2 Atomic structure

2.3 Electrons: energy levels, atomic orbitals, ionisation energy

ELECTRONIC CONFIGURATION

2 Atomic structure

This topic describes the type, number and distribution of the fundamental particles which make up an atom and the impact of this on some atomic properties.

	Learning outcomes Candidates should be able to:
2.3 Electrons: energy levels, atomic orbitals, ionisation energy, electron affinity	a) describe the number and relative energies of the s, p and d orbitals for the principal quantum numbers 1, 2 and 3 and also the 4s and 4p orbitals b) describe and sketch the shapes of s and p orbitals c) state the electronic configuration of atoms and ions given the proton number and charge, using the convention $1s^2 2s^2 2p^6$, etc. d) (i) explain and use the term *ionisation energy* (ii) explain the factors influencing the ionisation energies of elements (iii) explain the trends in ionisation energies across a Period and down a Group of the Periodic Table (see also Section 9.1) e) deduce the electronic configurations of elements from successive ionisation energy data f) interpret successive ionisation energy data of an element in terms of the position of that element within the Periodic Table g) explain and use the term *electron affinity*

DISCLAIMER

A complete discussion of the experimental evidence for the modern theory of atomic structure is beyond the scope of the CIE A Level Syllabus.

In this chapter only the results of the theoretical treatment will be described. These results will have to be memorized as “rules of the game,” but they will be used so extensively throughout the general chemistry course that the notation used will soon become familiar.

1

ELECTRON ARRANGEMENT

The electronic configuration describes the arrangement of electrons in atoms.

An atom’s electrons are arranged outside the nucleus in energy levels (or shells).

Each shell or energy level holds a certain maximum number of electrons.

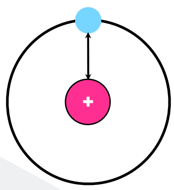
The energy of levels becomes greater as they go further from the nucleus and electrons fill energy levels in order.

2

IONISATION ENERGY

Ionisation energy is a measure of the energy needed to remove an electron from a gaseous atom or ion. It measures how strongly an atom or ion holds on to its electrons.

Attraction between the nucleus and an electron



The greater the pull of the nucleus, the harder it will be to pull an electron away from an atom.

FIRST IONISATION ENERGY

Ionisation energies give evidence for the arrangement of electrons in atoms in shells and sub-shells.

The first ionisation energy for an element is the energy needed to remove **one mole** of electrons from **one mole** of **gaseous atoms**.

$Na (g) \rightarrow Na^{+}(g) + e^{-}$ $\Delta H_{i1} = 496$ kJ mol^{-1}

$Ca (g) \rightarrow Ca^{+}(g) + e^{-}$ $\Delta H_{i1} = 590$ kJ mol^{-1}

$F (g) \rightarrow F^{+}(g) + e^{-}$ $\Delta H_{i1} = 1680$ kJ mol^{-1}

SUCCESSIVE IONISATION ENERGIES

Successive ionisation energies for the same element measure the energy to remove a second, third, fourth electron and so on.

$Na^{+}(g) \rightarrow Na^{2+}(g) + e^{-}$ $\Delta H_{i2} = 4563\ kJ\ mol^{-1}$

$Na^{2+}(g) \rightarrow Na^{3+}(g) + e^{-}$ $\Delta H_{i3} = 6913\ kJ\ mol^{-1}$

It is possible to measure energy changes involving ions which do not normally appear in chemical reactions.

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SKILL CHECK 1

Which equation represents the second ionisation energy of an element X?

A $X_{(g)} \longrightarrow X^{2+}(g) + 2e^{-}$

B $X^{+}_{(g)} \longrightarrow X^{2+}_{(g)} + e^{-}$

C $X_{(g)} + 2e^{-} \longrightarrow X^{2-}_{(g)}$

D $X^{-}_{(g)} + e^{-} \longrightarrow X^{2-}_{(g)}$

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SKILL CHECK 2

Write equations to represent the first ionisation of:

(a) potassium

(b) argon

(c) bromine

(d) fluorine

SKILL CHECK 3

The successive ionisation energies $kJmol^{-1}$ of element X are listed below. **Identify the group in the periodic table in which X occurs.**

Ionisation energies of X:

1st	2nd	3rd	4th	5th	6th	7th
950	1800	2700	4800	6000	12300	15000

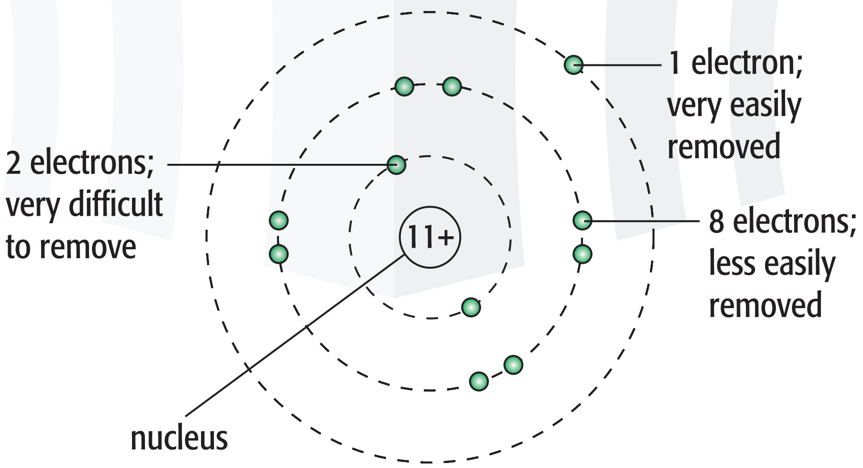
SUCCESSIVE IONISATION ENERGIES

After an electron has been removed the rest of them will be more strongly attracted by the nucleus.

Hence more energy is required to pull the 2^{nd} electron and thus the 2^{nd} I.E. is greater than the 1^{st} I.E.

Successive ionisation energies are always greater than the previous one.

SUCCESSIVE IONISATION ENERGIES



EVIDENCE OF ENERGY LEVELS

The arrangement of electrons in an atom of any element can be deduced from the values of successive ionisation energies.

The successive I.E of sodium illustrate the change clearly.

electron removed	1	2	3	4	5	6	7	8	9	10	11
ionisation energy	500	4600	6900	9500	13400	16600	20100	25500	28900	141000	158000

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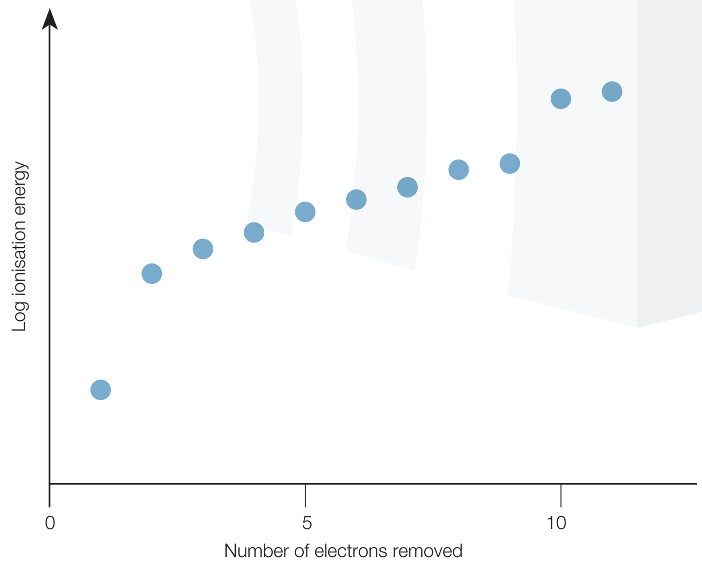
IONISATION ENERGIES OF SODIUM



The graph shows shows **successive** ionisation energies against the number of electrons removed for sodium.

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IONISATION ENERGIES OF SODIUM

There is a big difference between **some** successive ionisation energies. For sodium the first big difference occurs between the 1st and 2nd ionisation energies.

These **large changes** indicate that for the **second** of these two ionisation energies, the electron being removed is from a shell **closer** to the nucleus.

IONISATION ENERGIES OF SODIUM

electron removed	1	2	3	4	5	6	7	8	9	10	11
ionisation energy	500	4600	6900	9500	13400	16600	20100	25500	28900	141000	158000

There is a **big jump** in the value of the **second** ionisation energy. This suggests that the second electron is in a shell closer to the nucleus than the first electron.

Taken together, the 1st and 2nd ionisation energies suggest that sodium has **one electron** in its **outer shell**.

IONISATION ENERGIES OF SODIUM

electron removed	1	2	3	4	5	6	7	8	9	10	11
ionisation energy	500	4600	6900	9500	13400	16600	20100	25500	28900	141000	158000

From the second to the ninth electrons removed there is only a **gradual** change in successive ionisation energies. This suggests that all these eight electrons are in the **same** shell.

There is a **big jump** in the value of the 10th ionisation energy. This suggests that the 10th electron is in a shell **closer** to the nucleus than the 9th electron.

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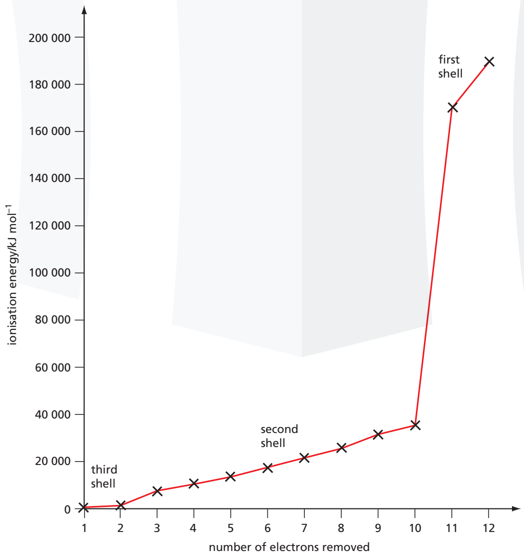
IONISATION ENERGIES OF MAGNESIUM



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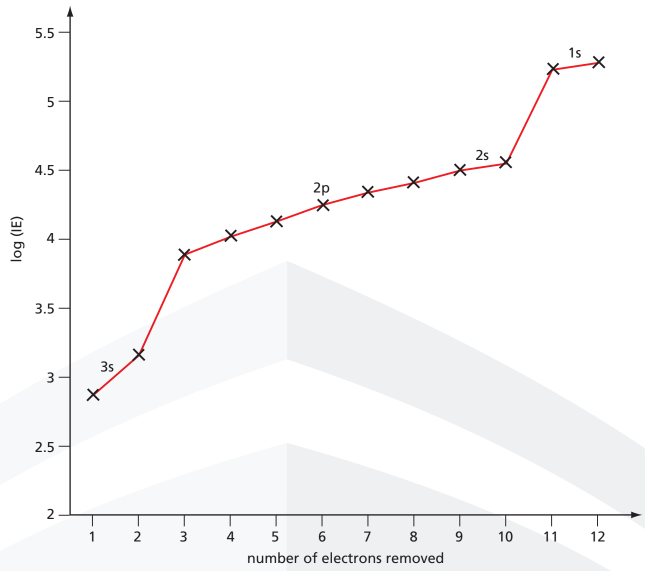
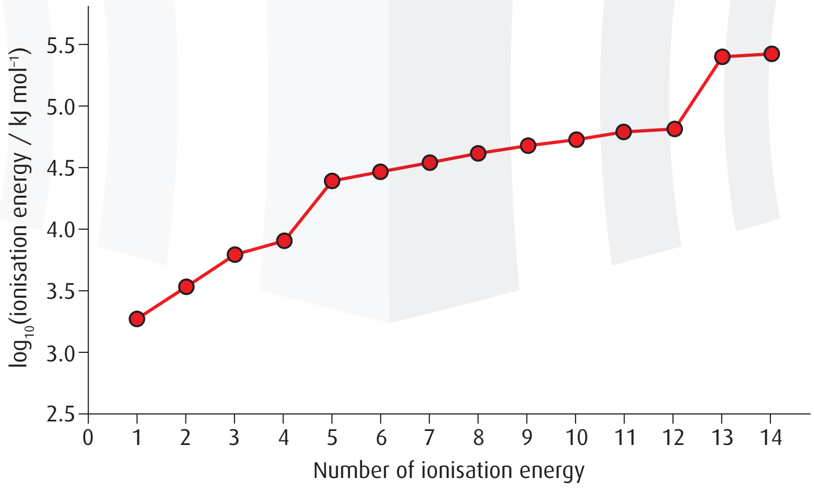
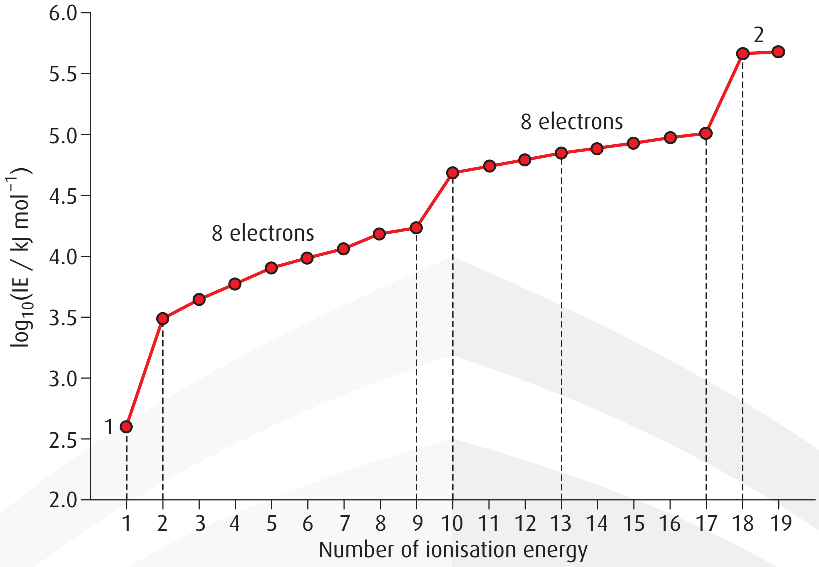
IONISATION ENERGIES OF MAGNESIUM



IONISATION ENERGIES OF SILICON



IONISATION ENERGIES OF POTASSIUM



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SKILL CHECK 4

Write an equation to represent the 5th ionisation energy of Fluorine.

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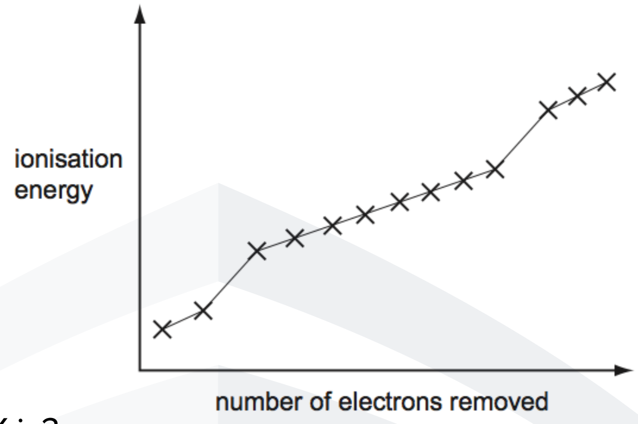
SKILL CHECK 5

The graph shows the first thirteen ionisation energies for element **X**.



What group is X in?

SKILL CHECK 6

The first ionisation energies of four consecutive elements in the Periodic table are:

Sodium = 494 kJ mol-1
Magnesium = 736 kJ mol-1
Aluminium = 577 kJ mol-1
Silicon = 786 kJ mol-1

a Explain the general increase in ionisation energies from Sodium to Silicon.
b Explain why Aluminium has a lower first ionisation energy than Magnesium.

SKILL CHECK 7

The first six ionisation energies of an element are, or 1090, 2250, 4610, 6220, 37,800, and 47,300kJ mol^{-1}. Which group in the Periodic Table does this element belong to? Explain your decision.

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SKILL CHECK 8

The successive ionisation energies ΔH_i, of an element X are shown in the table below. Which group in the Periodic Table does X belong to?

	Number of electrons removed									
	1	2	3	4	5	6	7	8	9	10
ΔH_i / kJ mol-1	1000	2260	3390	4540	7010	8500	27100	31670	36580	43140

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SUCCESSIVE IONISATION ENERGIES

We can use successive ionisation energies in this way to confirm:

- The simple electronic configuration of elements.
- The number of electrons in the outer shell of an element and hence the group to which the element belongs.

The successive ionisation energies for an element rise and there are big jumps in value each time electrons start to be removed from the next shell in towards the nucleus.

IONISATION ENERGIES OF OXYGEN

electron removed	1	2	3	4	5	6	7	8
ionisation energy	1310	3390	5320	7450	11000	13300	71300	84100

Large increases can be used to predict the group of any element. The electron configuration of oxygen is 2,6.

Since the large change is after the removal of 6 electrons, it signifies that there are 6 electrons in the shell farthest from the nucleus.

Therefore, Oxygen is in Group VI.

SUCCESSIVE IONISATION ENERGIES

Element		ELECTRONS REMOVED 1	2	3	4	5	6	7	8	9	10	11
1	H	1310										
2	He	2370	5250									
3	Li	519	7300	11800								
4	Be	900	1760	14850	21000							
5	B	799	2420	3660	25000	32800						
6	C	1090	2350	4620	6220	37800	47300					
7	N	1400	2860	4580	7480	9450	53300	64400				
8	O	1310	3390	5320	7450	11000	13300	71300	84100			
9	F	1680	3370	6040	8410	11000	15200	17900	92000	106000		
10	Ne	2080	3950	6150	9290	12200	15200	20000	23000	117000	131400	
11	Na	494	4560	6940	9540	13400	16600	20100	25500	28900	141000	158700

SKILL CHECK 9

The successive ionisation energies, in kJ mol^{-1}, of different elements are given below. Which groups are the following elements in?

	1	2	3	4	5	6	7	8
A	799	2420	3660	25000				
B	736	1450	7740	10500				
C	418	3070	4600	5860				
D	870	1800	3000	3600	5800	7000	13200	
E	950	1800	2700	4800	6000	12300		

SKILL CHECK 10

The successive ionisation energies of beryllium are 900, 1757, 14,849 and 21,007 kJ mol^{-1}.

a What is the atomic number of beryllium?

b Why do successive ionisation energies of beryllium always get more endothermic?

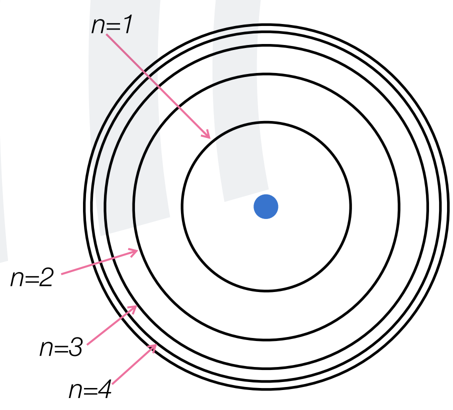
c To which group of the Periodic Table does this element belong?

SHELLS (ENERGY LEVELS)

The principal energy levels are designated n = 1, 2, 3, and so forth.

The energy levels are **not** equally spaced.

The energy gap between successive levels gets increasingly smaller as the levels move further from the nucleus.



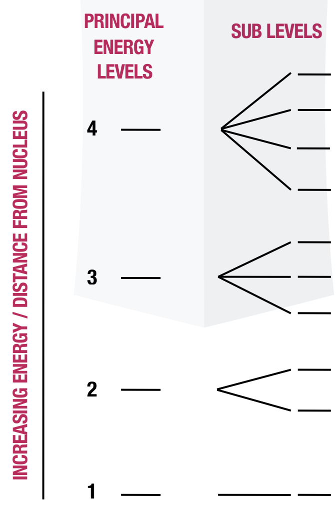
SUB-LEVELS (SUB-SHELLS)

Electron shells are numbered 1,2,3 etc. These numbers are known as the **principle quantum numbers**.

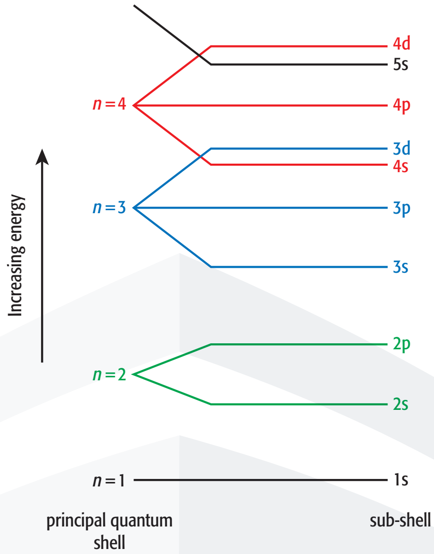
Each energy level (shell) consists of a number of sub-levels (sub-shells), labeled **s**, **p**, **d**, or **f**.

Energy Level	Number of sub-levels	Name of sub-levels
1	1	s
2	2	s, p
3	3	s, p, d
4	4	s, p, d, f

SUB-LEVELS (SUB-SHELLS)



SUB-LEVELS (SUB-SHELLS)



SUB-LEVELS

Each sub-level can hold a certain maximum number of electrons.

Type of sub-level	Maximum # of electrons
s	2
p	6
d	10
f	14

ORBITALS

An **atomic orbital** is a region of space around the nucleus of an atom which can be occupied by one or two electrons **only**.

Each sub-level contains a **fixed number** of orbitals that contain electrons.

Type of sub-level	Maximum # of electrons	Number of orbitals
s	2	1
p	6	3
d	10	5
f	14	7

ORBITALS

Electrons are viewed as charged clouds and the region, which encloses almost all the charge cloud, is the orbital.

The region in space where the probability of finding an electron in maximum is called the orbital.

The boundary surface encloses the region where the probability of finding an electron is high.

The orbitals are of different **three-dimensional** shapes and are named **s**, **p**, **d**, **f** etc.

ORBITALS

The cluster of dots show the probability of finding an electron at different distances from the nucleus.

SHELLS, SUB SHELLS AND ORBITALS

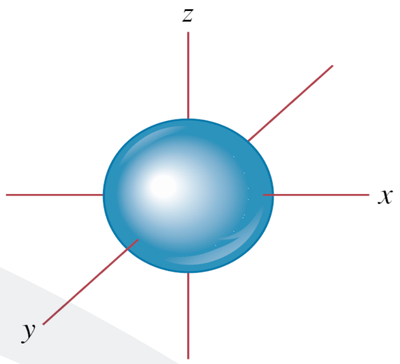
Energy Level	Type of sub-level	Number of orbitals	Maximum # of electrons
1	s	1	2
2	s	1	2
	p	3	6
3	s	1	2
	p	3	6
	d	5	10
4	s	1	2
	p	3	6
	d	5	10

"S" ORBITAL

The s orbital is spherically shaped. There is **only one** s orbital for each shell.

The 2s orbital in the second principal quantum shell has the same shape as the 1s orbital in the first quantum shell.

They are both spherical, but electrons in the 2s orbital have more energy than electrons in the 1s orbital.



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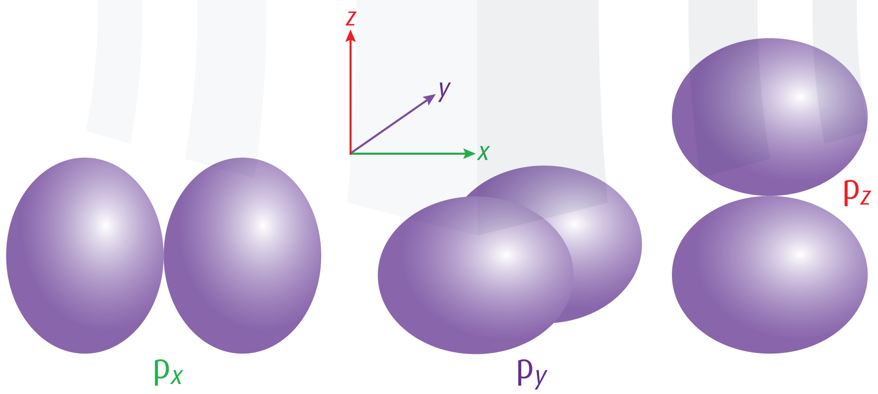
"P" ORBITAL

For each shell (**except the first**), there are three p orbitals.



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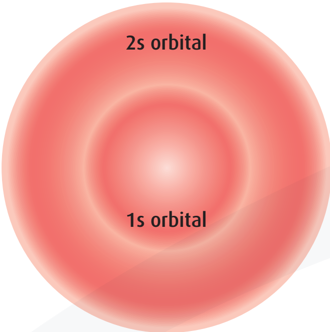
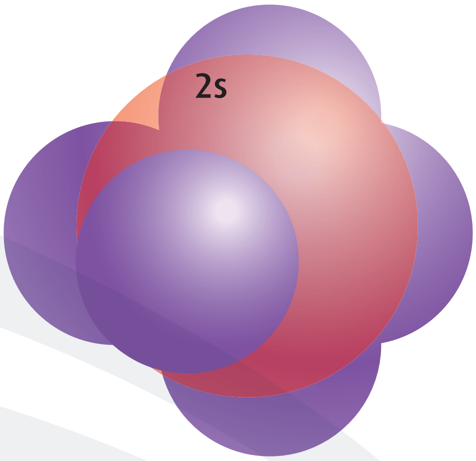
THE FIRST TWO SHELLS





ELECTRONIC CONFIGURATION

Electrons are distributed in different energy levels in the atom of the element.

The order in which they fill up the sub-levels is governed by **stability**.

When electrons fill up the orbitals having the **least energy** they attain **maximum stability**.

There are three principles that describe how electrons fill up in orbitals.

Aufbau Principle: Electrons enter the orbital that is available with the **lowest** energy. The orbitals are arranged in the order of increasing energy and the electrons are added until the proper number of electrons for the element have been accommodated

ELECTRONIC CONFIGURATION

Pauli's Exclusion Principle: No orbital can accommodate more than two electrons. If there are two electrons in an orbital, they must have opposite spin.

Hund's Rule of Maximum Multiplicity: When there are a number of orbitals of equal energy, electrons first fill them up individually and then get paired. By filling up individually, **mutual repulsion** between electrons is avoided and thereby maximum stability is achieved.

ELECTRON SPIN

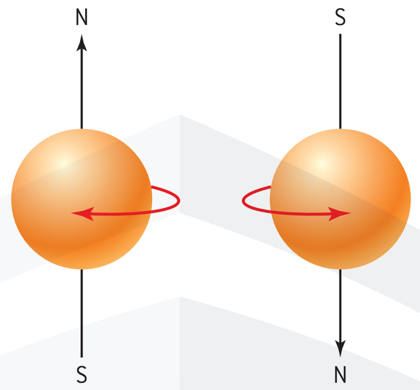
Electrons are all identical. The only way of distinguishing them is by describing how their **energies** and **spatial distributions** differ.

Thus an electron in a 1s orbital is different from an electron in a 2s orbital because it occupies a different region of space **closer to the nucleus**, causing it to have **less** potential energy.

An electron in a $2p_x$ orbital differs from an electron in a $2p_y$ orbital because although they have exactly the same potential energy, they occupy different regions of space.

ELECTRON SPIN

There can only be two electrons in each orbital, and they must have opposite directions of spin.



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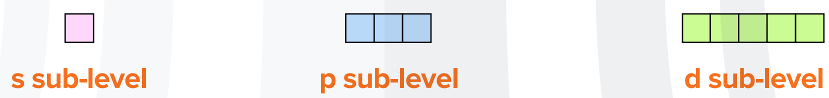
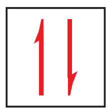
ELECTRON SPIN

For convenience, an “**electrons-in-boxes**” notation is used to represent the electrons in these atomic orbitals:



Each box represents one orbital:



an orbital with electrons in opposite spins

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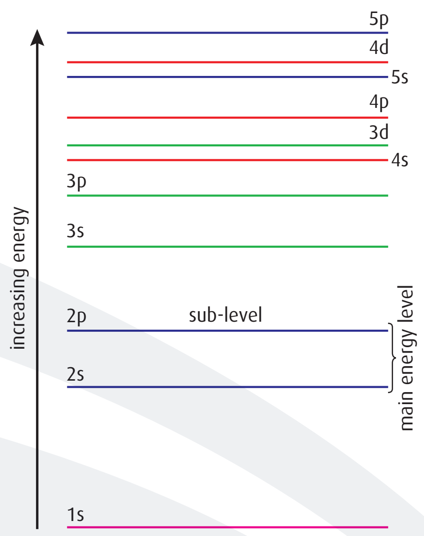
ORDER OF FILLING ORBITALS

The **principal energy levels** (shells) get closer together as you get further from the nucleus. This results in an overlap of sub- levels.



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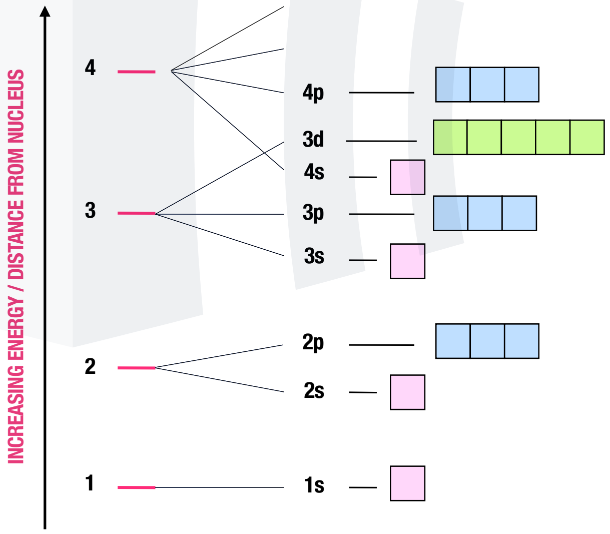
ORDER OF FILLING ORBITALS

The first example of this overlap occurs when the 4s orbital is filled before the 3d orbital



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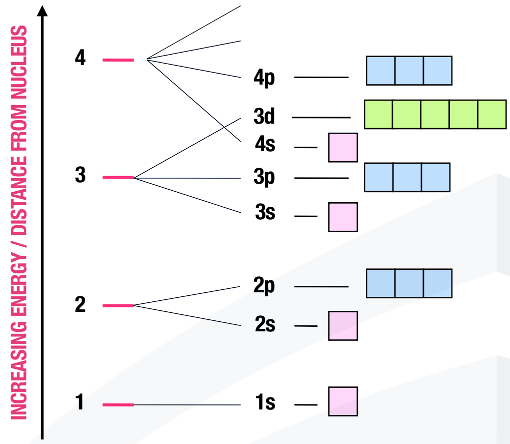
THE 'AUFBAU' PRINCIPAL



The following sequence will show the 'building up' of the electronic structures of the first 36 elements in the periodic table.

Electrons are shown as **half-headed** arrows and can spin in one of two directions.

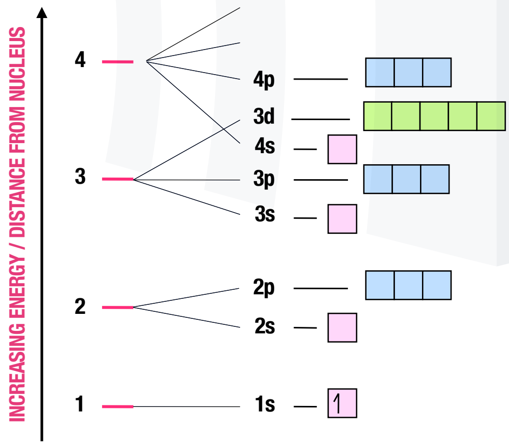
Orbitals are color-coded as below:

s orbitals

p orbitals

d orbitals

HYDROGEN

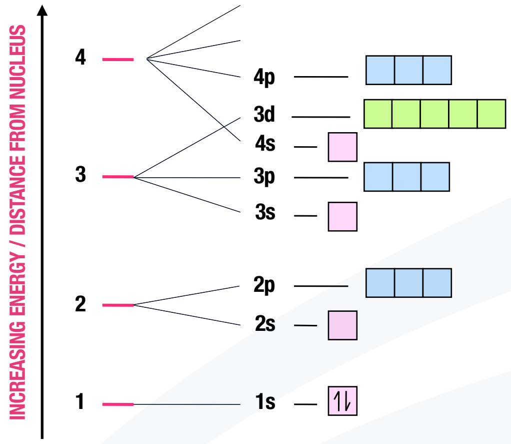


$1s^1$

Hydrogen atoms have one electron. This goes into a vacant orbital in the lowest available energy level.

HELIUM

$1s^2$

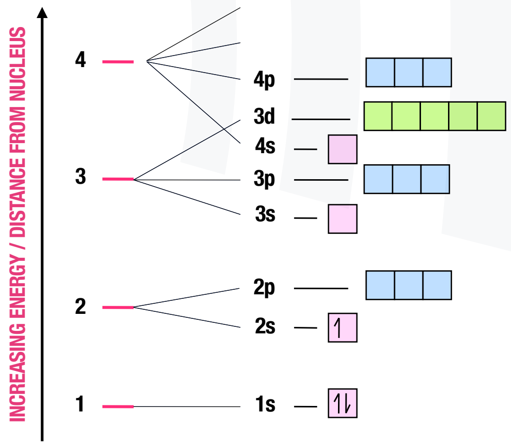


Every orbital can contain 2 electrons, provided the electrons are spinning in opposite directions.

The two electrons in a helium atom both go in the 1s orbital.

LITHIUM

$1s^2\ 2s^1$

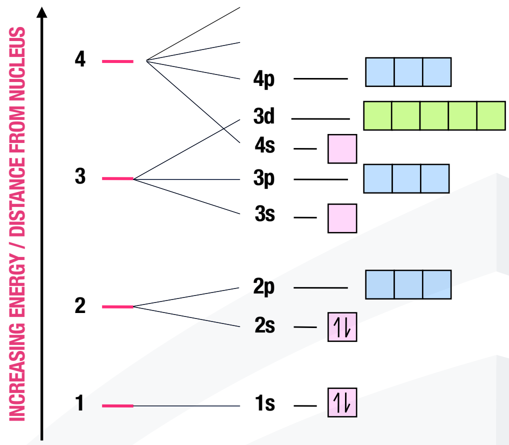


1s orbitals can hold a maximum of two electrons so the third electron in a lithium atom must go into the next available orbital of higher energy. This will be further from the nucleus in the second principal energy level.

The second principal level has two types of orbital (s and p). An s orbital is lower in energy than a p.

BERYLLIUM

$1s^2\ 2s^2$



Beryllium atoms have four electrons so the fourth electron pairs up in the 2s orbital. The 2s sub level is now full.

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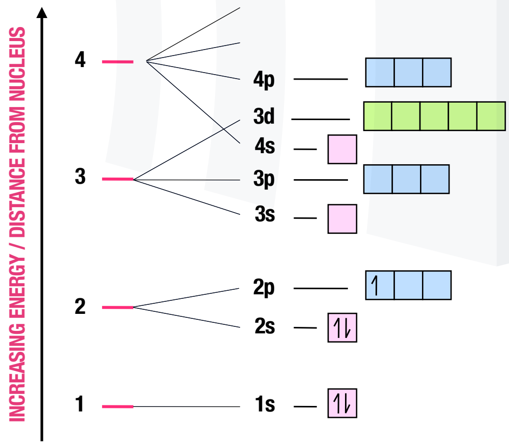
BORON

$1s^2\ 2s^2\ 2p^1$



As the 2s sub level is now full, the fifth electron goes into one of the three p orbitals in the 2p sub level.

The 2p orbitals are slightly higher in energy than the 2s orbital.

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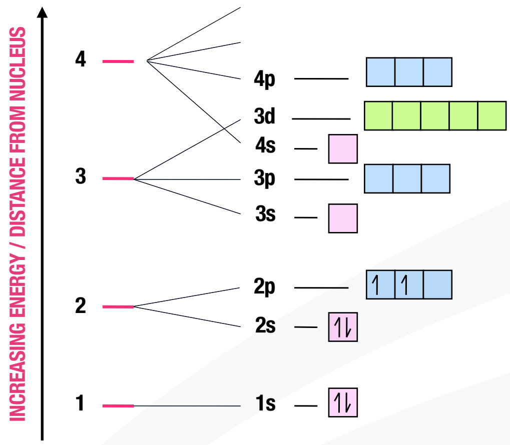
CARBON

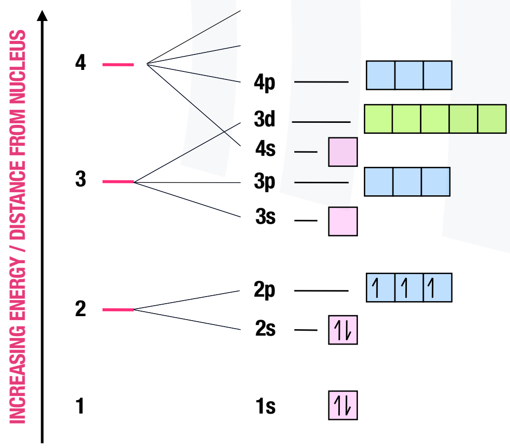
$1s^2\ 2s^2\ 2p^2$



The next electron in doesn't pair up with the one already there. This would give rise to **repulsion between the similarly charged species**. Instead, it goes into another p orbital which means **less repulsion, lower energy and more stability**.

NITROGEN

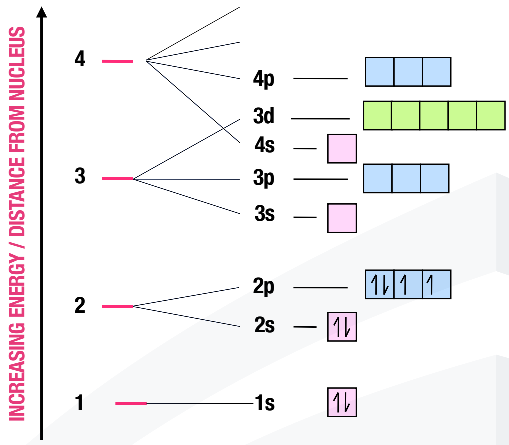
$1s^2\ 2s^2\ 2p^3$



The next electron will not pair up but goes into a vacant p orbital. All three electrons are now unpaired. This gives **less repulsion, lower energy** and therefore **more stability**.

OXYGEN

$1s^2\ 2s^2\ 2p^4$



With all three orbitals half-filled, the eighth electron in an oxygen atom must now pair up with one of the electrons already there.

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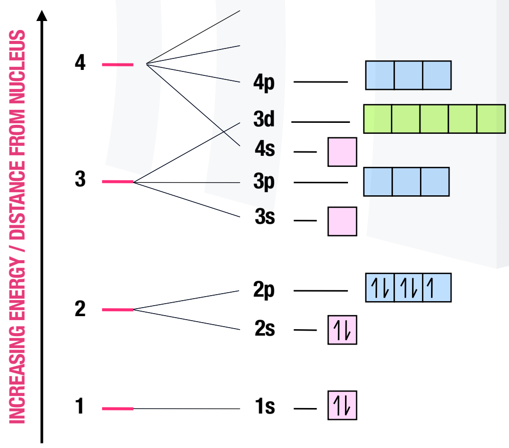
FLUORINE

$1s^2\ 2s^2\ 2p^5$



The electrons continue to pair up with those in the half-filled orbitals.

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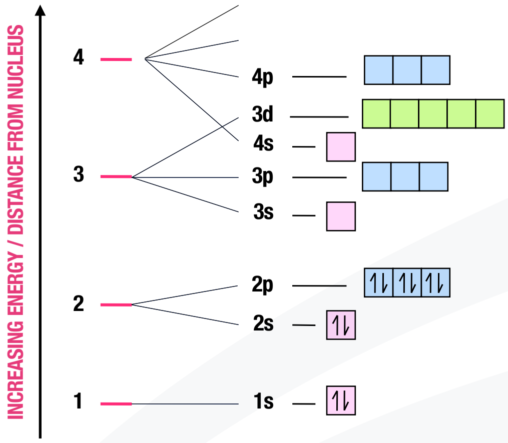
NEON

$1s^2\ 2s^2\ 2p^6$



The electrons continue to pair up with those in the half-filled orbitals. The 2p orbitals are now completely filled and so is the second principal energy level.

In the older system of describing electronic configurations, this would have been written as 2,8.

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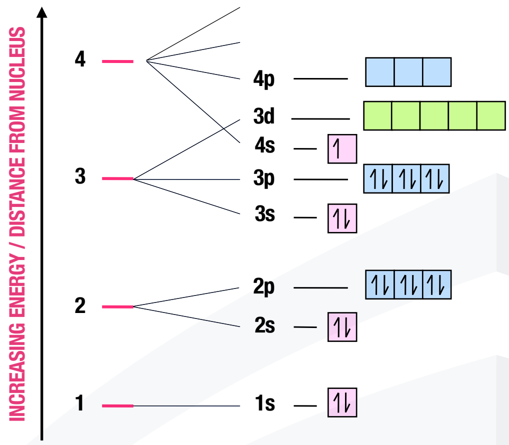
SODIUM TO ARGON

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	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$
P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$
Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

60

POTASSIUM

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

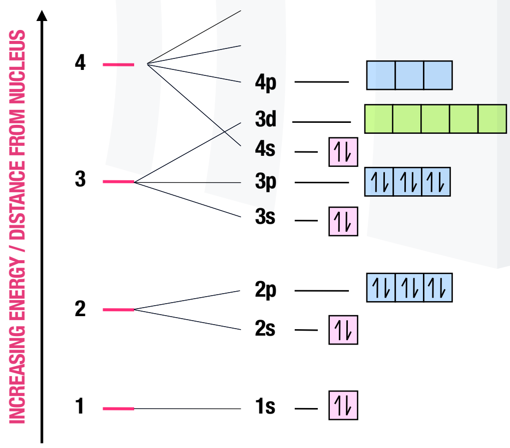


Because the **principal energy levels get closer together as you go further from the nucleus** coupled with the **splitting into sub energy levels**, the **4s orbital is of a LOWER ENERGY than the 3d orbitals** so it gets filled first.

61

CALCIUM

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$



As expected, the next electron pairs up to complete a filled 4s orbital.

62

ELEMENTS WITH PROTON NUMBERS 1 TO 10

Period 1			H					He
Atomic no.			1					2
Electron shell structure			1					2
Electron sub-shell structure			$1s^1$					$1s^2$
Period 2	**Li**	**Be**	**B**	**C**	**N**	**O**	**F**	**Ne**
Atomic no.	3	4	5	6	7	8	9	10
Electron shell structure	2, 1	2, 2	2, 3	2, 4	2, 5	2, 6	2, 7	2, 8
Electron sub-shell structure	$1s^2$ $2s^1$	$1s^2$ $2s^2$	$1s^2$ $2s^22p^1$	$1s^2$ $2s^22p^2$	$1s^2$ $2s^22p^3$	$1s^2$ $2s^22p^4$	$1s^2$ $2s^22p^5$	$1s^2$ $2s^22p^6$

63

ELEMENTS WITH PROTON NUMBERS 11 TO 20

Period 3	Na	Mg	Al	Si	P	S	Cl	Ar
Atomic no.	11	12	13	14	15	16	17	18
Electron shell structure	2, 8, 1	2, 8, 2	2, 8, 3	2, 8, 4	2, 8, 5	2, 8, 6	2, 8, 7	2, 8, 8
Electron sub-shell structure	$1s^2$ $2s^22p^6$ $3s^1$	$1s^2$ $2s^22p^6$ $3s^2$	$1s^2$ $2s^22p^6$ $3s^23p^1$	$1s^2$ $2s^22p^6$ $3s^23p^2$	$1s^2$ $2s^22p^6$ $3s^23p^3$	$1s^2$ $2s^22p^6$ $3s^23p^4$	$1s^2$ $2s^22p^6$ $3s^23p^5$	$1s^2$ $2s^22p^6$ $3s^23p^6$
Period 4	**K**	**Ca**						
Atomic no.	19	20						
Electron shell structure	2, 8, 8, 1	2, 8, 8, 2						
Electron sub-shell structure	$1s^2$ $2s^22p^6$ $3s^23p^6$ $4s^1$	$1s^2$ $2s^22p^6$ $3s^23p^6$ $4s^2$						

SKILL CHECK 11

Copy and complete the following information for the quantum shell with **principal quantum number 3**.

(a) total number of sub-shells

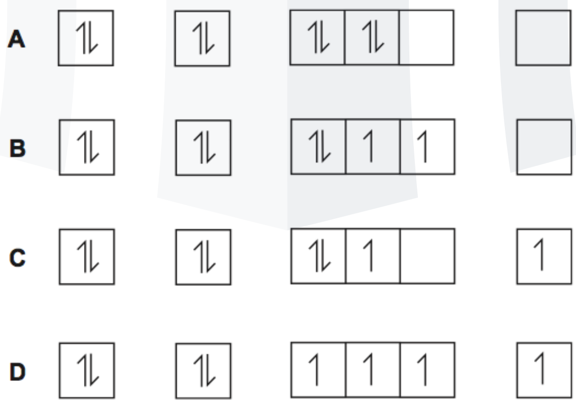
(b) total number of orbitals

(c) number of different types of orbital

(d) maximum number of electrons in the shell

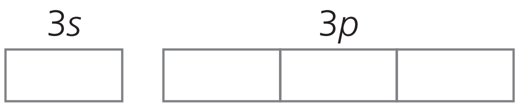
SKILL CHECK 12

An atom has eight electrons. Which diagram shows the electronic configuration of this atom in its lowest energy state?



SKILL CHECK 13

Fill in the outer electrons of a phosphorus atom in the boxes below.



SKILL CHECK 14

Give the electron orbital configuration for the ground state of the following atoms or ions:

(a) N

(b) O^{2-}

(c) Ca^{2+}

(d) Al^{3+}

(e) P^{3-}

SKILL CHECK 15

Write the full electronic configuration of:

(a) ^{75}As

(b) $^{75}As^{3-}$

(c) ^{32}S

(d) $^{32}S^{2-}$

SKILL CHECK 16

Name all element in the third period (row) of the periodic table with the following:

a three valence electrons
b four 3*p* electrons
c six 3*p* electrons
d two 3*s* electrons and no 3*p* electrons

EXCEPTIONS IN TRANSITION ELEMENTS

Though the s orbital is at lower energy level than the d of the penultimate shell, after the filling of the d sub-level, the order changes.

The d electrons, because of the shape of the d orbital, penetrate into the region of space between the nucleus and the s orbital and repel the s electrons and push them to higher energy level.

Before filling: 1s 2s 2p 3s 3p 4s 3d

After filling: $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^x$ $4s^2$

71

EXCEPTIONS IN TRANSITION ELEMENTS

Thus the electronic configuration of iron is

$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $3d^6$ $4s^2$

and not

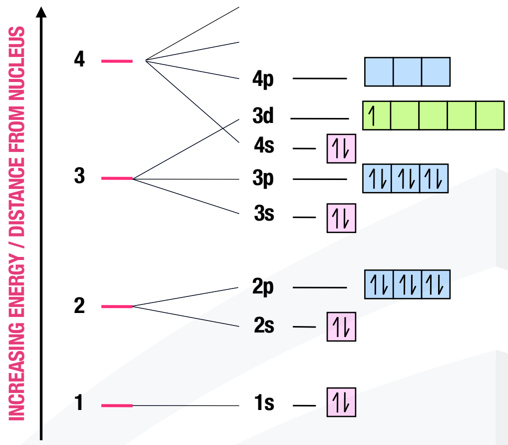
$1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^6$

Remember: 3d is higher than 4s in terms of energy levels!

72

SCANDIUM

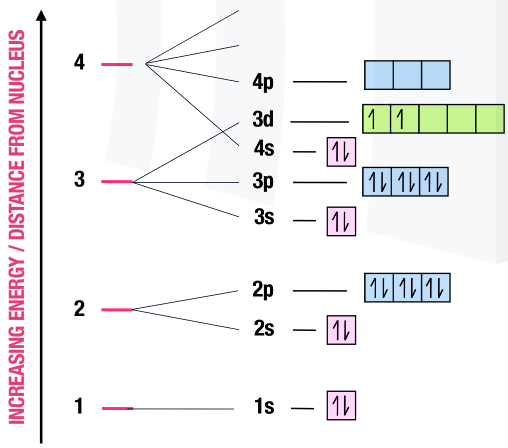
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1\ 4s^2$



With the lower energy 4s orbital filled, the next electrons can now fill the 3d orbitals. There are five d orbitals. They are filled according to Hund's Rule.

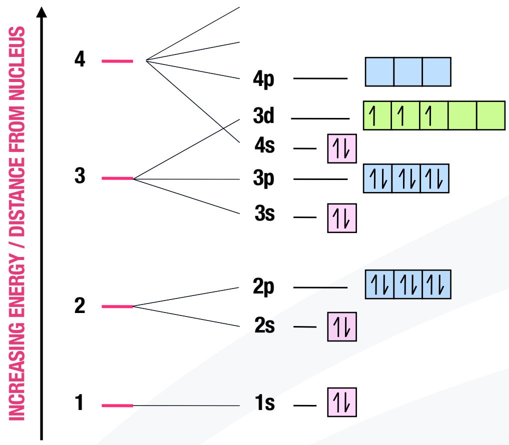
TITANIUM

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$



The 3d orbitals are filled according to Hund's rule so the next electron doesn't pair up but goes into an empty orbital in the same sub level.

VANADIUM

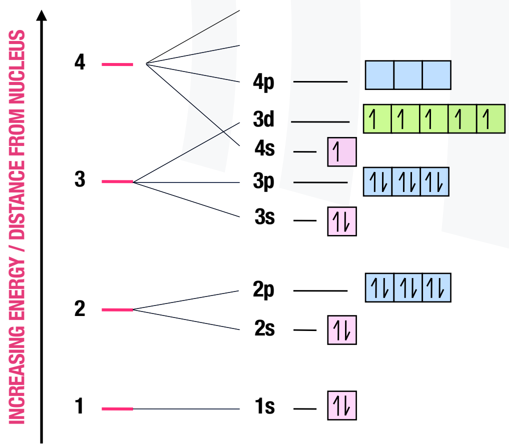


$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$

The 3d orbitals are filled according to Hund's rule so the next electron doesn't pair up but goes into an empty orbital in the same sub level.

75

CHROMIUM



$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^1$

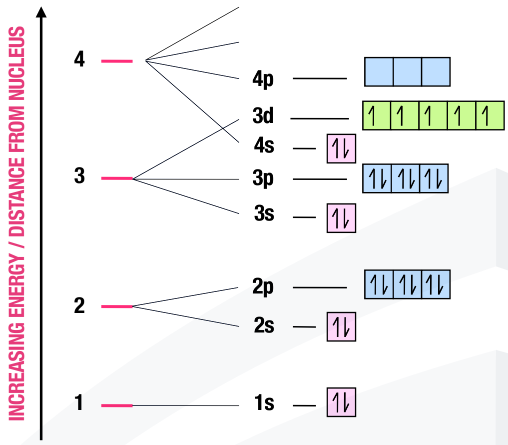
One would expect the configuration of chromium atoms to end in $3d^4$ $4s^2$.

To achieve a more stable arrangement of lower energy, one of the 4s electrons is promoted into the 3d to give six unpaired electrons with lower repulsion.

76

MANGANESE

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5\ 4s^2$

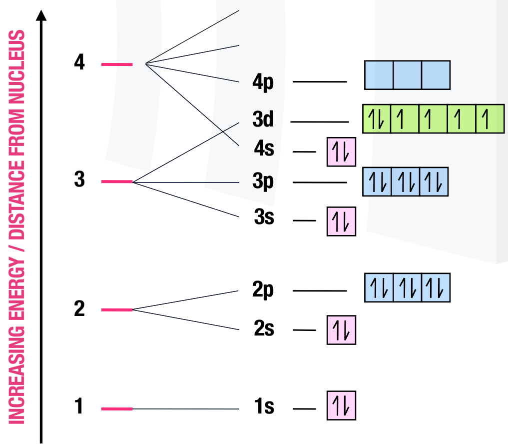


The new electron goes into the 4s to restore its filled state.

77

IRON

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6\ 4s^2$

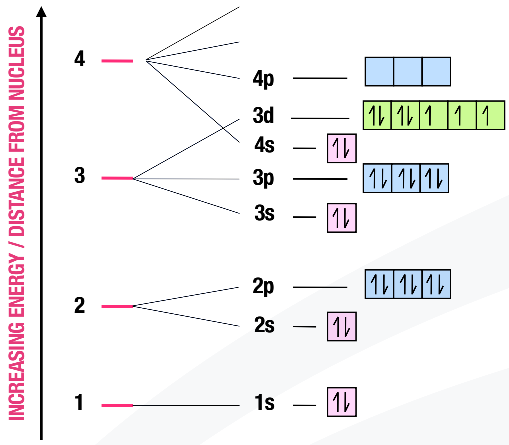


Orbitals are filled according to Hund's Rule. They continue to pair up.

78

COBALT

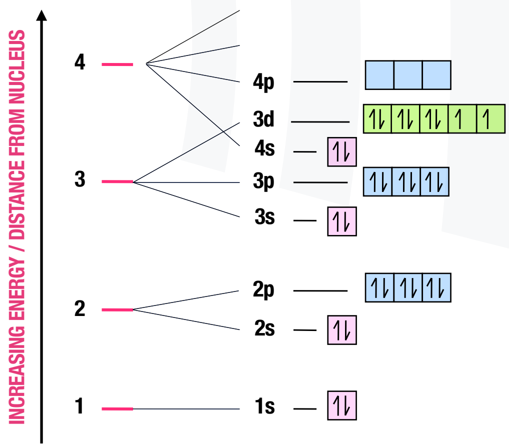
$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^7\ 4s^2$



Orbitals are filled according to Hund's Rule. They continue to pair up.

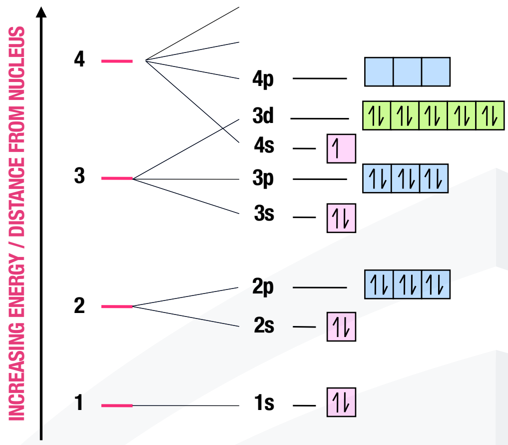
NICKEL

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^8\ 4s^2$



Orbitals are filled according to Hund's Rule. They continue to pair up.

COPPER

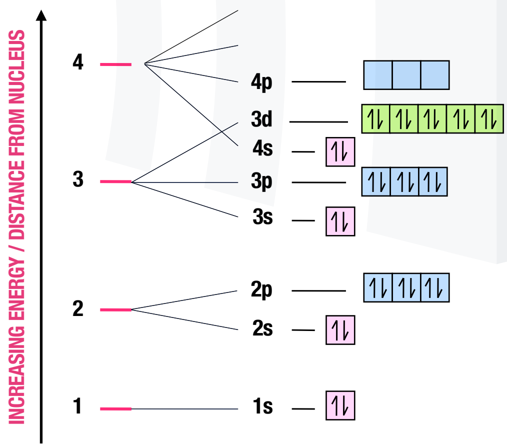


$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^1$

One would expect the configuration of chromium atoms to end in $3d^9$ $4s^2$.

To achieve a more stable arrangement of lower energy, one of the 4s electrons is promoted into the 3d.

ZINC



$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

The electron goes into the 4s to restore its filled state and complete the 3d and 4s orbital filling.

GALLIUM TO KRYPTON

The 4p orbitals are filled in exactly the same way as those in the 2p and 3p sub levels:

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2$

Ga - $4p^1$

Ge - $4p^2$

As - $4p^3$

Se - $4p^4$

Br - $4p^5$

Kr - $4p^6$

ELECTRONIC CONFIGURATION OF IONS

Metallic elements, belonging to Group I, II and III form positively charged ions by losing electrons in their outermost shell, so that the ions may be iso-electronic with the preceding noble gas.

SODIUM Na $1s^2\ 2s^2\ 2p^6\ 3s^1$

Na^+ $1s^2\ 2s^2\ 2p^6$

(1 electron removed from the 3s orbital)

ELECTRONIC CONFIGURATION OF IONS

Non-metallic elements belonging to Group V, VI and VII form negative ions by accepting electrons so that they achieve the noble gas configuration:

CHLORINE Cl $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Cl^- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

(1 electron added to the 3p orbital)

TRANSITION METAL IONS

When transition metals form ions, electrons are lost first from the outermost s orbital and then from the penultimate d sub-level.

Electrons in the 4s orbital are removed before any electrons in the 3d orbitals.

TITANIUM	Ti	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^2$
	Ti+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2\ 4s^1$
	Ti2+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^2$
	Ti3+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^1$
	Ti4+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$`

SKILL CHECK 17

Write electronic configurations for the following ions:

a Al^{3+}

b O^{2-}

c Fe^{3+}

d Cu^{2+}

e Cu^{+}

87

SKILL CHECK 18

Write the electron configuration for each ion.

a O^{2-}
b Br^{-}
c Sr^{2+}
d Co^{3+}

88

SKILL CHECK 19

Q. What is the order of increasing energy of the listed orbitals in the atom of titanium?

A 3s 3p 3d 4s

B 3s 3p 4s 3d

C 3s 4s 3p 3d

D 4s 3s 3p 3d

89

SKILL CHECK 20

A simple ion X^+ contains eight protons.

What is the electronic configuration of X^+?

A $1s^2$ $2s^1$ $2p^6$

B $1s^2$ $2s^2$ $2p^3$

C $1s^2$ $2s^2$ $2p^5$

D $1s^2$ $2s^2$ $2p^7$

90

SKILL CHECK 21

Write down the electronic configurations of the following ions:

A bromide

B magnesium

C iron(II)

D copper (II)

E iron(III)

SKILL CHECK 22

Q. Determine the number of unpaired electrons in each of these atoms:

A phosphorus

B chromium

C oxygen

SKILL CHECK 23

Use the periodic table to determine the element corresponding to each electron configuration.

A [Ar] $4s^2\ 3d^{10}\ 4p^6$
B [Ar] $4s^2\ 3d^2$
C [Kr] $5s^2\ 4d^{10}\ 5p^2$
D [Kr] $5s^2$

PERIODIC TABLE

The modern Periodic Table is arranged such that elements with similar electronic configuration lie in vertical groups.

Thus elements with electronic configuration ending up as ns^1 are put in one group while those that end up as ns^2 are put in another group.

The elements that fill up the s orbital of the highest energy level are said to belong to the s block, comprising of Groups 1 and 2.

PERIODIC TABLE

The elements that similarly fill up the p orbitals of the highest energy are said to belong to the p block, which comprises of Groups 13 to 0.

So there are two columns in the s block and six columns in the p block.

The elements that fill up the d sub-level of the penultimate shell are called the d block elements or the Transition Metal series.

SKILL CHECK 24

Write down the last term of the electronic configuration for the elements in the table, the sequence for Hydrogen has been entered for you. Label the s, p and d blocks on the table.

$1s^1$																	