

MSJChem

Tutorials for IB Chemistry

Topic 3 Periodicity SL

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Tutorials for IB Chemistry

The periodic table

The periodic table

Period – horizontal row

Group – vertical column

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.90
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.96	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
6	55 Cs 132.91	56 Ba 137.33	57 † La 138.91	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 207.20	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (281)	112 Cn (285)	113 Uut (286)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (294)	118 Uuo (294)
			†	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 173.05	71 Lu 174.97	
			‡	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)	

Atomic number
Element
Relative atomic mass

The periodic table

Metals, non-metals and metalloids

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01																	2 He 4.00
2	3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
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4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.90
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.96	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
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7	87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (281)	112 Cn (285)	113 Uut (286)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (294)	118 Uuo (294)
			†	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 173.05	71 Lu 174.97	
			‡	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)	

Atomic number
Element
Relative atomic mass

The periodic table

Group names in the periodic table

Group 1 : Alkali metals (Li, Na, K, Rb, Cs, Fr)

Group 2 : Alkaline Earth metals (Be, Mg, Ca, Sr, Ba)

Group 17 : Halogens (salt formers) (F, Cl, Br, I, At)

Group 18 : Noble gases (Ne, He, Kr, Xe, Rn)

Groups 3 – 12 : Transition metals

La – Lu : Lanthanoids (lanthanides)

Ac – Lr : Actinoids (actinides)

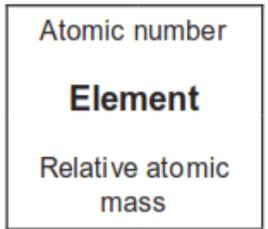
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**Electron configurations
and the periodic table**

Electron configurations

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H 1.01	s-block elements											p-block elements					2 He 4.00
2	3 Li 6.94	4 Be 9.01	d-block elements										5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31	d-block elements										13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
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7	87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)	104 Rf (267)	105 Db (268)	106 Sg (269)	107 Bh (270)	108 Hs (269)	109 Mt (278)	110 Ds (281)	111 Rg (281)	112 Cn (285)	113 Uut (286)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (294)	118 Uuo (294)



f-block elements

†
‡

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 173.05	71 Lu 174.97
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)

Electron configurations

Element	Block	Electron configuration
Na	s-block	$1s^2 2s^2 2p^6 3s^1$
O	p-block	$1s^2 2s^2 2p^4$
Co	d-block	$[\text{Ar}] 4s^2 3d^7$
Nd	f-block	$[\text{Xe}] 6s^2 4f^4$

Electron configurations

Group number	Block	Valence electron configuration	Number of valence electrons
1	s	ns^1	1
2	s	ns^2	2
13	p	$ns^2 np^1$	3
14	p	$ns^2 np^2$	4
15	p	$ns^2 np^3$	5
16	p	$ns^2 np^4$	6
17	p	$ns^2 np^5$	7
18	p	$ns^2 np^6$	8

Electron configurations

Lithium $Z = 3$

Electron configuration: $1s^2 2s^1$



Number of valence electrons (1)
Group 1

Highest occupied main energy level and period number

Block (s)

Electron configurations

Phosphorus $Z = 15$

Electron configuration: $[\text{Ne}] 3s^2 3p^3$



Number of valence electrons (5)
Group 15

Highest occupied main energy level and period number

Block (p)

Electron configurations

Strontium $Z = 38$

Electron configuration: $[\text{Kr}] 5s^2$



Number of valence
electrons (2)
Group 2

Highest occupied main energy
level and period number

Block (s)

Electron configurations

Fluorine $Z = 9$

Electron configuration: $[\text{He}] 2s^2 2p^5$



Number of valence electrons (7)
Group 17

Highest occupied main energy level and period number

Block (p)

Electron configurations

Krypton $Z = 36$

Electron configuration: $[\text{Ar}] 3d^{10} 4s^2 4p^6$



Number of valence electrons (8)
Group 18

Highest occupied main energy level and period number

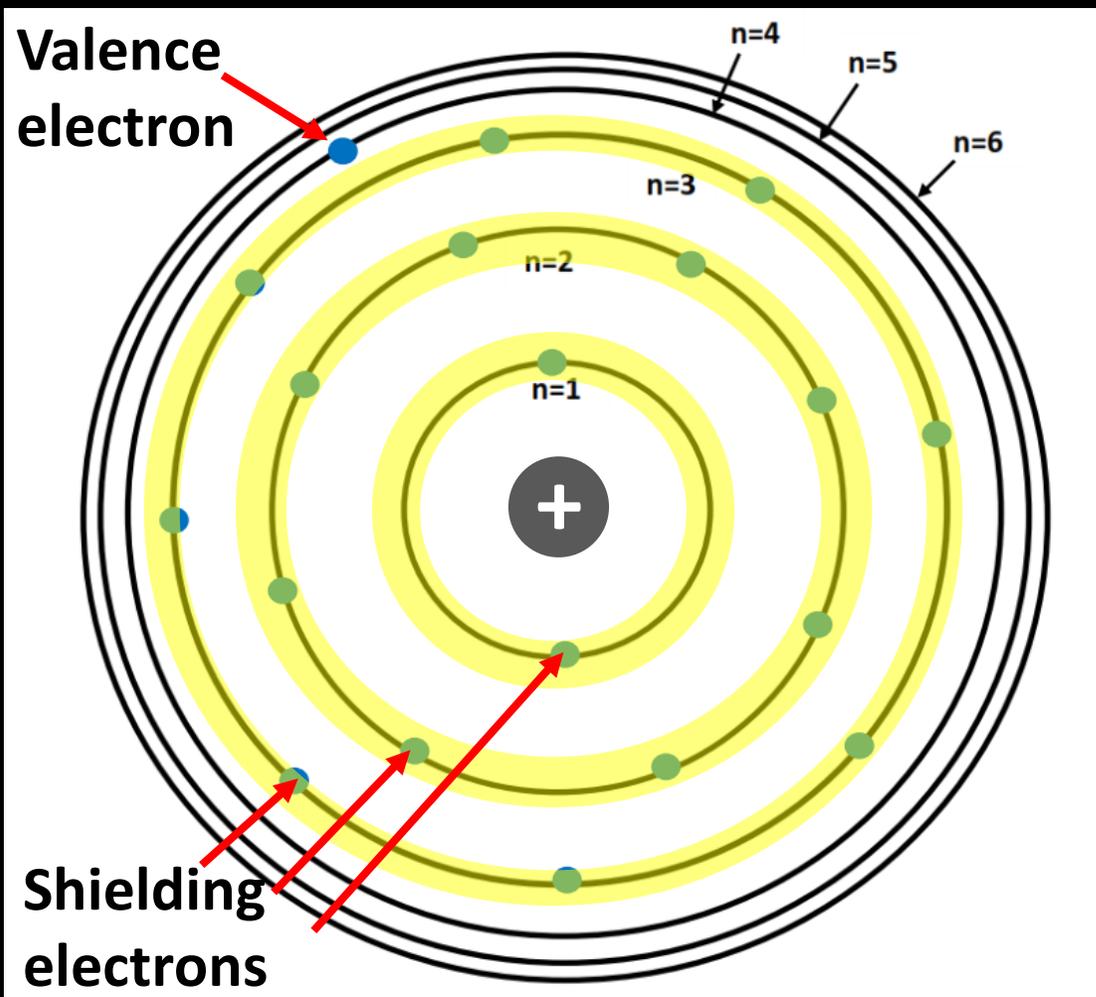
Block (p)

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**Electron shielding and
effective nuclear charge**

Electron shielding



Electron shielding occurs when the inner (shielding) electrons shield the outer (valence) electrons from the full attraction of the nucleus. The valence electron(s) within an atom require less energy to remove than the inner electrons.



Electron shielding

Electron shielding remains constant across a period (left to right).

Electron shielding increases down a group.

	1	2	3			13	14	15	16	17	18
1	1 H 1.01			Na	$1s^2 2s^2 2p^6 3s^1$						2 He 4.00
2	3 Li 6.94	4 Be 9.01		Mg	$1s^2 2s^2 2p^6 3s^2$	5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31		Al	$1s^2 2s^2 2p^6 3s^2 3p^1$	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	Si	$1s^2 2s^2 2p^6 3s^2 3p^2$	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.90
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	P	$1s^2 2s^2 2p^6 3s^2 3p^3$	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 † La 138.91	S	$1s^2 2s^2 2p^6 3s^2 3p^4$	81 Tl 204.38	82 Pb 207.20	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 ‡ Ac (227)	Cl	$1s^2 2s^2 2p^6 3s^2 3p^5$	113 Uut (286)	114 Uuq (289)	115 Uup (288)	116 Uuh (293)	117 Uus (294)	118 Uuo (294)
			†	Ar	$1s^2 2s^2 2p^6 3s^2 3p^6$	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97	
			‡			99 Es (252)	100 Fm (257)	101 Md (258)	102 No (259)	103 Lr (262)	

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Effective nuclear charge

Effective nuclear charge

Effective nuclear charge (Z_{eff}) is the net positive charge experienced by valence electrons.

$$Z_{\text{eff}} = Z - S$$

Z is the atomic number

S is the number of shielding electrons

Effective nuclear charge

Effective nuclear charge (Z_{eff}) increases across a period (left to right).

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

$$Z_{\text{eff}}(\text{Na}) = 11 - 10 = +1$$

$$Z_{\text{eff}}(\text{Mg}) = 12 - 10 = +2$$

$$Z_{\text{eff}}(\text{Al}) = 13 - 10 = +3$$

$$Z_{\text{eff}}(\text{Si}) = 14 - 10 = +4$$

$$Z_{\text{eff}}(\text{P}) = 15 - 10 = +5$$

$$Z_{\text{eff}}(\text{S}) = 16 - 10 = +6$$

$$Z_{\text{eff}}(\text{Cl}) = 17 - 10 = +7$$

$$Z_{\text{eff}}(\text{Ar}) = 18 - 10 = +8$$

Effective nuclear charge

Effective nuclear charge (Z_{eff}) remains the same down a group.



$$Z_{\text{eff}}(\text{Li}) = 3 - 2 = +1$$

$$Z_{\text{eff}}(\text{Na}) = 11 - 10 = +1$$

$$Z_{\text{eff}}(\text{K}) = 19 - 18 = +1$$

$$Z_{\text{eff}}(\text{Rb}) = 37 - 36 = +1$$

Summary

Electron shielding occurs when the inner shielding electrons shield the outer valence electrons from the full attraction of the nucleus.

Electron shielding remains the same across a period (left to right) and increases down a group.

Effective nuclear charge is the net positive charge felt by the valence electrons.

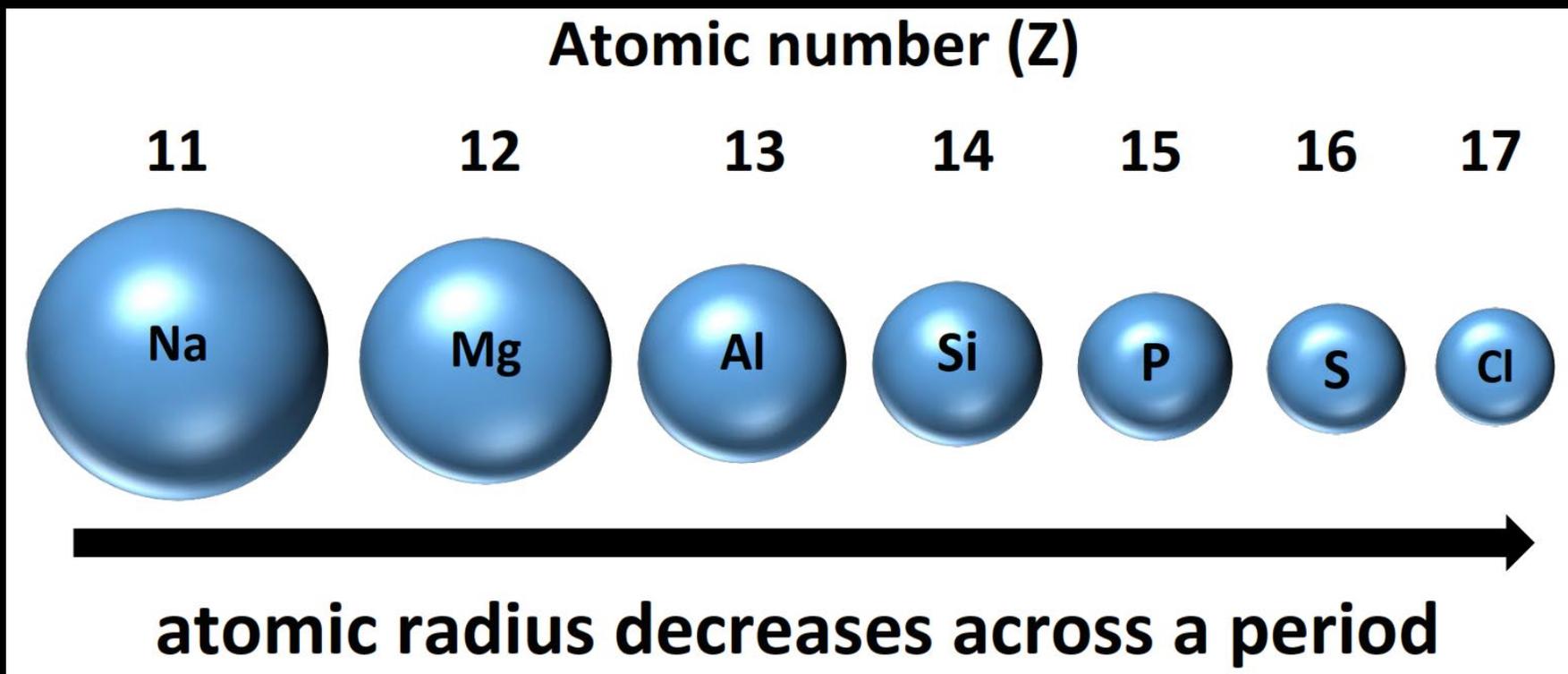
It increases across a period (left to right) and remains the same down a group.

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Tutorials for IB Chemistry

Trends in atomic radii

Trends in atomic radii



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	$1s^2$	$2s^2$	$2p^6$	$3s^2$	$3p^5$

Nuclear charge increases across a period.

Electron shielding remains constant across a period.

The attraction between the nucleus and the outer electrons increases which results in a decreasing atomic radius.

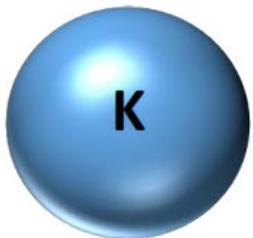
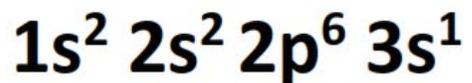
Trends in atomic radii



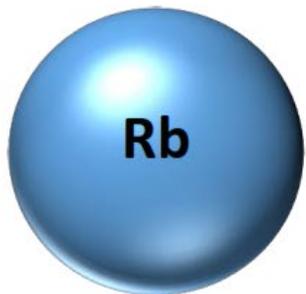
Li



Na



K



Rb



Atomic radius increases down a group as the number of occupied energy levels increases.

Trends in atomic radii

Atomic radius decreases across a period (left to right) because of increasing nuclear charge and the same amount of electron shielding.

Atomic radius increases down a group because of an increase in the number of occupied main energy levels.

9. Atomic and ionic radii of the elements

32 H		Atomic radius (10 ⁻¹² m)														37 He																			
Element		Element														Element																			
Ionic radius (10 ⁻¹² m)		Ionic radius (10 ⁻¹² m)														Ionic radius (10 ⁻¹² m)																			
130 Li 76 (1+)	99 Be 45 (2+)	84 B 27 (3+)	75 C 16 (4+)	71 N 146 (3-)	64 O 140 (2-)	60 F 133 (1-)	62 Ne	160 Na 102 (1+)	140 Mg 72 (2+)	124 Al 54 (3+)	114 Si 40 (4+)	109 P 38 (5+)	104 S 184 (2-)	100 Cl 181 (1-)	101 Ar	200 K 138 (1+)	174 Ca 100 (2+)	159 Sc 75 (3+)	148 Ti 86 (2+) 61 (4+)	144 V 79 (2+) 54 (5+)	130 Cr 62 (3+) 44 (6+)	129 Mn 83 (2+) 53 (4+)	124 Fe 61 (2+) 55 (3+)	118 Co 65 (2+) 55 (3+)	117 Ni 69 (2+)	122 Cu 77 (1+) 73 (2+)	120 Zn 74 (2+)	123 Ga 62 (3+)	120 Ge 53 (4+) 272 (4-)	120 As 58 (3+) 46 (5+)	118 Se 198 (2-)	117 Br 196 (1-)	116 Kr		
215 Rb 152 (1+)	190 Sr 118 (2+)	176 Y 90 (3+)	164 Zr 72 (4+)	156 Nb 72 (3+) 64 (5+)	146 Mo 65 (4+)	138 Tc 65 (4+)	136 Ru 68 (3+) 62 (4+)	134 Rh 67 (3+) 60 (4+)	130 Pd 86 (2+) 62 (4+)	136 Ag 115 (1+)	140 Cd 95 (2+)	142 In 80 (3+)	140 Sn 118 (2+) 69 (4+)	140 Sb 76 (3+)	137 Te 221 (2-)	136 I 220 (1-)	136 Xe	238 Cs 167 (1+)	206 Ba 135 (2+)	194 La 103 (3+)	164 Hf 71 (4+)	158 Ta 64 (5+)	150 W 66 (4+) 60 (6+)	141 Re 63 (4+) 53 (7+)	136 Os 63 (4+) 55 (6+)	132 Ir 68 (3+) 63 (4+)	130 Pt 80 (2+) 63 (4+)	130 Au 137 (1+) 85 (3+)	132 Hg 119 (1+) 102 (2+)	144 Tl 150 (1+) 89 (3+)	145 Pb 119 (2+) 78 (4+)	150 Bi 103 (3+) 76 (5+)	142 Po 97 (4+)	148 At	146 Rn
242 Fr	211 Ra	201 Ac																																	

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Trends in ionic radii

Trends in ionic radii

32 H																37 He	
130 Li 76 (1+)	99 Be 45 (2+)																62 Ne
160 Na 102 (1+)	140 Mg 72 (2+)																101 Ar
200 K 138 (1+)	174 Ca 100 (2+)	159 Sc 75 (3+)	148 Ti 86 (2+) 61 (4+)	144 V 79 (2+) 54 (5+)	130 Cr 62 (3+) 44 (6+)	129 Mn 83 (2+) 53 (4+)	124 Fe 61 (2+) 55 (3+)	118 Co 65 (2+) 55 (3+)	117 Ni 69 (2+)	122 Cu 77 (1+) 73 (2+)	120 Zn 74 (2+)	123 Ga 62 (3+)	120 Ge 53 (4+) 272 (4-)	120 As 58 (3+) 46 (5+)	118 Se 198 (2-)	117 Br 196 (1-)	116 Kr
215 Rb 152 (1+)	190 Sr 118 (2+)	176 Y 90 (3+)	164 Zr 72 (4+)	156 Nb 72 (3+) 64 (5+)	146 Mo 65 (4+)	138 Tc 65 (4+)	136 Ru 68 (3+) 62 (4+)	134 Rh 67 (3+) 60 (4+)	130 Pd 86 (2+) 62 (4+)	136 Ag 115 (1+)	140 Cd 95 (2+)	142 In 80 (3+)	140 Sn 118 (2+) 69 (4+)	140 Sb 76 (3+)	137 Te 221 (2-)	136 I 220 (1-)	136 Xe
238 Cs 167 (1+)	206 Ba 135 (2+)	194 La 103 (3+)	164 Hf 71 (4+)	158 Ta 64 (5+)	150 W 66 (4+) 60 (6+)	141 Re 63 (4+) 53 (7+)	136 Os 63 (4+) 55 (6+)	132 Ir 68 (3+) 63 (4+)	130 Pt 80 (2+) 63 (4+)	130 Au 137 (1+) 85 (3+)	132 Hg 119 (1+) 102 (2+)	144 Tl 150 (1+) 89 (3+)	145 Pb 119 (2+) 78 (4+)	150 Bi 103 (3+) 76 (5+)	142 Po 97 (4+)	148 At	146 Rn
242 Fr	211 Ra	201 Ac															

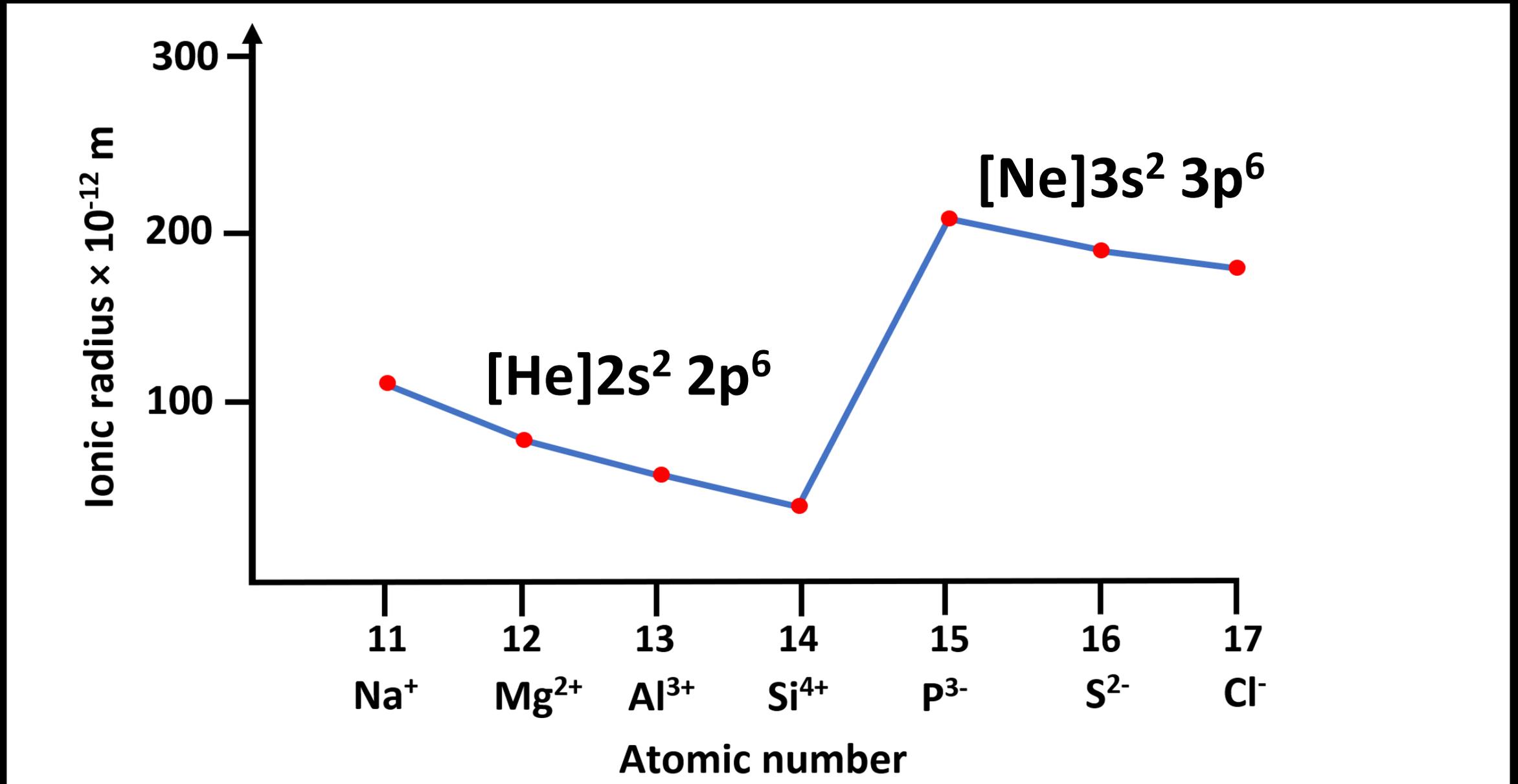
Atomic radius
(10^{-12} m)

Element

Ionic radius
(10^{-12} m)

109
P
212 (3-)

Trends in ionic radii



Trends in ionic radii

Ion	Atomic number	Electron configuration	Ionic radius ($\times 10^{-12}$ m)
Na^+	11	$1s^2 2s^2 2p^6$	102
Mg^{2+}	12	$1s^2 2s^2 2p^6$	72
Al^{3+}	13	$1s^2 2s^2 2p^6$	54

All three ions are isoelectronic (same electron configuration). The number of protons increases but the number of electrons remains the same.

The attraction between the nucleus and electrons increases, which causes the ionic radius to decrease.

Trends in ionic radii

Ion	Atomic number	Electron configuration	Ionic radius ($\times 10^{-12}$ m)
N^{3-}	7	$1s^2 2s^2 2p^6$	146
O^{2-}	8	$1s^2 2s^2 2p^6$	140
F^-	9	$1s^2 2s^2 2p^6$	133

All three ions are isoelectronic (same electron configuration). The number of protons increases but the number of electrons remains the same.

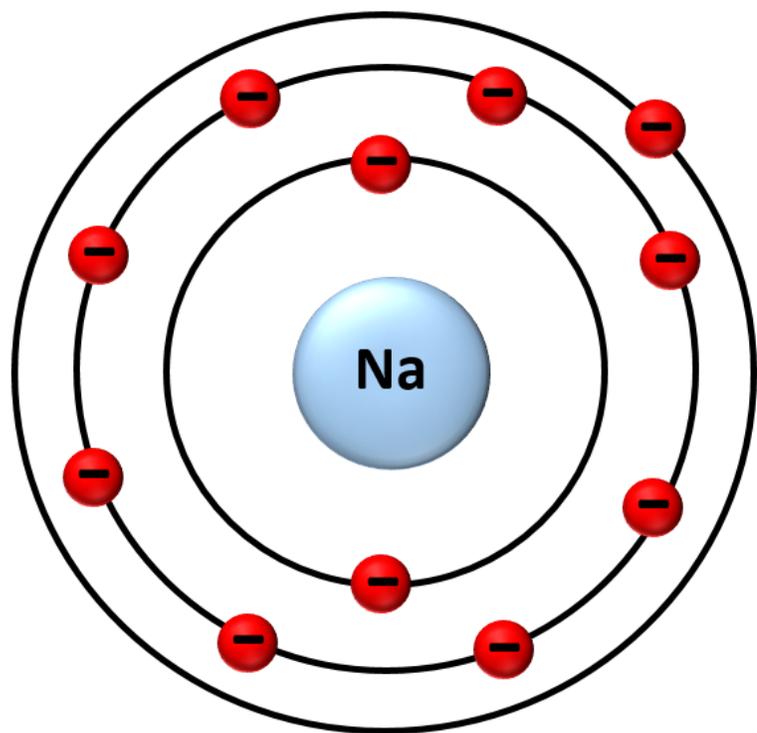
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Trends in ionic radii

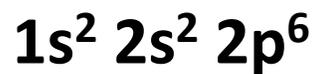
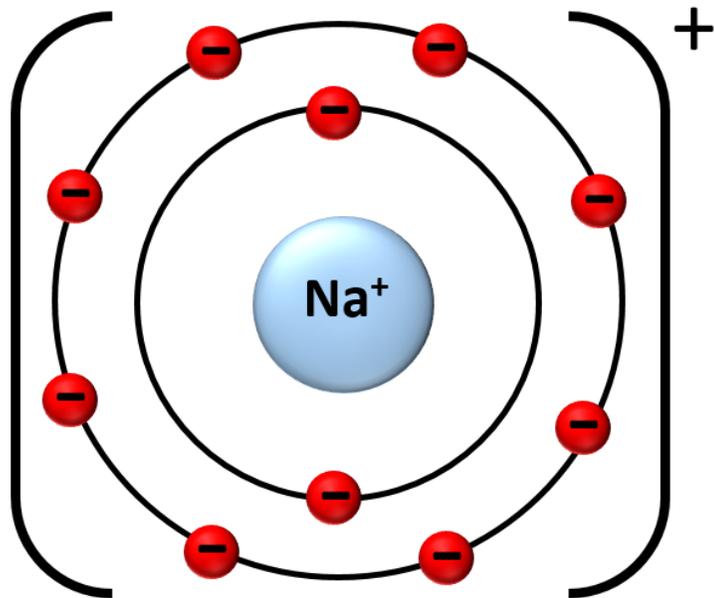
Ion	Atomic number	Electron configuration	Ionic radius ($\times 10^{-12}$ m)
N³⁻	7	1s² 2s² 2p⁶	146
O²⁻	8	1s² 2s² 2p⁶	140
F⁻	9	1s² 2s² 2p⁶	133
Na⁺	11	1s² 2s² 2p⁶	102
Mg²⁺	12	1s² 2s² 2p⁶	72
Al³⁺	13	1s² 2s² 2p⁶	54

Trends in ionic radii

sodium atom ($160 \times 10^{-12} \text{ m}$)



sodium ion ($102 \times 10^{-12} \text{ m}$)

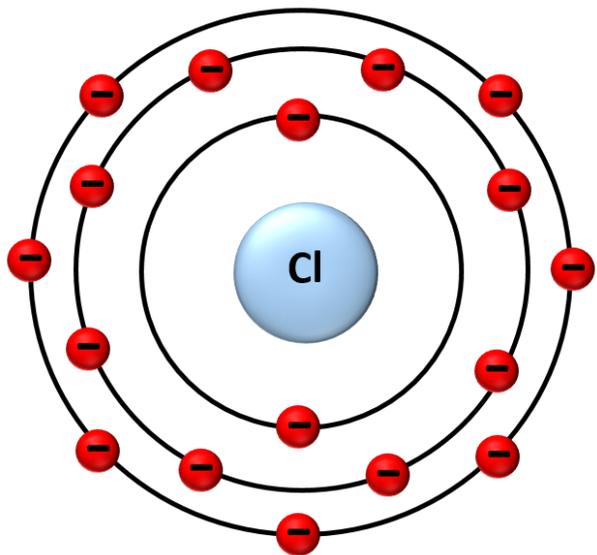


Positive ions lose electrons to obtain a full outer shell. Positive ions are smaller than their parent atoms.

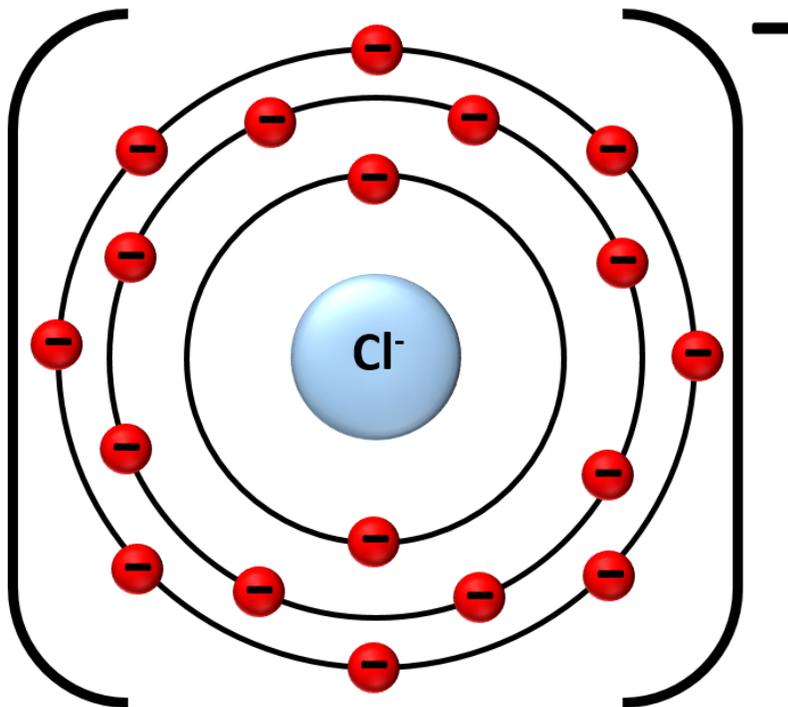
The ion has more protons than electrons, therefore, there is a stronger attraction between the nucleus and electrons.

Trends in ionic radii

chlorine atom ($100 \times 10^{-12} \text{ m}$)



chloride ion ($181 \times 10^{-12} \text{ m}$)



Negative ions gain electrons to obtain a full outer shell. Negative ions are bigger than their parent atoms.

The ion has more electrons than protons, therefore, there is a weaker attraction between the nucleus and electrons.

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**Trends in
ionisation energy**

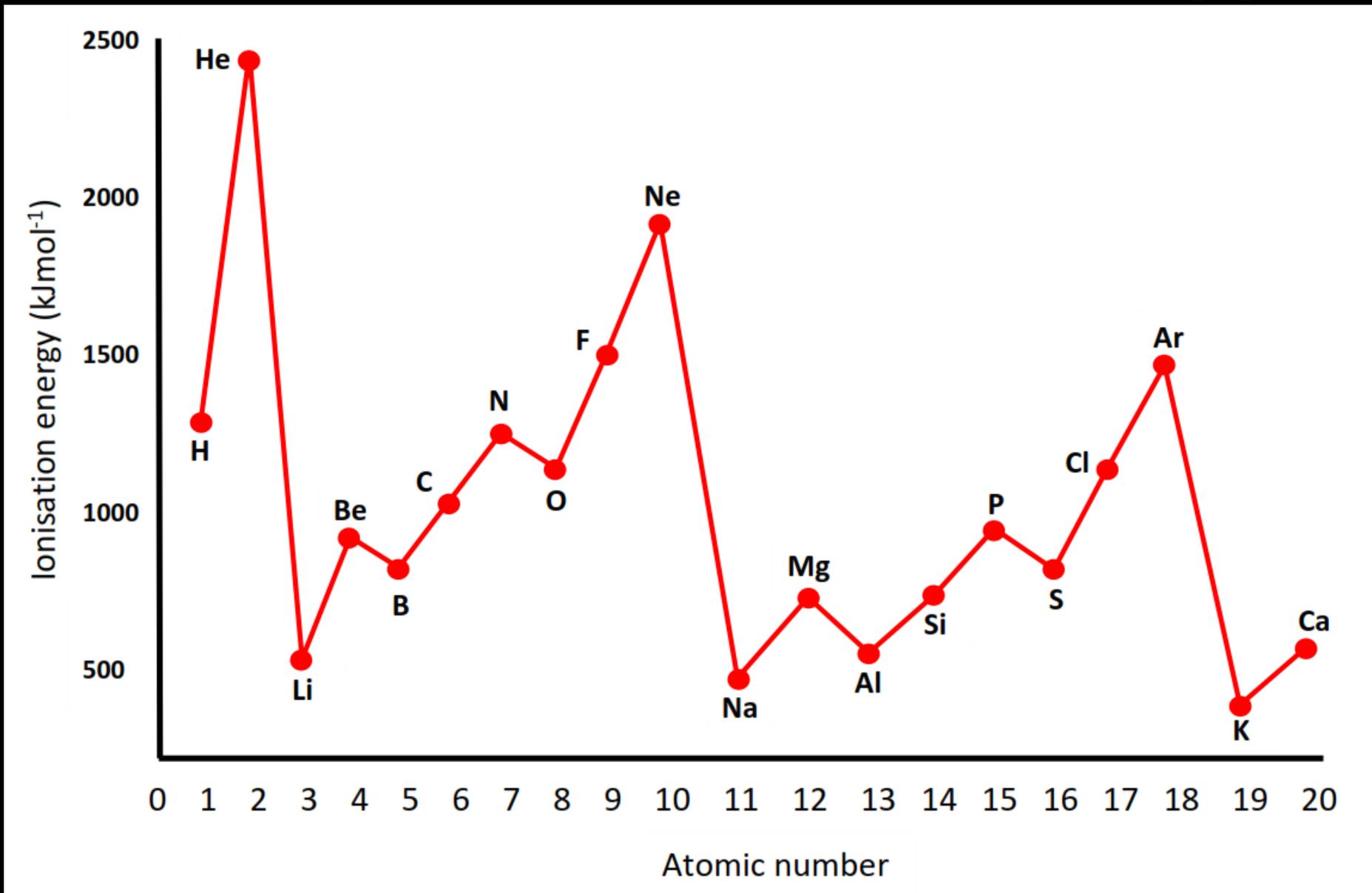
Ionisation energy

The first ionisation energy is the energy required to remove one mole of electrons from one mole of gaseous atoms to produce one mole of gaseous 1^+ ions.



First ionisation energy values are endothermic as energy is required to overcome the attraction between the positively charged nucleus and the outer electrons.

Ionisation energy



Ionisation energy

Ionisation energy increases across a period.

Nuclear charge increases and the atomic radius decreases across a period which means more energy is required to remove the outer electrons.

Ionisation energy decreases down a group.

The number of occupied energy levels increases down a group (increasing atomic radius) and increased electron shielding means less energy is required to remove the outer electrons.

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**Trends in
electronegativity**

Electronegativity

Electronegativity is a measure of the attraction of an atom for a bonding pair of electrons.

1312 -73		First ionization energy (kJ mol ⁻¹)																2372																			
H 2.2		Element																He																			
520 -60 900		Electronegativity																1681 -328 2081																			
Li 1.0		Be 1.6																B 2.0		C 2.6		N 3.0		O 3.4		F 4.0		Ne									
496 -53 738		Na 0.9		Mg 1.3																Al 1.6		Si 1.9		P 2.2		S 2.6		Cl 3.2		Ar							
419 -48 590 -2 633 -18 659 -8 651 -51 653 -64 717		K 0.8		Ca 1.0		Sc 1.4		Ti 1.5		V 1.6		Cr 1.7		Mn 1.6		Fe 1.8		Co 1.9		Ni 1.9		Cu 1.9		Zn 1.6		Ga 1.8		Ge 2.0		As 2.2		Se 2.6		Br 3.0		Kr	
403 -47 549 -5 600 -30 640 -41 652 -88 684 -72 702 -53 710 -101 720 -110 804 -54 731 -126 868		Rb 0.8		Sr 1.0		Y 1.2		Zr 1.3		Nb 1.6		Mo 2.2		Tc 2.1		Ru 2.2		Rh 2.3		Pd 2.2		Ag 1.9		Cd 1.7		In 1.8		Sn 2.0		Sb 2.0		Te 2.1		I 2.7		Xe 2.6	
376 -46 503 -14 538 -45 659 -1 728 -31 759 -79 756 -14 814 -106 865 -151 864 -205 890 -223 1007		Cs 0.8		Ba 0.9		La 1.1		Hf 1.3		Ta 1.5		W 1.7		Re 1.9		Os 2.2		Ir 2.2		Pt 2.2		Au 2.4		Hg 1.9		Tl 1.8		Pb 1.8		Bi 1.9		Po 2.0		At 2.2		Rn	
393 -47 509 -10 499 -34		Fr 0.7		Ra 0.9		Ac 1.1																															

Electronegativity

Electronegativity increases from left to right across a period for two reasons; the increase in nuclear charge and the decrease in atomic radius.

Electronegativity decreases down a group because of increasing atomic radius (bonding electrons are further from the attraction of the nucleus).

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**Trends in
metallic character**

Metallic character

The metallic character of an element can be defined as how easily an atom can lose electrons.

Metallic elements have low ionisation energies and tend to lose electrons to form positive ions.

Non-metal elements have higher ionisation energies and tend to gain electrons to form negative ions.

Metallic character

increasing metallic character



7 N 14.01
15 P 30.97
33 As 74.92
51 Sb 121.76
83 Bi 208.98

nitrogen – gaseous non-metal (3^- ions)

phosphorus – solid non-metal (3^- ions)

arsenic – metalloid

antimony – metalloid (more metallic)

bismuth – metal (2^+ ions)

Metallic character

Na, Mg and Al – metals
that form positive ions

P – non-metal that
forms 3⁻ ions

Cl – gaseous
non-metal that
forms 1⁻ ions

11 Na 22.99	12 Mg 24.31	13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45
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Si – shiny
metalloid

S – non-metal
that forms 2⁻ ions



decreasing metallic character

Metallic character

Metallic character increases down a group in the periodic table.

Increasing atomic radius results in a weaker attraction between the nucleus and valence electrons.

Metallic character decreases from left to right across a period in the periodic table.

Increasing nuclear charge (and decreasing atomic radius) results in a stronger attraction between the nucleus and valence electrons.

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**Electronegativity and
bonding**

Electronegativity is a measure of the attraction of an atom for a bonding pair of electrons.

It increases across a period due to increasing nuclear charge, and decreases down a group as atomic radius increases.

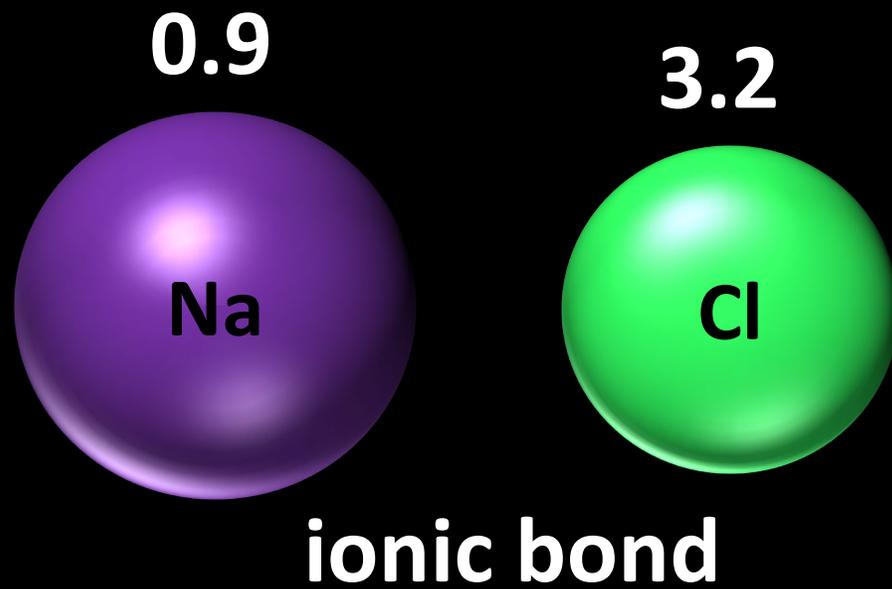
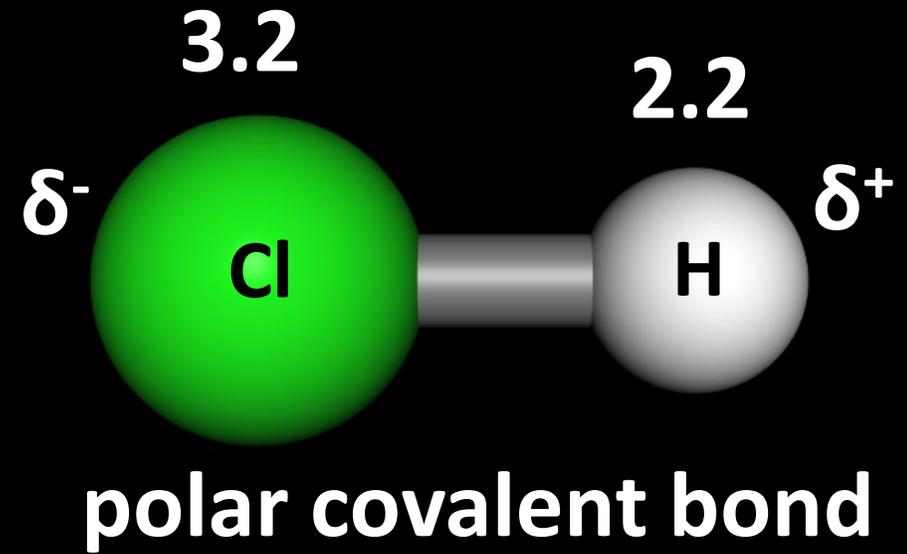
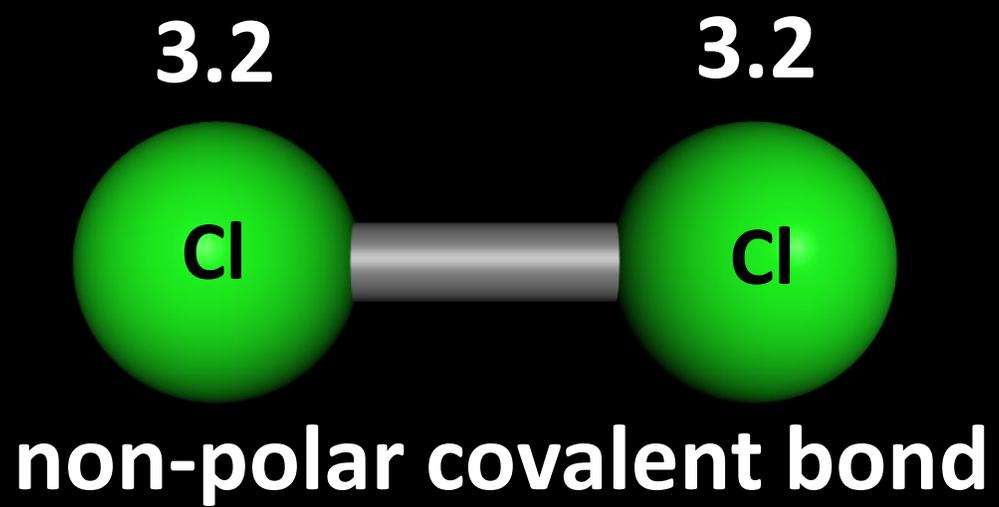
Difference in electronegativity	Type of bonding
0–0.4	non-polar covalent
0.5–1.7	polar covalent
≥ 1.8	ionic

Electronegativity and bonding

1312 -73																		2372	
H 2.2		First ionization energy (kJ mol ⁻¹) Electron affinity (kJ mol ⁻¹) (2nd EA / kJ mol ⁻¹) Element Electronegativity																He	
520 -60	900													801 -27	1086 -122	1402	1314 -141 (+753)	1681 -328	2081
Li 1.0	Be 1.6													B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne
496 -53	738													578 -42	787 -134	1012 -72	1000 -200 (+545)	1251 -349	1520
Na 0.9	Mg 1.3													Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar
419 -48	590 -2	633 -18	659 -8	651 -51	653 -64	717	762 -15	760 -64	737 -112	745 -119	906	579 -41	762 -119	944 -78	941 -195	1140 -325	1351		
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr		
403 -47	549 -5	600 -30	640 -41	652 -88	684 -72	702 -53	710 -101	720 -110	804 -54	731 -126	868	558 -29	709 -107	831 -101	869 -190	1008 -295	1170		
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.0	Te 2.1	I 2.7	Xe 2.6		
376 -46	503 -14	538 -45	659 -1	728 -31	759 -79	756 -14	814 -106	865 -151	864 -205	890 -223	1007	589 -36	716 -35	703 -91	812 -183	-270	1037		
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn		
393 -47	509 -10	499 -34																	
Fr 0.7	Ra 0.9	Ac 1.1																	

**Ionic bonding takes place between metals and non-metals.
Covalent bonding takes place between non-metals.**

Electronegativity and bonding



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**Bonding and acid-base
properties of the period 3
oxides**

Bonding

Formula of oxide	$\text{Na}_2\text{O}_{(s)}$	$\text{MgO}_{(s)}$	$\text{Al}_2\text{O}_{3(s)}$	$\text{SiO}_{2(s)}$	$\text{P}_4\text{O}_{10(s)}$ $\text{P}_4\text{O}_{6(s)}$	$\text{SO}_{3(l)}$ $\text{SO}_{2(g)}$	$\text{Cl}_2\text{O}_{7(l)}$ $\text{Cl}_2\text{O}_{(g)}$
Structure	ionic			giant covalent	molecular covalent		

Bonding changes from ionic to covalent across period 3. This change in bonding follows the decreasing difference in electronegativity between oxygen and the period 3 element.

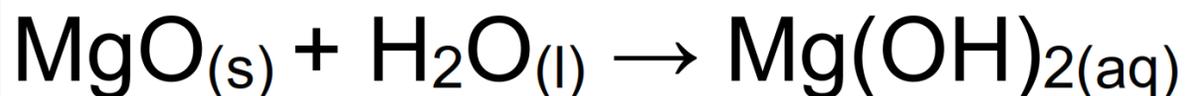
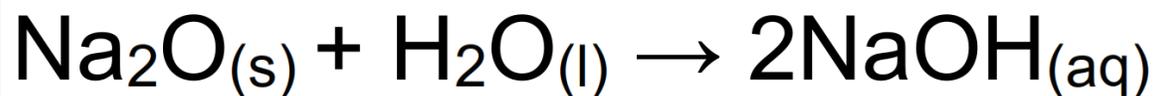
Na and O have a difference of 2.5 (ionic bond).

S and O have a difference of 0.8 (covalent bond).

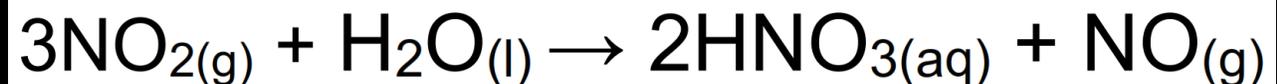
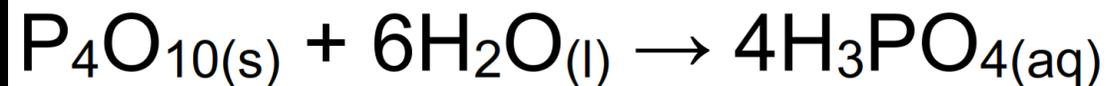
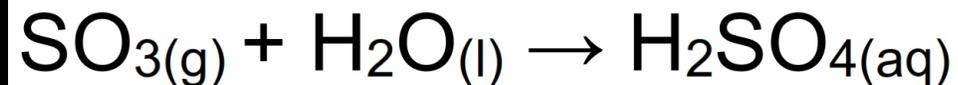
Acid-base properties

Formula of oxide	$\text{Na}_2\text{O}_{(s)}$	$\text{MgO}_{(s)}$	$\text{Al}_2\text{O}_{3(s)}$	$\text{SiO}_{2(s)}$	$\text{P}_4\text{O}_{10(s)}$ $\text{P}_4\text{O}_{6(s)}$	$\text{SO}_{3(l)}$ $\text{SO}_{2(g)}$	$\text{Cl}_2\text{O}_{7(l)}$ $\text{Cl}_2\text{O}_{(g)}$
Acid-base properties	basic		amphoteric	acidic			

basic oxides



acidic oxides

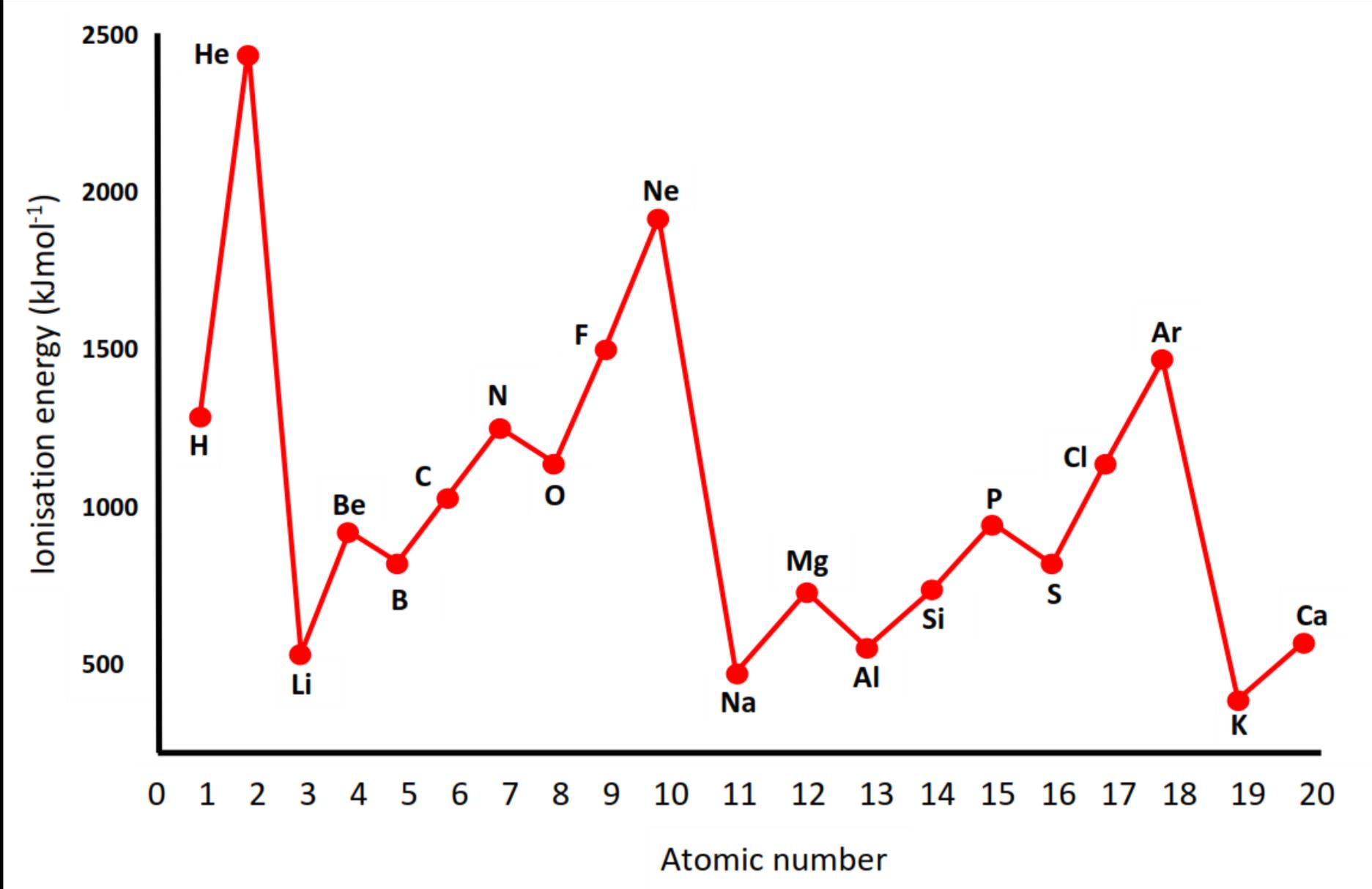


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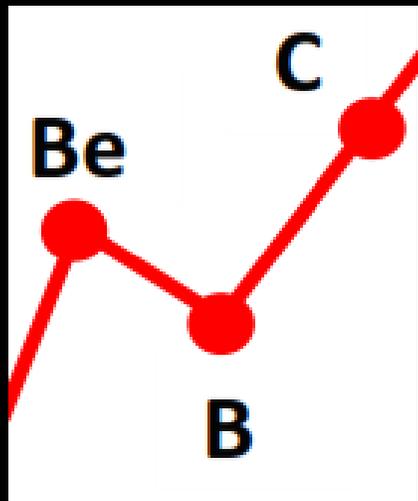
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**Exceptions to the trend in
ionisation energy**

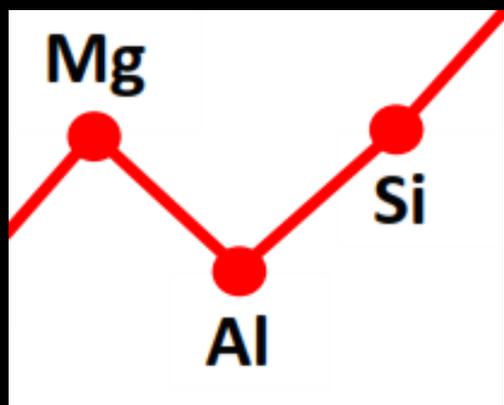
Ionisation energy



Ionisation energy



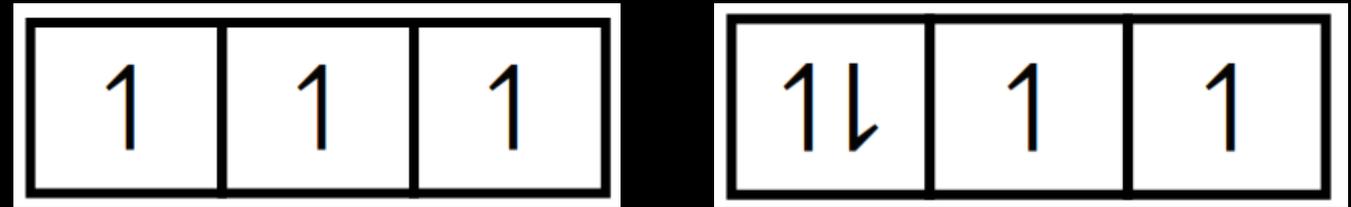
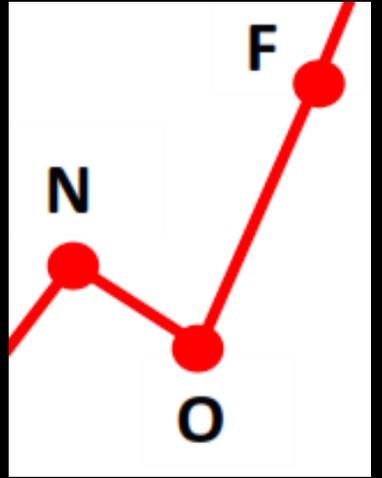
Be has the electron configuration $1s^2 2s^2$
B has the electron configuration $1s^2 2s^2 2p^1$
Electrons in p orbitals are slightly higher in energy and further from the nucleus than electrons in s orbitals, therefore they require less energy to remove.



A similar explanation can be applied for the decrease in ionisation energy from Mg to Al, except that the electron configurations are $1s^2 2s^2 2p^6 3s^2$ and $1s^2 2s^2 2p^6 3s^2 3p^1$

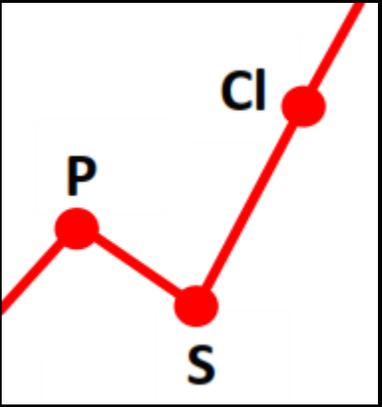
Ionisation energy

N has the electron configuration $1s^2 2s^2 2p^3$
O has the electron configuration $1s^2 2s^2 2p^4$



In oxygen, the electron is removed from a doubly occupied p orbital.

An electron in a doubly occupied orbital is repelled by the other electron and requires less energy to remove than an electron in a half-filled orbital.



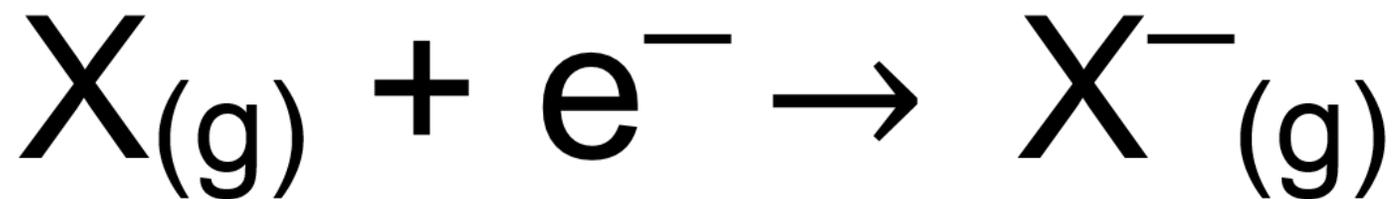
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Electron affinity

Electron affinity

The first electron affinity is the energy released when one mole of electrons is added to one mole of gaseous atoms to form one mole of 1^- ions.



The second electron affinity corresponds to the addition of one mole of electrons to one mole of gaseous 1^- ions.



Electron affinity

Electron affinity decreases (becomes less exothermic) down a group due to increasing atomic radius and increasing electron shielding.

1312 -73																		2372					
H 2.2		First ionization energy (kJ mol ⁻¹) Electron affinity (kJ mol ⁻¹) (2nd EA / kJ mol ⁻¹) Element Electronegativity																	He				
520 -60 900																		801 -27 1086 -122 1402		1314 -141 (+753) 1681 -328 2081			
Li 1.0	Be 1.6																	B 2.0	C 2.6	N 3.0	O 3.4	F 4.0	Ne
496 -53 738																		578 -42 787 -134 1012 -72		1000 -200 (+545) 1251 -349 1520			
Na 0.9	Mg 1.3																	Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2	Ar
419 -48 590 -2 633 -18 659 -8 651 -51 653 -64 717																		762 -15 760 -64 737 -112 745 -119 906		579 -41 762 -119 944 -78 941 -195 1140 -325 1351			
K 0.8	Ca 1.0	Sc 1.4	Ti 1.5	V 1.6	Cr 1.7	Mn 1.6	Fe 1.8	Co 1.9	Ni 1.9	Cu 1.9	Zn 1.6	Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0	Kr						
403 -47 549 -5 600 -30 640 -41 652 -88 684 -72 702 -53 710 -101 720 -110 804 -54 731 -126 868																		558 -29 709 -107 831 -101 869 -190 1008 -295 1170					
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.3	Nb 1.6	Mo 2.2	Tc 2.1	Ru 2.2	Rh 2.3	Pd 2.2	Ag 1.9	Cd 1.7	In 1.8	Sn 2.0	Sb 2.0	Te 2.1	I 2.7	Xe 2.6						
376 -46 503 -14 538 -45 659 -1 728 -31 759 -79 756 -14 814 -106 865 -151 864 -205 890 -223 1007																		589 -36 716 -35 703 -91 812 -183 -270 1037					
Cs 0.8	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn						
393 -47 509 -10 499 -34																							
Fr 0.7	Ra 0.9	Ac 1.1																					

Electron affinity

First electron affinity values are closely related to atomic radius and electron shielding.

In general, the greater the atomic radius and the greater the electron shielding, the less energy is released when an electron is added.

Second electron affinity values are positive due to the extra repulsion when adding negative electrons to an already negative ion.

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Group 1 metals

Group 1 metals

The group 1 metals (alkali metals) are found in group 1 of the periodic table.

1 H 1.01
3 Li 6.94
11 Na 22.99
19 K 39.10
37 Rb 85.47
55 Cs 132.91
87 Fr (223)

Element	Symbol	Electron configuration
Lithium	Li	[He] 2s ¹
Sodium	Na	[Ne] 3s ¹
Potassium	K	[Ar] 4s ¹
Rubidium	Rb	[Kr] 5s ¹
Caesium	Cs	[Xe] 6s ¹
Francium	Fr	[Rn] 7s ¹

Group 1 metals

3 Li 6.94
11 Na 22.99
19 K 39.10
37 Rb 85.47
55 Cs 132.91
87 Fr (223)

Group 1 metals are soft shiny metals that can be easily cut.

The melting point decreases down the group as the metallic bond gets progressively weaker.

Atomic and ionic radii increase down the group due to the increasing number of occupied energy levels.

Lithium, sodium and potassium float on water due to their low densities ($<1 \text{ g cm}^{-3}$).

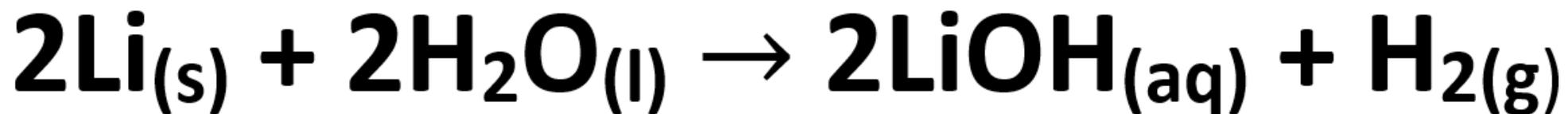
Group 1 metals

3 Li 6.94
11 Na 22.99
19 K 39.10
37 Rb 85.47
55 Cs 132.91
87 Fr (223)

Group 1 metals are stored in oil to prevent the reaction with oxygen in the air.

The reactivity of the group 1 metals increases down the group.

They react vigorously with water to produce an alkaline solution and hydrogen gas.



Group 1 metals

3 Li 6.94
11 Na 22.99
19 K 39.10
37 Rb 85.47
55 Cs 132.91
87 Fr (223)

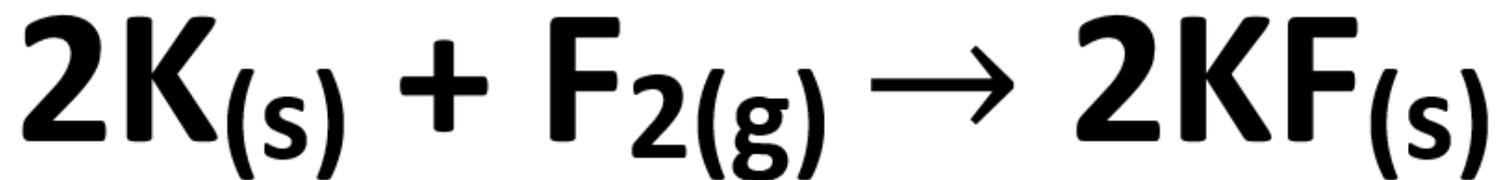
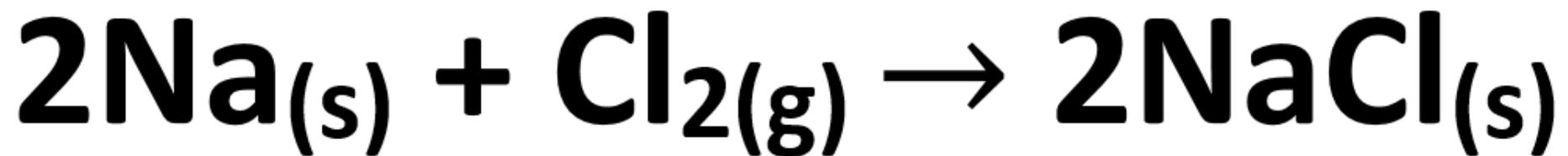
Ionisation energy decreases down the group as atomic radius increases which results in a weaker attraction between the nucleus and valence electron.

Metallic character increases down the group.

Electronegativity decreases down the group due to increasing atomic radius.

Group 1 metals

Group 1 metals react with group 17 elements (the halogens) to produce salts.



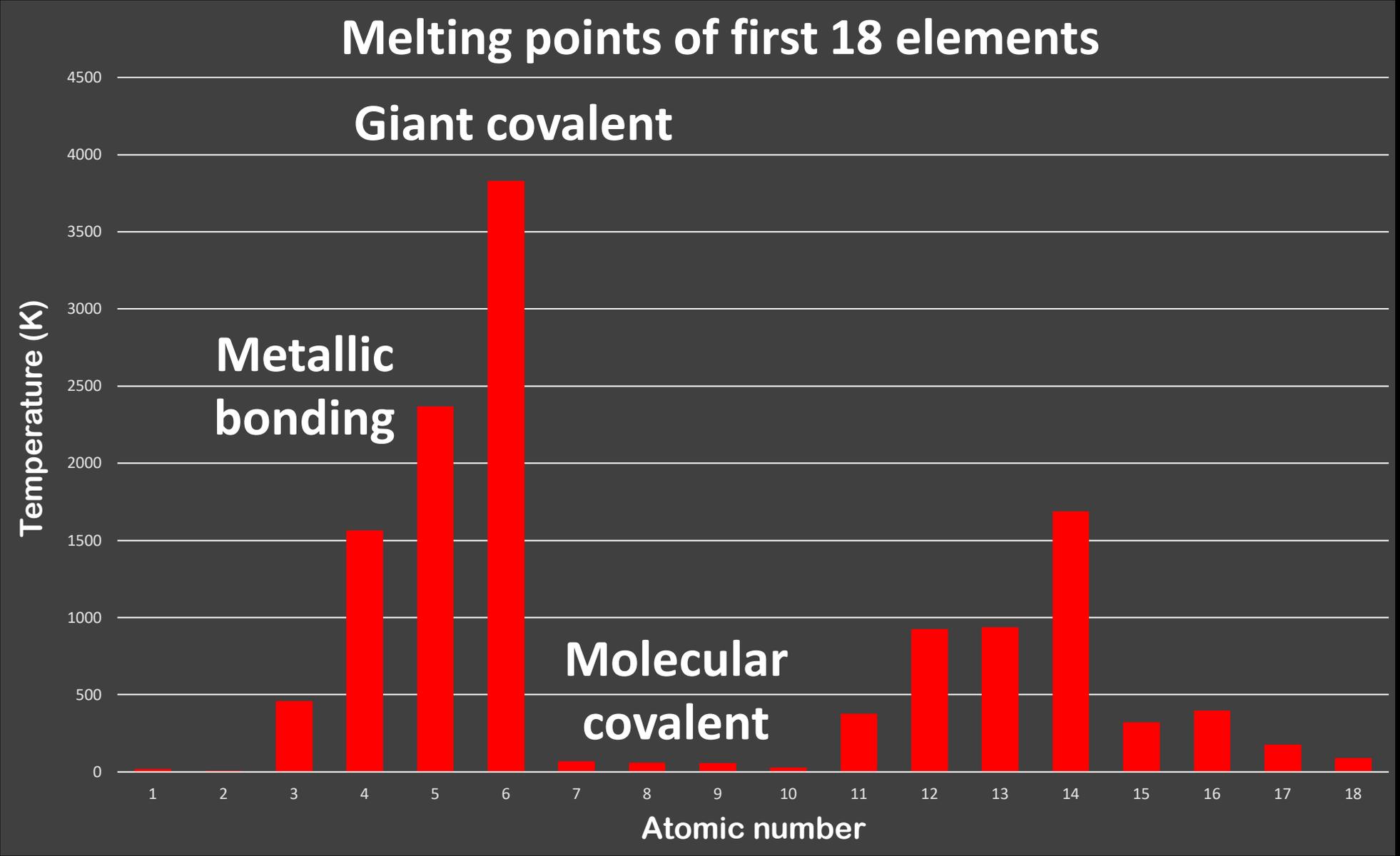
An ionic bond is formed between the elements in the compound.

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Trends in melting point

Trends in melting point



Trends in melting point

Melting point depends on the type of bonding (covalent, ionic or metallic) and structure (ionic lattice, molecular covalent, giant covalent, or metallic).

Melting point increases for metals in periods 2 and 3 as the strength of the metallic bond increases.

Melting point reaches a peak at C and Si which have giant covalent structures.

Melting point then decreases for elements that have molecular covalent structures.

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Group 17 elements

Group 17 elements

Group 17 elements (the halogens) are located in group 17 of the periodic table.

9 F 19.00
17 Cl 35.45
35 Br 79.90
53 I 126.90
85 At (210)
117 Uus (294)

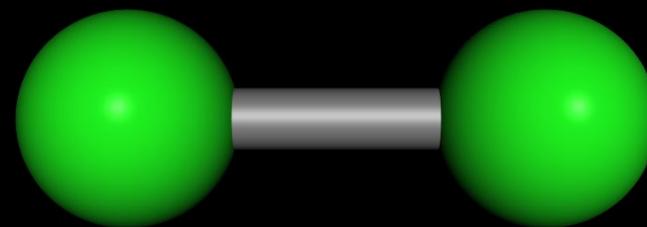
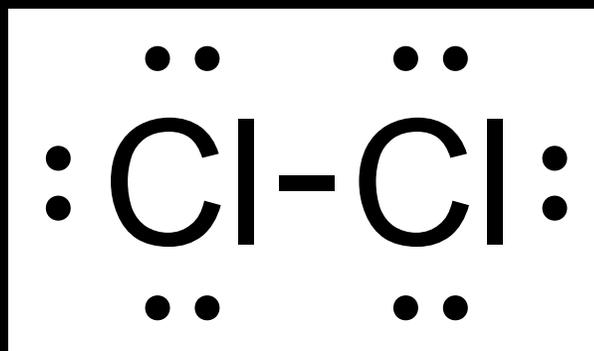
Element	Symbol	Electron configuration
Fluorine	F	[He] 2s ² 2p ⁵
Chlorine	Cl	[Ne] 3s ² 3p ⁵
Bromine	Br	[Ar] 4s ² 4p ⁵
Iodine	I	[Kr] 5s ² 5p ⁵
Astatine	As	[Xe] 6s ² 6p ⁵

Group 17 elements



The group 17 elements are coloured; fluorine is a pale yellow gas, chlorine a greenish-yellow gas, bromine a red liquid and iodine a purple solid.

They exist as diatomic molecules; F_2 , Cl_2 , Br_2 , I_2 (two atoms bonded together).



Group 17 elements

The melting points and boiling points increase down the group due to increasing molar mass which results in stronger London dispersion forces between the molecules.

-219.7 F -188.1
-101.5 Cl -34.04
-7.050 Br 58.78
113.7 I 184.4
301.8 At 336.8

Element	Molar mass (g mol ⁻¹)	Melting point (°C)	Boiling point (°C)
F ₂	38.00	-220	-188
Cl ₂	70.90	-102	-34
Br ₂	159.80	-7.1	59
I ₂	253.80	114	184

Group 17 elements

1681	-328
F	4.0
1251	-349
Cl	3.2
1140	-325
Br	3.0
1008	-295
I	2.7
	-270
At	2.2

Atomic and ionic radii increase down the group as the number of occupied energy levels increases.

Electronegativity decreases down the group as atomic radius increases.

Ionisation energy decreases down the group.
Electron affinity becomes less exothermic down the group.

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**Displacement reactions of
the group 17 elements**

Displacement reactions

9 F 19.00
17 Cl 35.45
35 Br 79.90
53 I 126.90

Elements at the top of the group are stronger oxidising agents.

An oxidising agent oxidises another species, it is reduced in the process.

The more reactive halogen displaces the ions of the less reactive halogen from solution.

Group 17 elements

