Periodic trends

We need to explain the trends of **ionisation energy**, **atomic radii**, **ionic radii** and **electronegativity** across periods and down groups of the periodic table.

Ionisation Energy

What is ionisation energy?

The energy required to remove one electron from an atom in its gaseous state

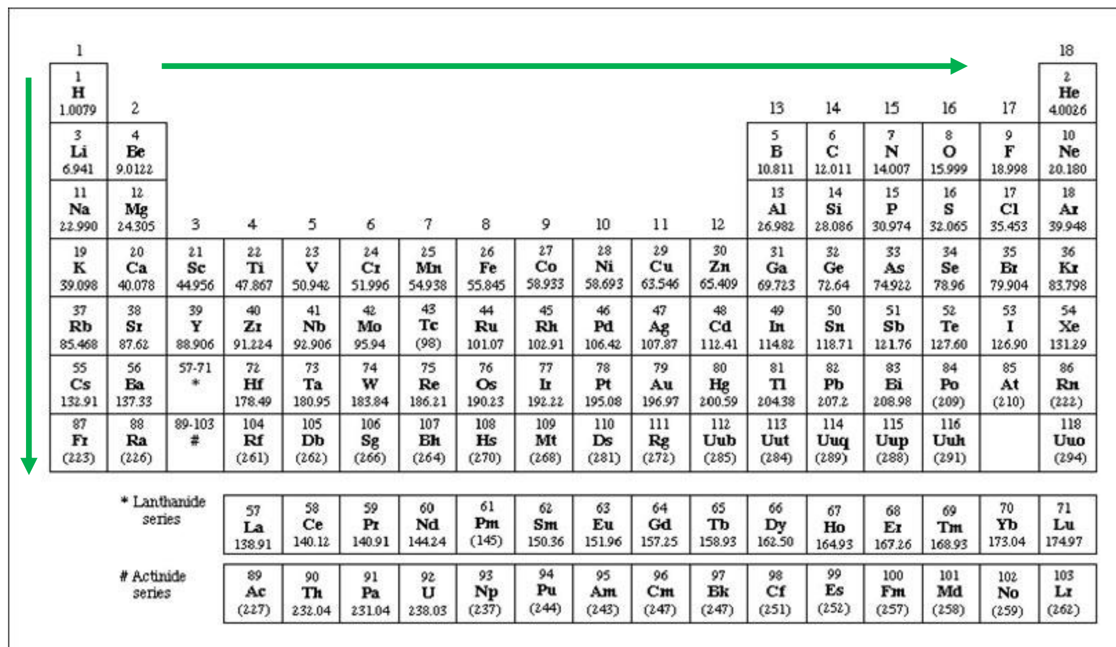
$Mg_{(g)} \rightarrow Mg_{(g)}^{+} + e$

How do you think it changes across a period? Down a group? Why?

Ionisation energy:

- <u>Increases</u> across a period
- <u>Decreases</u> down a group

Plot graph on pg 13

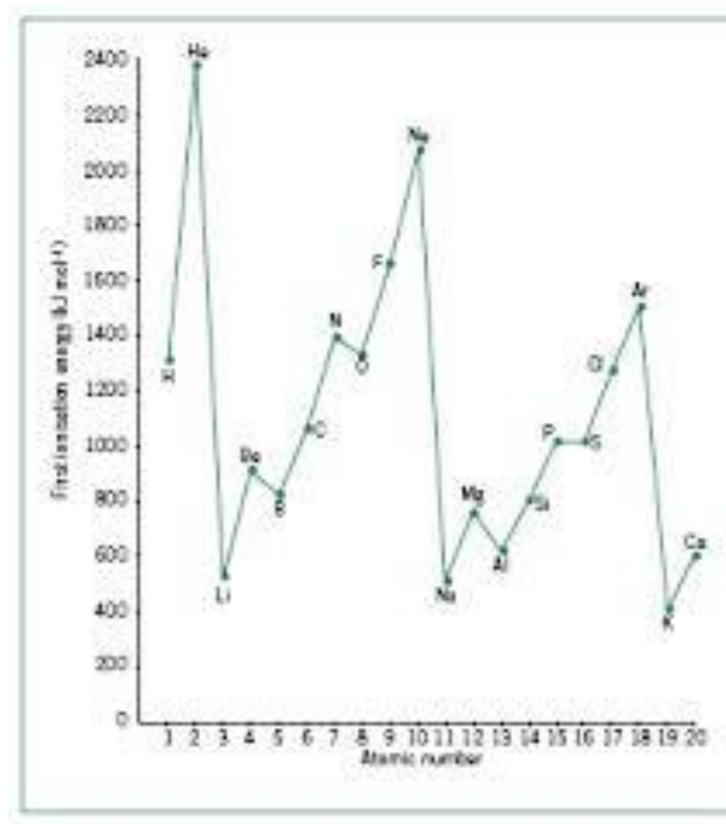


1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.0079																	2 He 4.0026
3 Li 6.941	4 Be 9.0122											5 B 10.811	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.180
11 Na 22.990	12 Mg 24.305											13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.065	17 Cl 35.453	18 Ar 39.948
19 K 39.098	20 Ca 40.078	21 Sc 44.956	22 Ti 47.867	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.845	27 Co 58.933	28 Ni 58.693	29 Cu 63.546	30 Zn 65.409	31 Ga 69.723	32 Ge 72.64	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.798
37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 *	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
87 Fr (223)	88 Ra (226)	89-103 #	104 Rf (261)	105 Db (262)	106 Sg (266)	107 Bh (264)	108 Hs (270)	109 Mt (268)	110 Ds (281)	111 Rg (272)	112 Uub (285)	113 Uut (284)	114 Uuq (289)	115 Uup (288)	116 Uuh (291)		118 Uuo (294)
		* Lanthanide series	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm (145)	62 Sm 150.36	63 Eu 151.96	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.04	71 Lu 174.97
		# Actinide series	89 Ac (227)	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np (237)	94 Pu (244)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (251)	99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)

Ionisation Energy

What affects ionisation energy? ie what affects how hard it is to remove an electron from an atom

- The energy level the electron is in (distance of orbital from the nucleus)
- Charge of the nucleus
- What sub orbital the electron is in



Electron configurations help to explain the ‘bumps’ in the graph

Shielding

We can use shielding to explain the decrease in ionisation energy going <u>down a group</u> of the periodic table.

You must explain what shielding is in your answer, not just state due to shielding the ionisation energy decreases.

When we move down a group of the periodic table the valence electrons are in higher energy levels. The higher the energy the further the level is from the nucleus. This means that there will be less attraction between the nucleus and the electron in the outside shells. This is called “shielding”.

Effective Nuclear Charge

We can use effective nuclear charge to explain the increase in ionisation energy going <u>across a row</u> of the periodic table.

You must explain what effective nuclear charge is in your answer, not just state due to increasing effective nuclear charge the ionisation energy increases.

When we move across a row of the periodic table electrons go into the same outside energy level. Moving right across a row the number of protons (and electrons) increases. The increase in the number of protons in the nucleus increases the attraction of the electrons in the outside shell to the nucleus. This is the “effective nuclear charge”.

Do now:

Complete the following table, this is from the 2012 exam Q1 a

Symbol	Electron configuration
Ge	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^2$ OR $[Ar] 4s^2 3d^{10} 4p^2$
Cu	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$ OR $[Ar] 4s^1 3d^{10}$
Cu^+	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$ OR $[Ar] 3d^{10}$

What does the term "effective nuclear charge" mean?

Ionisation Energy

Why does chlorine have a lower ionisation energy than fluorine?

What is the definition of the term.
Where are the atoms located on the periodic table (and valence electrons).
How this affects the stated trend.
Why it affects the stated trend and relate back to question.

Ionisation energy is the removal of one electron from the valence shell of an atom. Both chlorine and fluorine are in group 17. Chlorine has its valence electrons in the 3rd energy level and fluorine has its valence electrons in the 2nd energy level.

The further from the nucleus the electrons are (higher energy level) the more shielded they will be from the pull of the nucleus.

This weaker attraction for electrons in higher energy level **means that** the electrons will be easier to remove from the chlorine nucleus resulting in a lower ionisation energy.

Ionisation Energy

Why does carbon have a lower ionisation energy than fluorine?

What is the definition of the term.
Where are the atoms located on the periodic table (and valence electrons).
How this affects the stated trend.
Why it affects the stated trend and relate back to question.

Ionisation energy is the removal of one electron from the valence shell of an atom. Carbon and fluorine are both in the second row of the periodic table, this means they both have their valence electrons in the 2^{nd} energy level.

Fluorine atoms have more protons in their nucleus than carbon atoms. **This means** the effective nuclear charge acting on the fluorine electrons is greater than the effective nuclear charge on the carbon electrons.

This stronger attraction **means that** it will be harder to remove the electrons from the fluorine nucleus **resulting in** a higher ionisation energy.

2013 Sample Exam Q1 b (ii)

Complete worksheet about ionisation energy and exam questions

(b) Give a justification for each of the following:

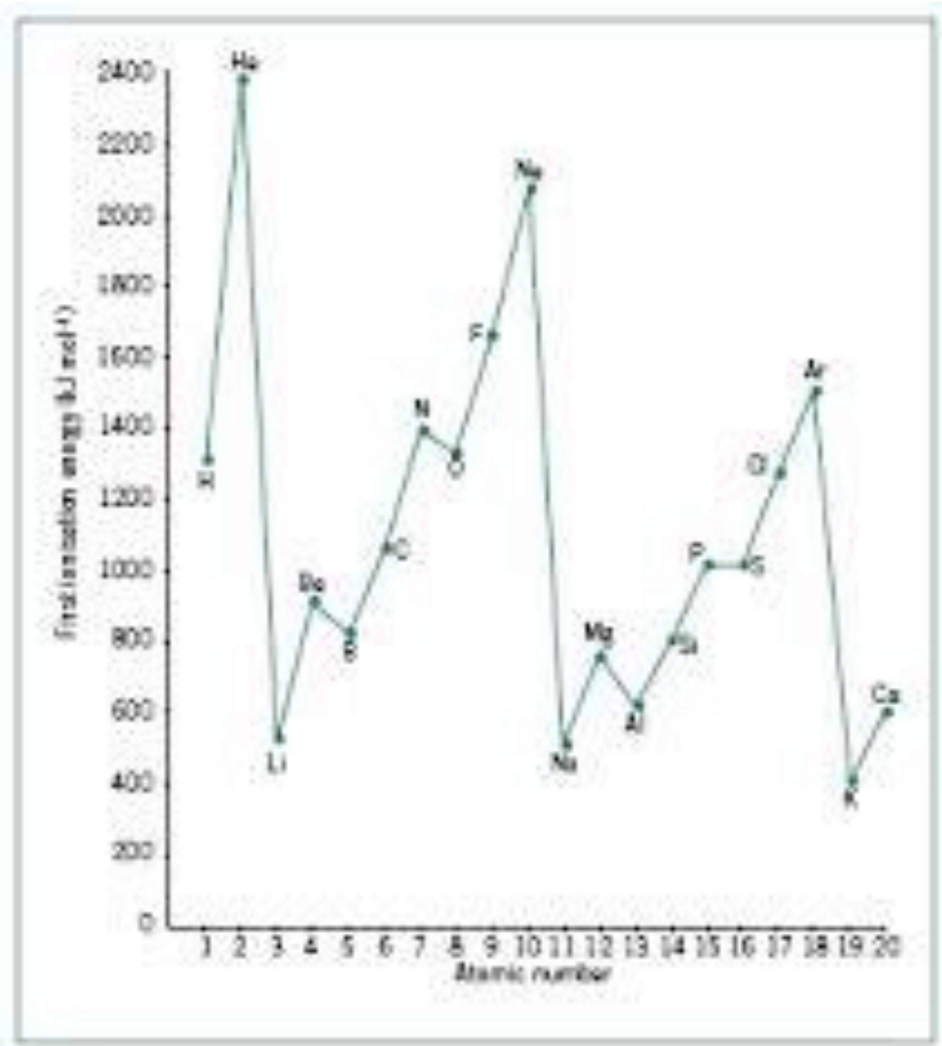
(ii) A chlorine atom has a greater first ionisation energy than a sodium atom.

2013 Sample Exam Q1 b (ii)

This is because 1st IE increases across a period. Across a period, the number of protons in the nucleus is increasing while electrons are being added to the same energy level so the effective nuclear charge on the outer electrons is increasing. This means that there is a stronger attractive force holding the outer electrons in chlorine than there is in sodium, so it takes more energy to remove the outermost electron from Cl than from Na so Cl has a higher 1st IE (energy required to remove outermost electron).

Ionisation energy

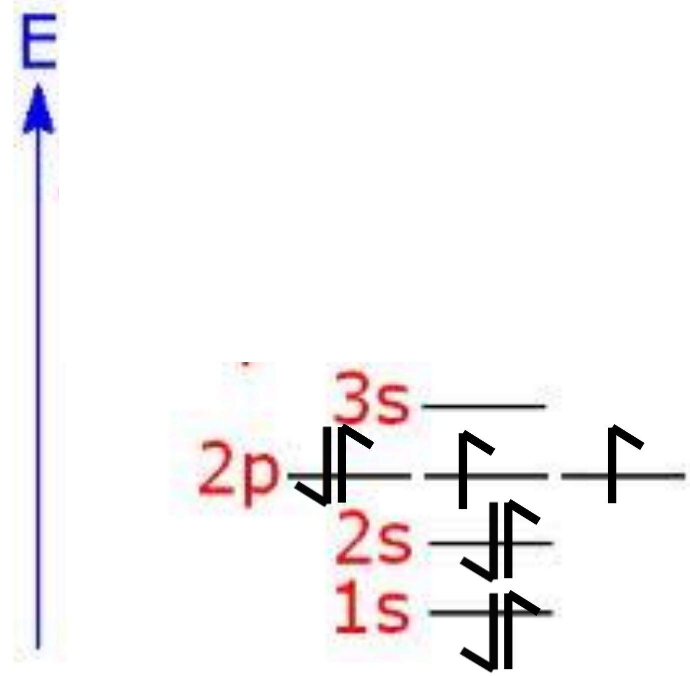
Why does oxygen have a lower ionisation energy than nitrogen?



Think back to electron configurations.

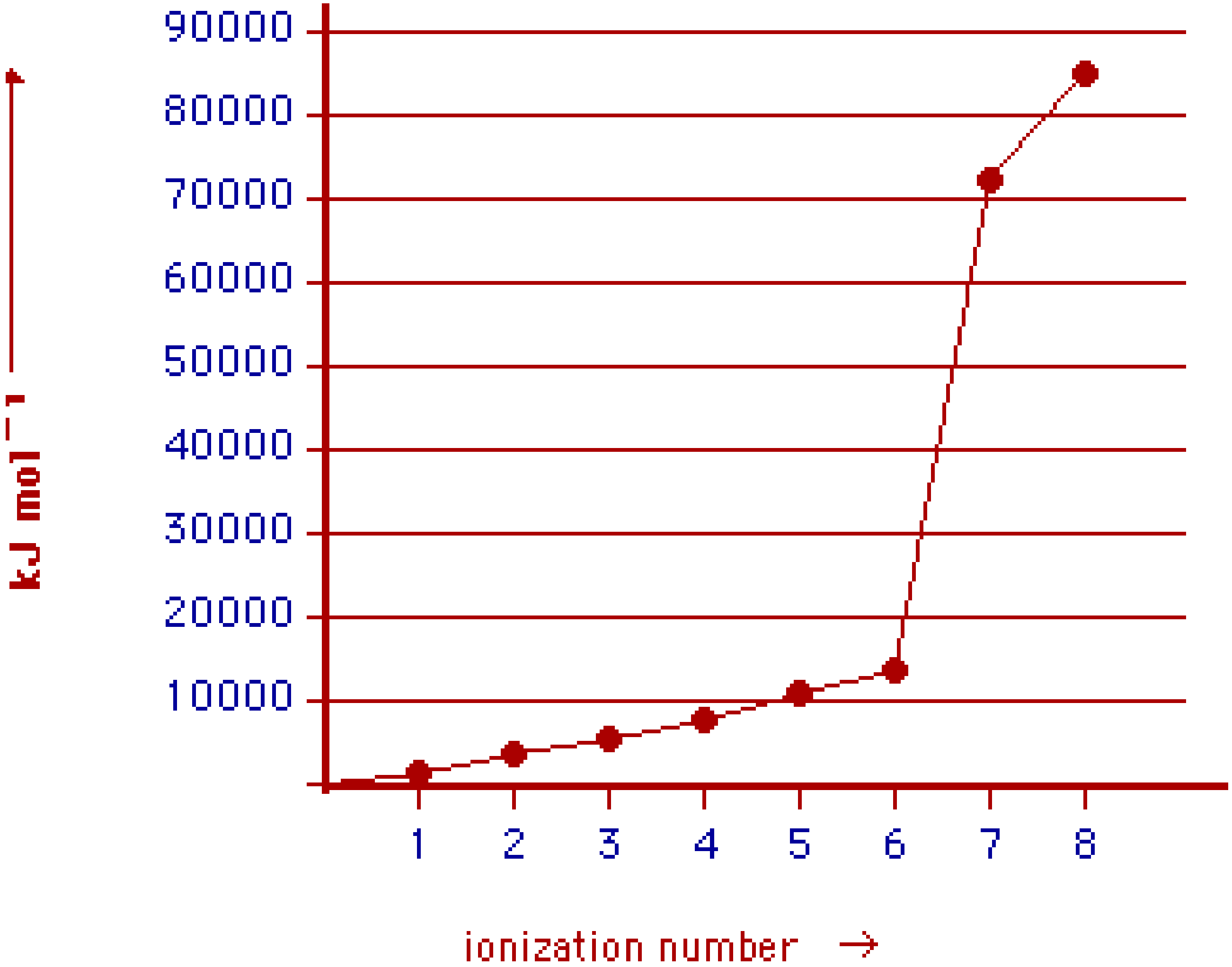
N: $1s^2 2s^2 2p^3$

O: $1s^2 2s^2 2p^4$



Successive ionisation energies

Removing electrons from oxygen ($1s^2 2s^2 2p^4$)



Gives us information about the electronic structure of the atom

Energy levels the electrons are in