruppe IV. RH ⁴ R ⁰	Gruppe V. RH ³ R ⁰	Gruppe VI. RH ² R ⁰	Gruppe VII. RH R ⁰	Gruppe VIII. R ⁰
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	---44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	---68	---72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	---100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	---	---	---	---
9	(-)	---	---	---	---	---	---	---
10	---	---	?Er=178	?La=180	Ta=182	W=184	---	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	---	---	---
12	---	---	---	Th=231	---	U=240	---	---

- Mendeleev arranged all of the known atoms in order of increasing atomic mass and placed all atoms with similar properties in the same column
- Mendeleev used his Periodic Table to predict the discovery of at least six new atoms, and to predict the properties of a number of new atoms before they were discovered; by creating the first Periodic Table, Mendeleev became the first scientist to develop a system of **classification** for atoms



Thinkabout 9.4: Mendeleev's Periodic Table

Mendeleev produced his Periodic Table before electrons and shells were discovered; can you see any mistakes he made as a result?

10.1 END-OF-TOPIC QUIZ
TOPIC 1 – ATOMS AND THE PERIODIC TABLE



1. (a) What are isotopes?
(b) Calculate the relative atomic mass of boron, given that it consists of 80% boron-11 and 20% boron-10
(c) Why do both isotopes of boron have identical chemical properties?
2. Deduce the number of protons, neutrons and electrons in $^{37}\text{Cl}^-$
3. Deduce the symbol for a species containing 11 protons, 12 neutrons and 10 electrons
4. What do emission spectra tell us about the structure of atoms?
5. Which types of orbital are found in the third energy level, and how many of each type?
6. Give the electronic configuration of Ni
7. Give the electronic configuration of Br^-
8. Define the term “first ionisation energy”
9. Explain why the first ionisation energy generally increases across a Period
10. Explain why the first ionisation energy decreases from Mg to Al
11. Explain why the first ionisation energy decreases from N to O
12. Explain why the first ionisation energy decreases down a group
13. Define the term “first electron affinity”
14. Explain why the first electron affinity decreases from Si to P
15. Explain Thomson’s contribution to our understanding of atomic structure