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Structure 3.1

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Groups and periods of the periodic table

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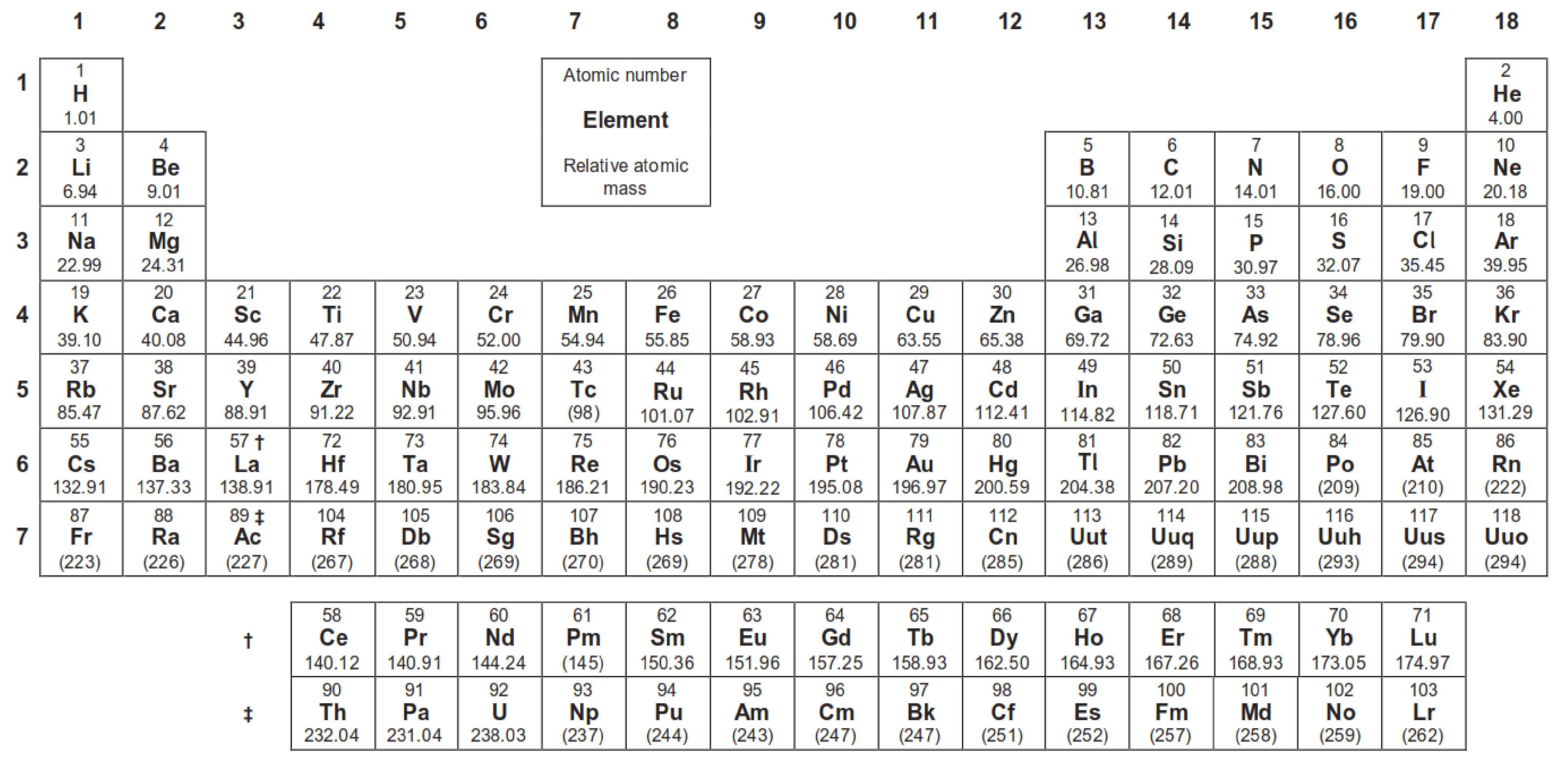
The periodic table

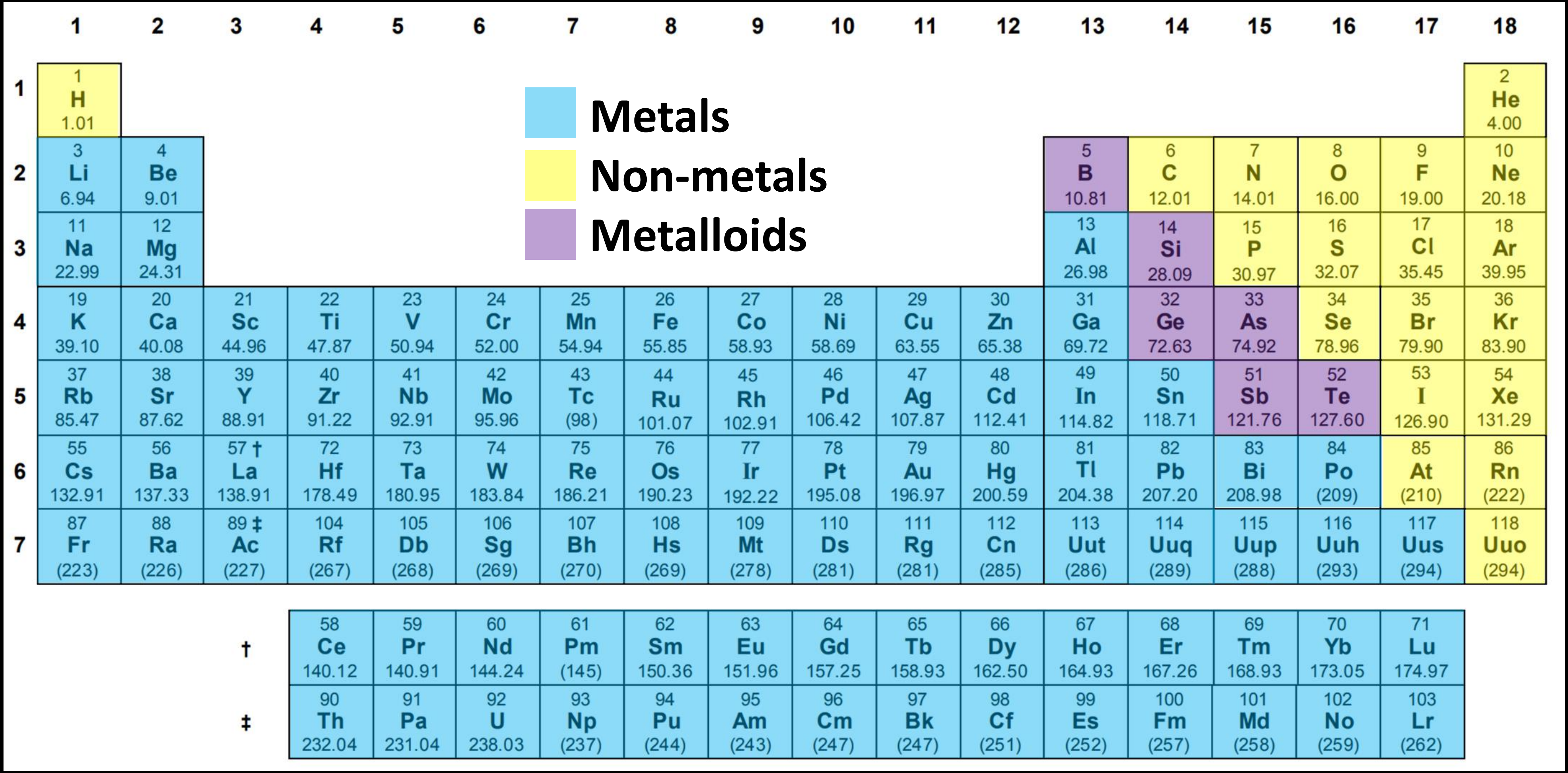
Period – horizontal row

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						Atomic number **Element** Relative atomic mass											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)	

Metals and non-metals



🟦 Metals
🟨 Non-metals
🟪 Metalloids

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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
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	
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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, K, Xe, Rn)

Groups 3 – 11 : Transition metals (excluding Zn)

La – Lu : Lanthanoids (lanthanides)

Ac – Lr : Actinoids (actinides)

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Electron configurations and the periodic table

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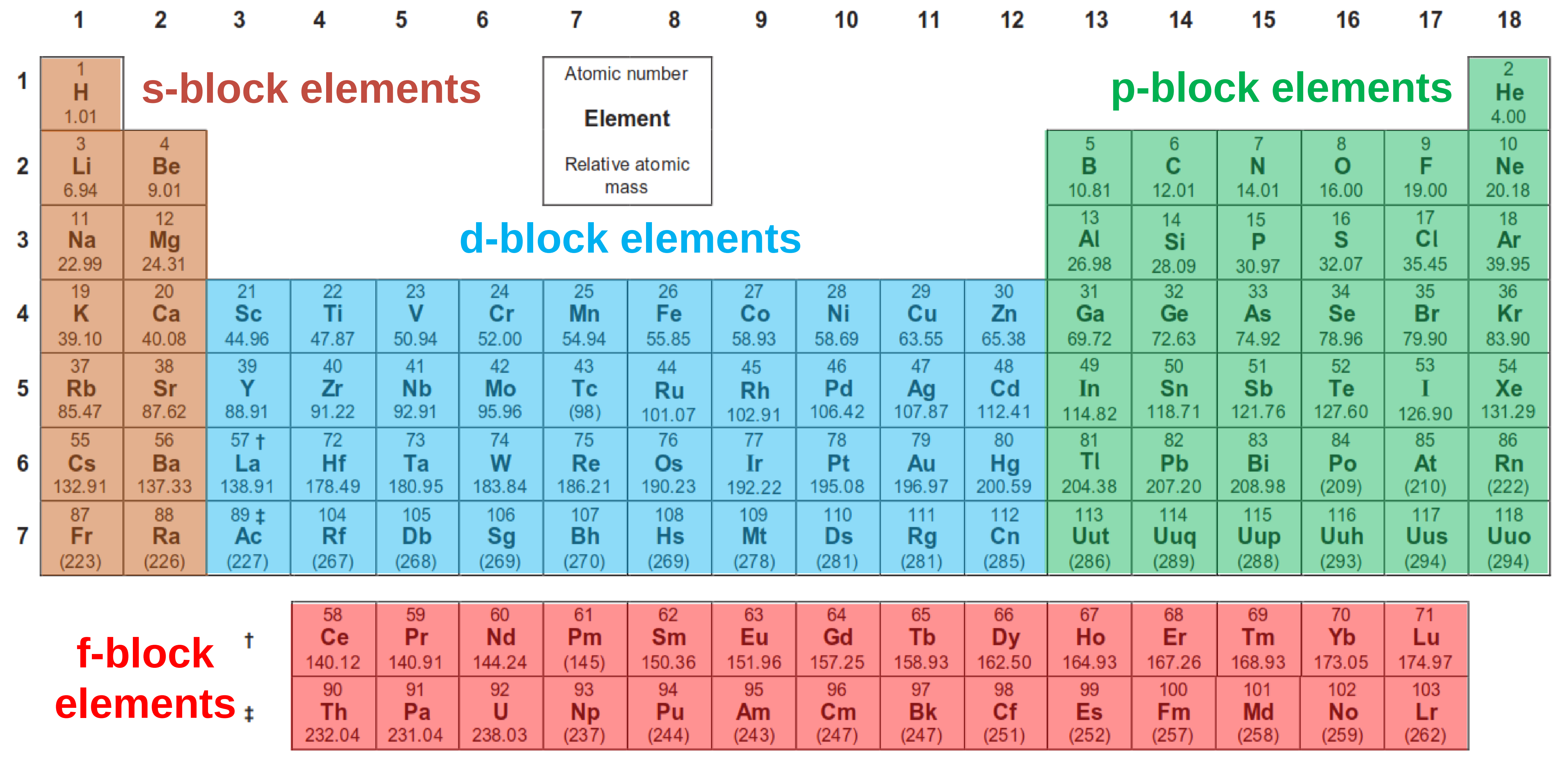
Electron configurations



s-block elements

p-block elements

d-block elements

Atomic number
Element
Relative atomic mass

	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
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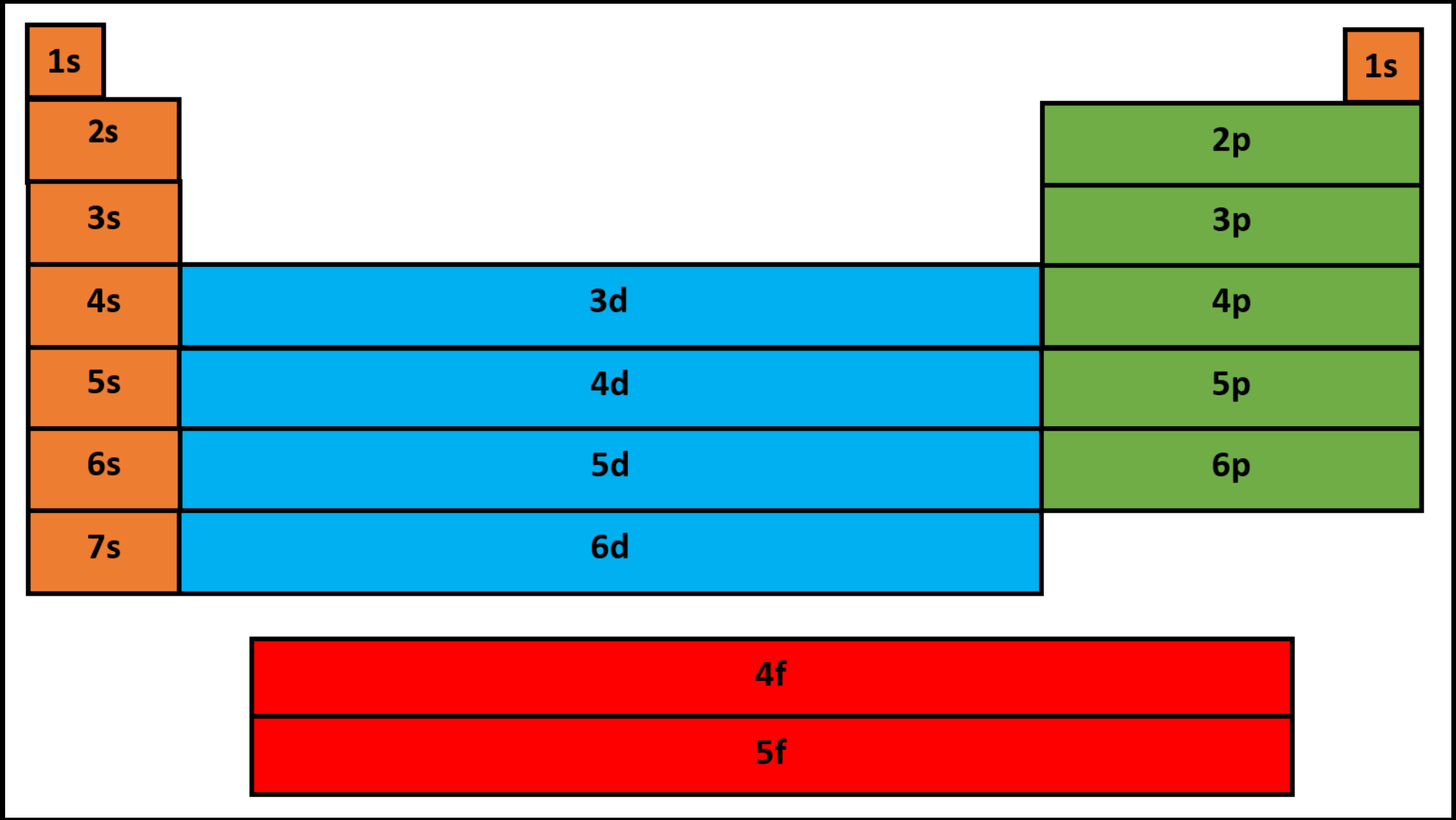
f-block elements

†	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

The block to which an element belongs indicates which sub-level is being filled with electrons.



s-block elements :
s sub-level

p-block elements :
p sub-level

d-block elements :
d sub-level

f-block elements :
f sub-level

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	$[Ar]\ 4s^2\ 3d^7$
Nd	f-block	$[Xe]\ 6s^2\ 4f^4$

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: $1s^2\ 2s^1$

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

Number of valence electrons (1)
Group 1

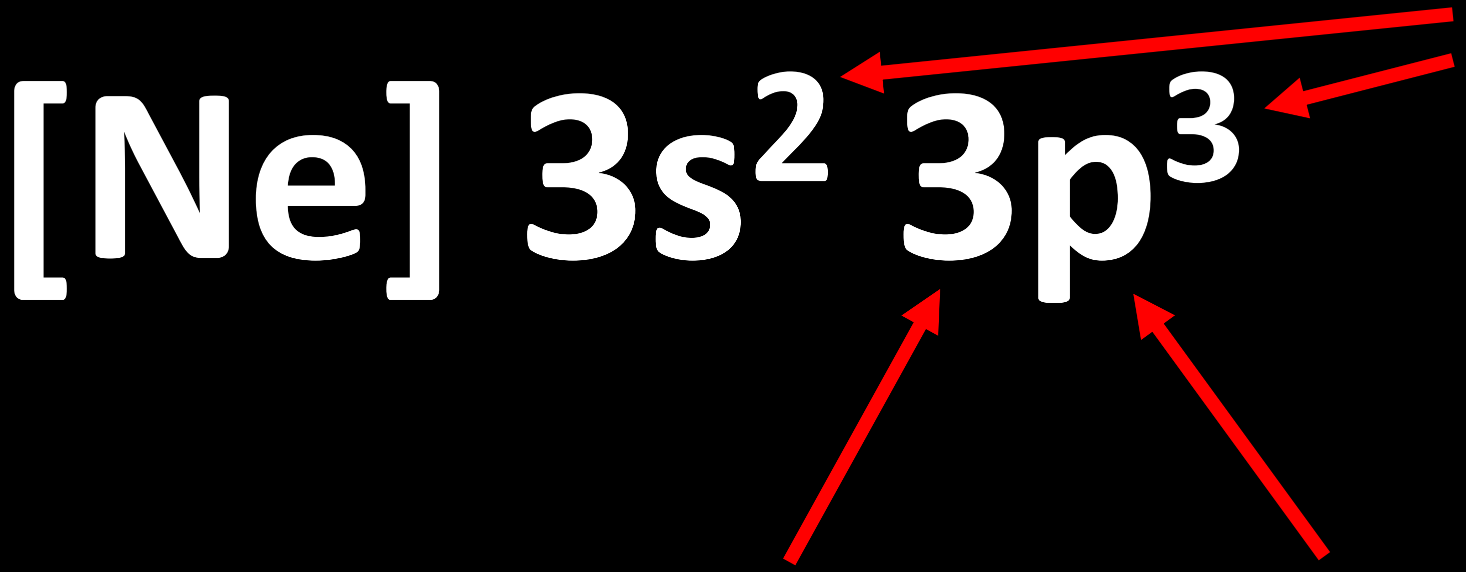
Highest occupied main energy level and period number

Block (s)

Electron configurations

Phosphorus $Z = 15$

Electron configuration: [Ne] $3s^2$ $3p^3$



[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: [Kr] $5s^2$

$$[Kr]\ 5s^2$$

Number of valence electrons (2)
Group 2

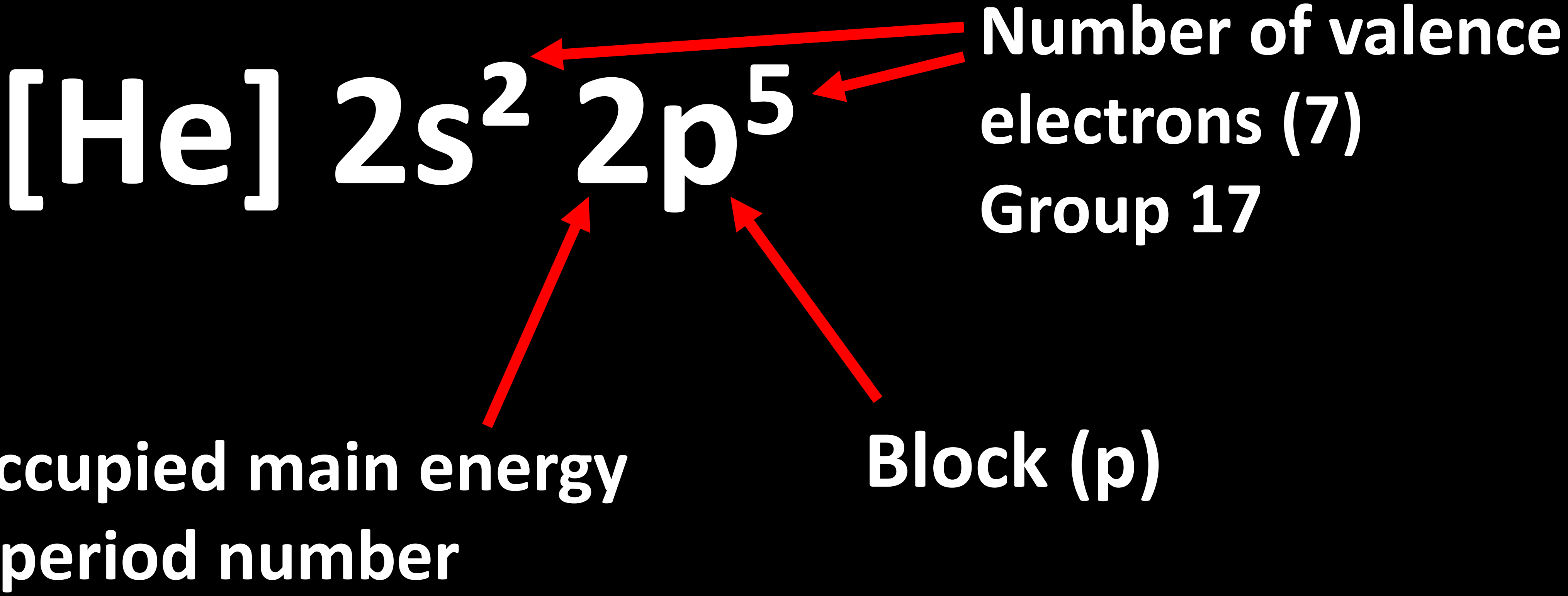
Highest occupied main energy level and period number

Block (s)

Electron configurations

Fluorine Z = 9

Electron configuration: [He] $2s^2$ $2p^5$



[He] $2s^2$ $2p^5$

Number of valence electrons (7)
Group 17

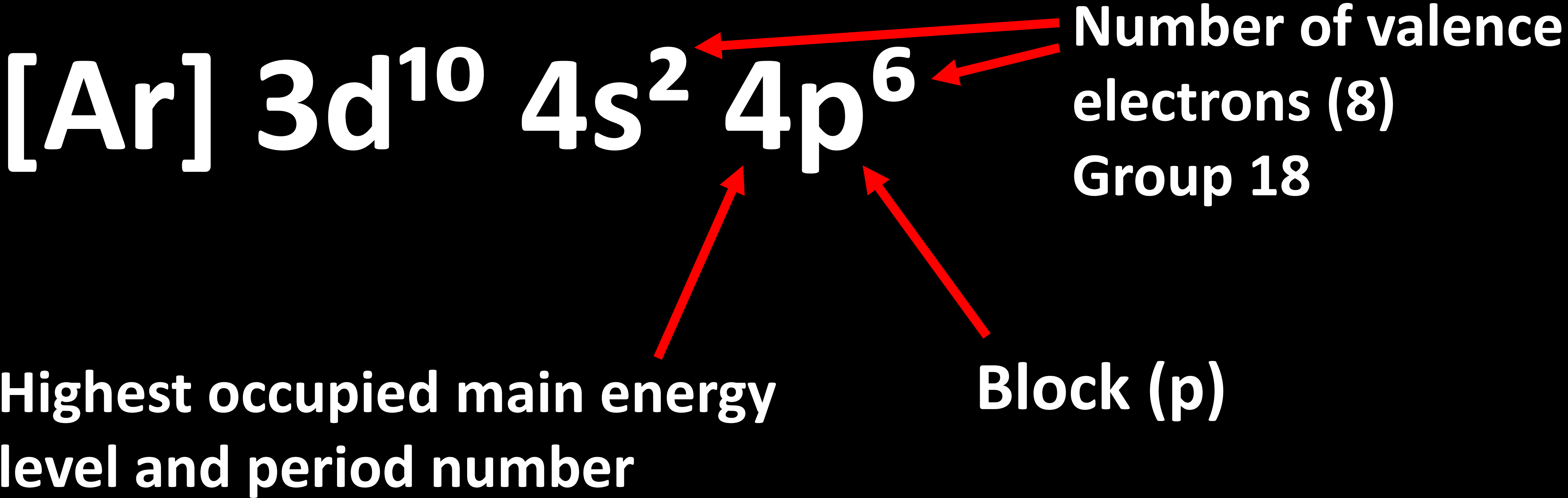
Highest occupied main energy level and period number

Block (p)

Electron configurations

Krypton Z = 36

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



[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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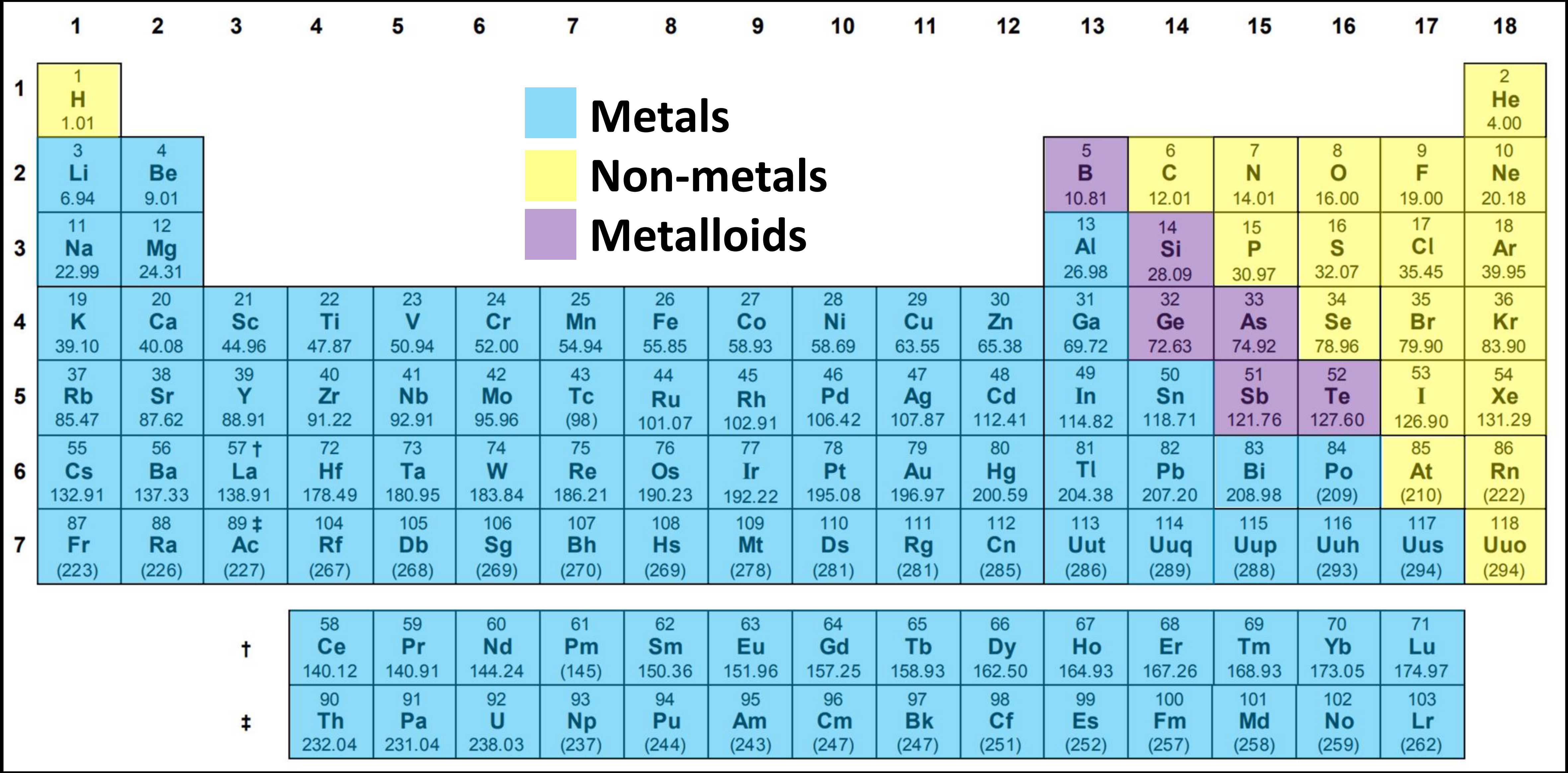
Properties of metals and non-metals

Metals and non-metals



🟦 Metals
🟨 Non-metals
🟪 Metalloids

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Metals and non-metals

Metals	Non-metals
Solids at room temperature	Mostly gases at room temperature
Have a metallic lustre	Dull
Malleable and ductile	Brittle
High electrical conductivity	Poor electrical conductivity
High densities	Low densities
Low ionisation energies	High ionisation energies
Low electronegativity values	High electronegativity values

Metals and non-metals

Metals	Non-metals
Lose electrons to form positive ions (cations)	Gain electrons to form negative ions (anions)
Behave as reducing agents (undergo oxidation)	Behave as oxidising agents (undergo reduction)
Form ionic bonds with non-metals	Form covalent bonds with other non-metals
Form basic oxides	Form acidic oxides

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Electron shielding

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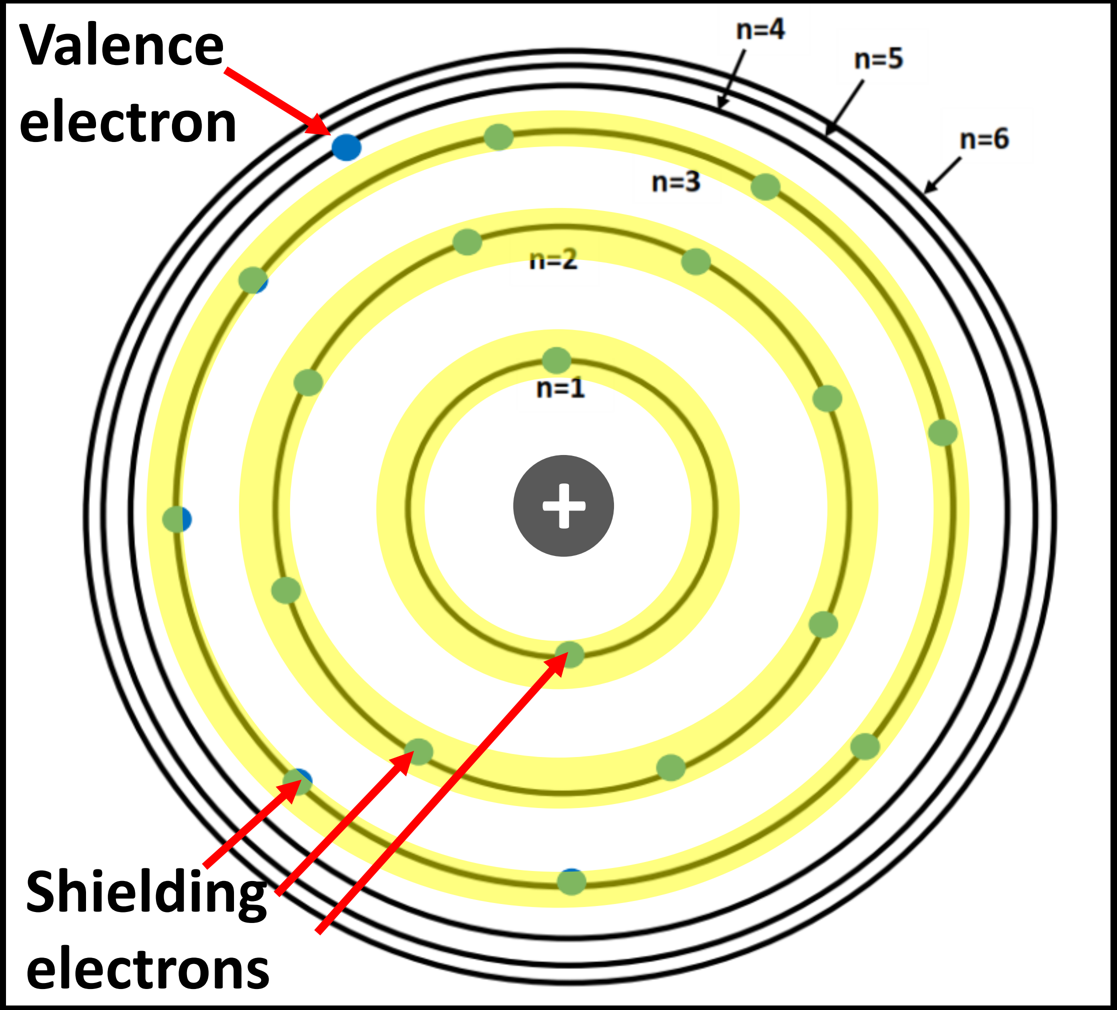
Electron shielding



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.

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

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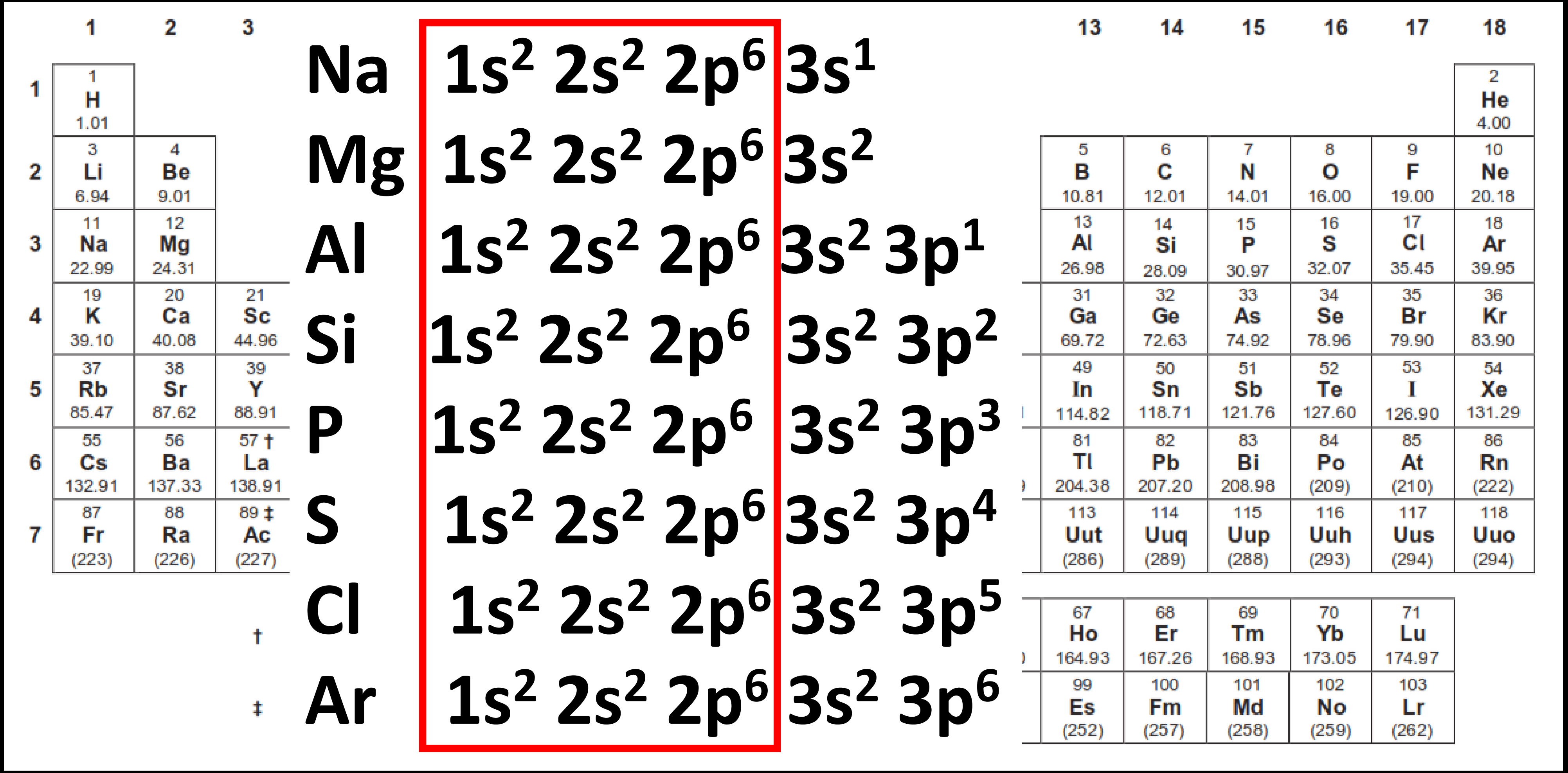
Electron shielding

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

Electron shielding increases down a group.



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

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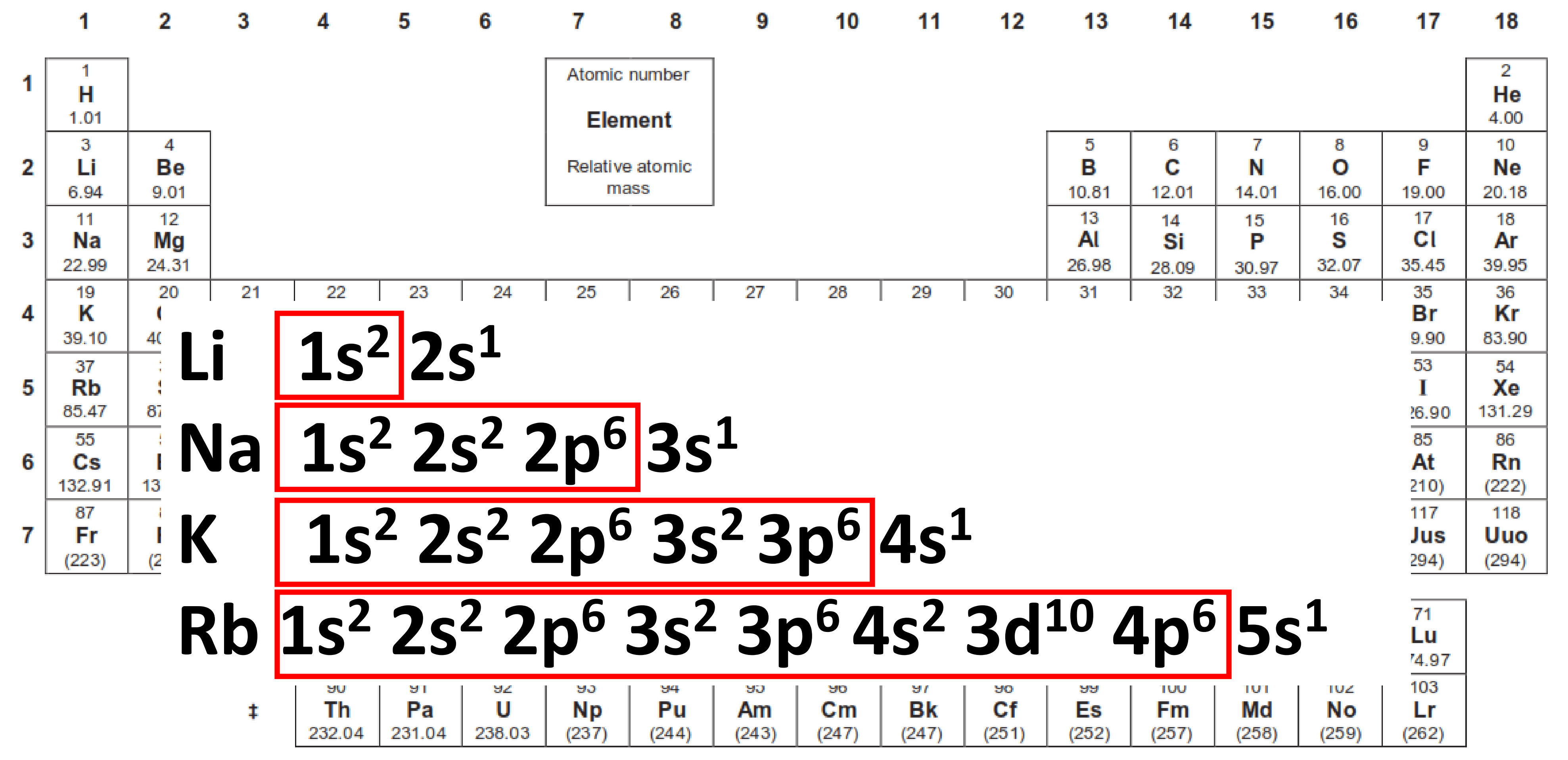
Electron shielding

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

Electron shielding increases down a group.



Li $\boxed{1s^2}\ 2s^1$

Na $\boxed{1s^2\ 2s^2\ 2p^6}\ 3s^1$

K $\boxed{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}\ 4s^1$

Rb $\boxed{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6}\ 5s^1$

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Effective nuclear charge

Effective nuclear charge

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

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

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_{eff} (Na) = 11 – 10 = +1

Z_{eff} (Mg) = 12 – 10 = +2

Z_{eff} (Al) = 13 – 10 = +3

Z_{eff} (Si) = 14 – 10 = +4

Z_{eff} (P) = 15 – 10 = +5

Z_{eff} (S) = 16 – 10 = +6

Z_{eff} (Cl) = 17 – 10 = +7

Z_{eff} (Ar) = 18 – 10 = +8

Effective nuclear charge

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

Li $\boxed{1s^2}\ 2s^1$

Na $\boxed{1s^2\ 2s^2\ 2p^6}\ 3s^1$

K $\boxed{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6}\ 4s^1$

Rb $\boxed{1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6}\ 5s^1$

Z_{eff} (Li) = 3 – 2 = +1

Z_{eff} (Na) = 11 – 10 = +1

Z_{eff} (K) = 19 – 18 = +1

Z_{eff} (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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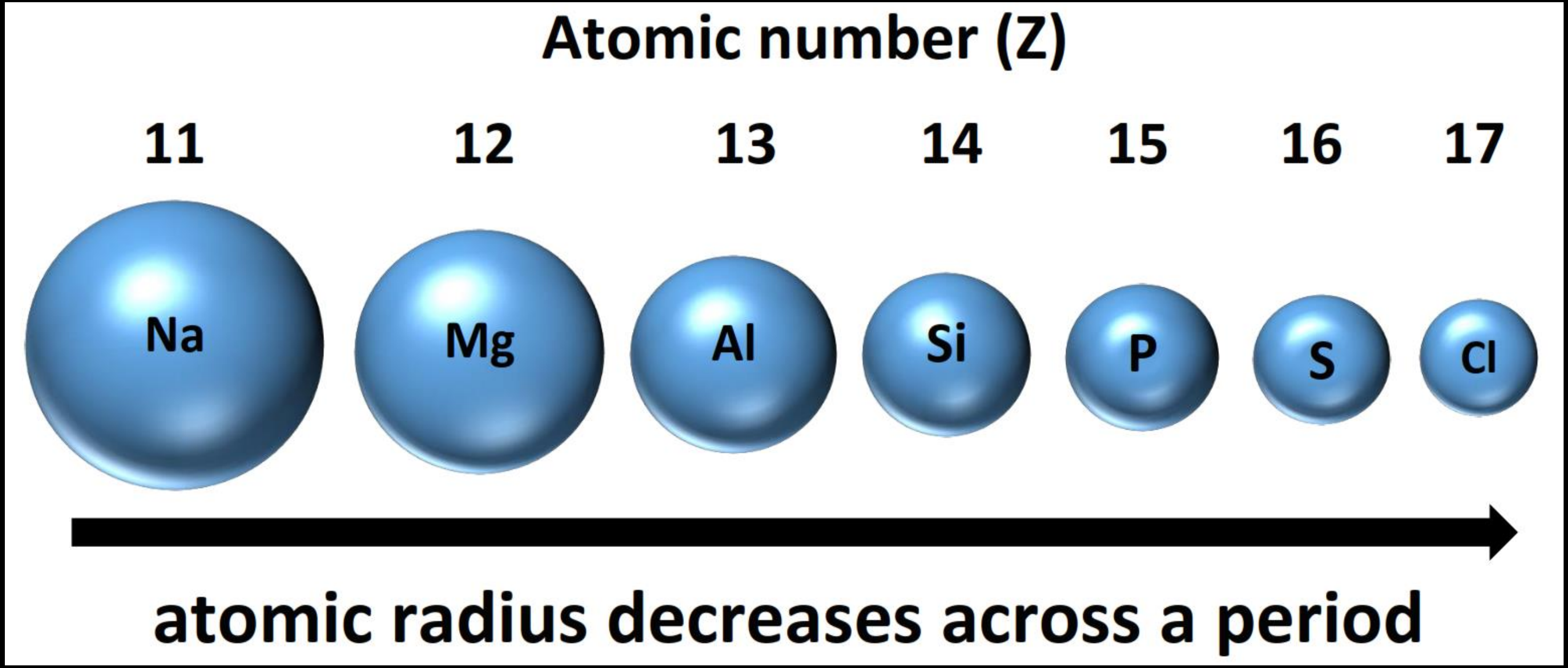
Trends in atomic radii

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.

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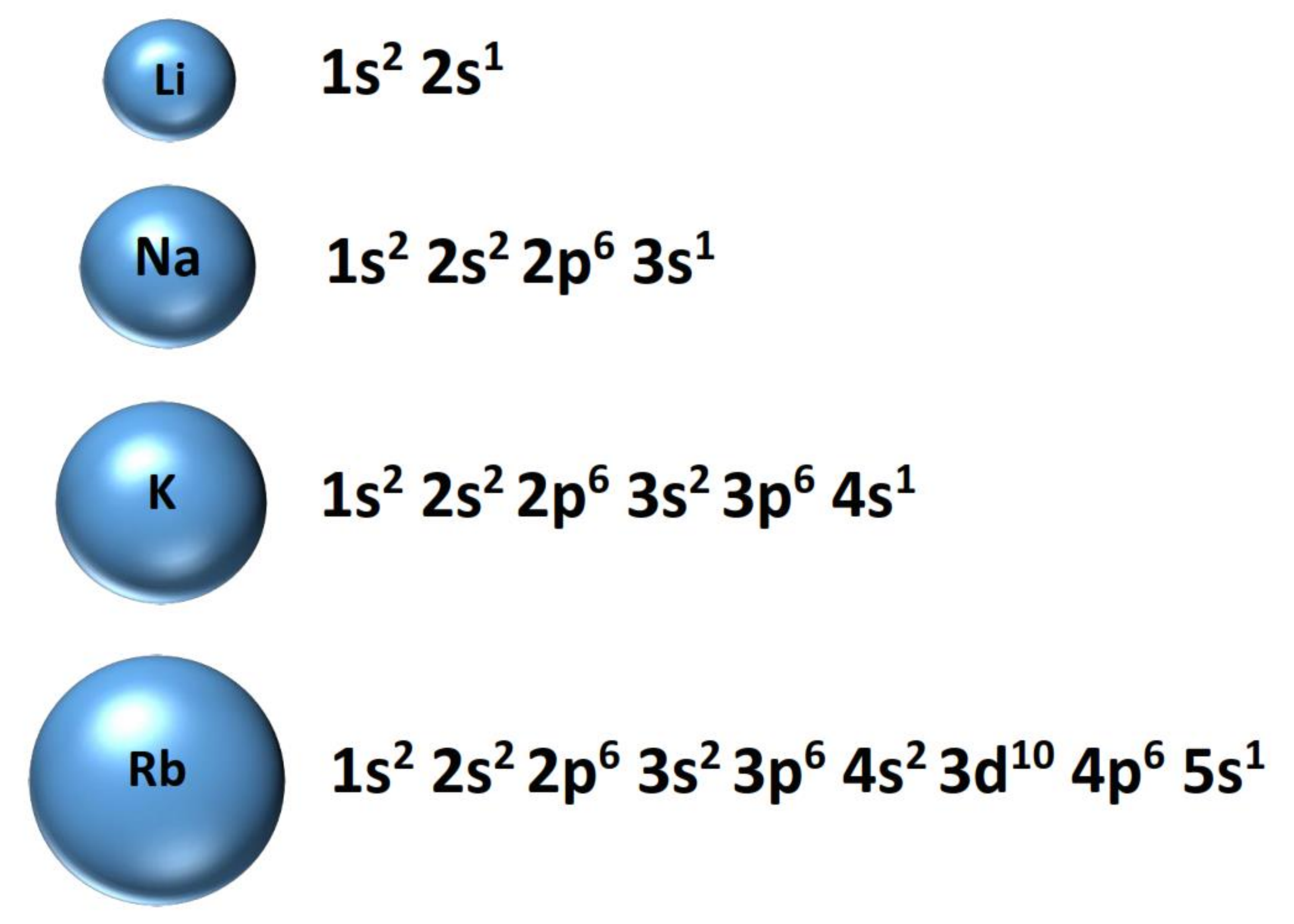
Trends in atomic radii



Li $1s^2\ 2s^1$

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

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

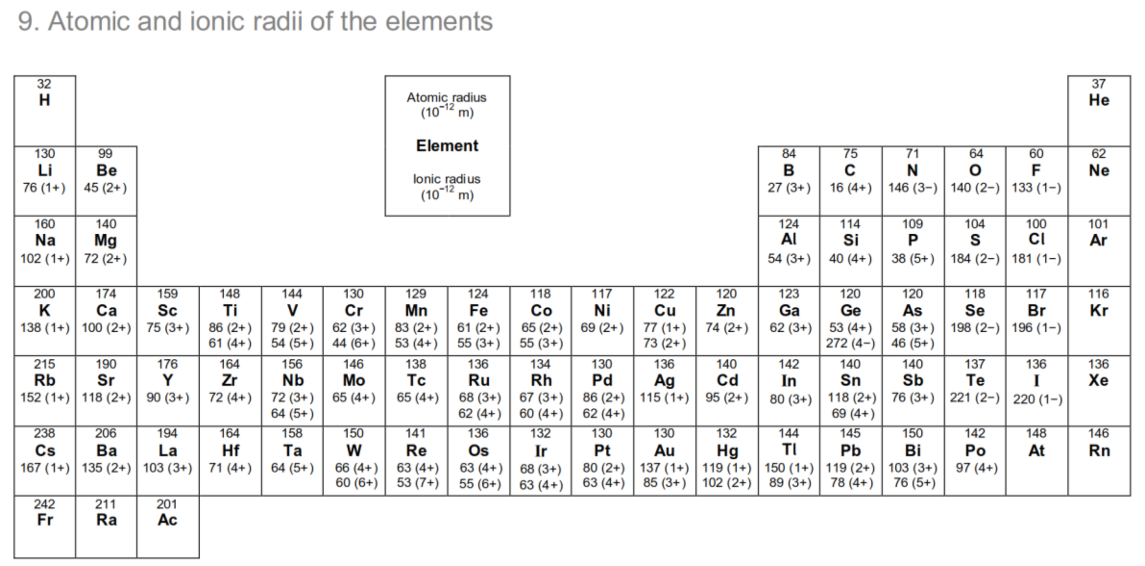
Rb $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^1$

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) **Element** Ionic radius (10^{-12} m)											37 **He**
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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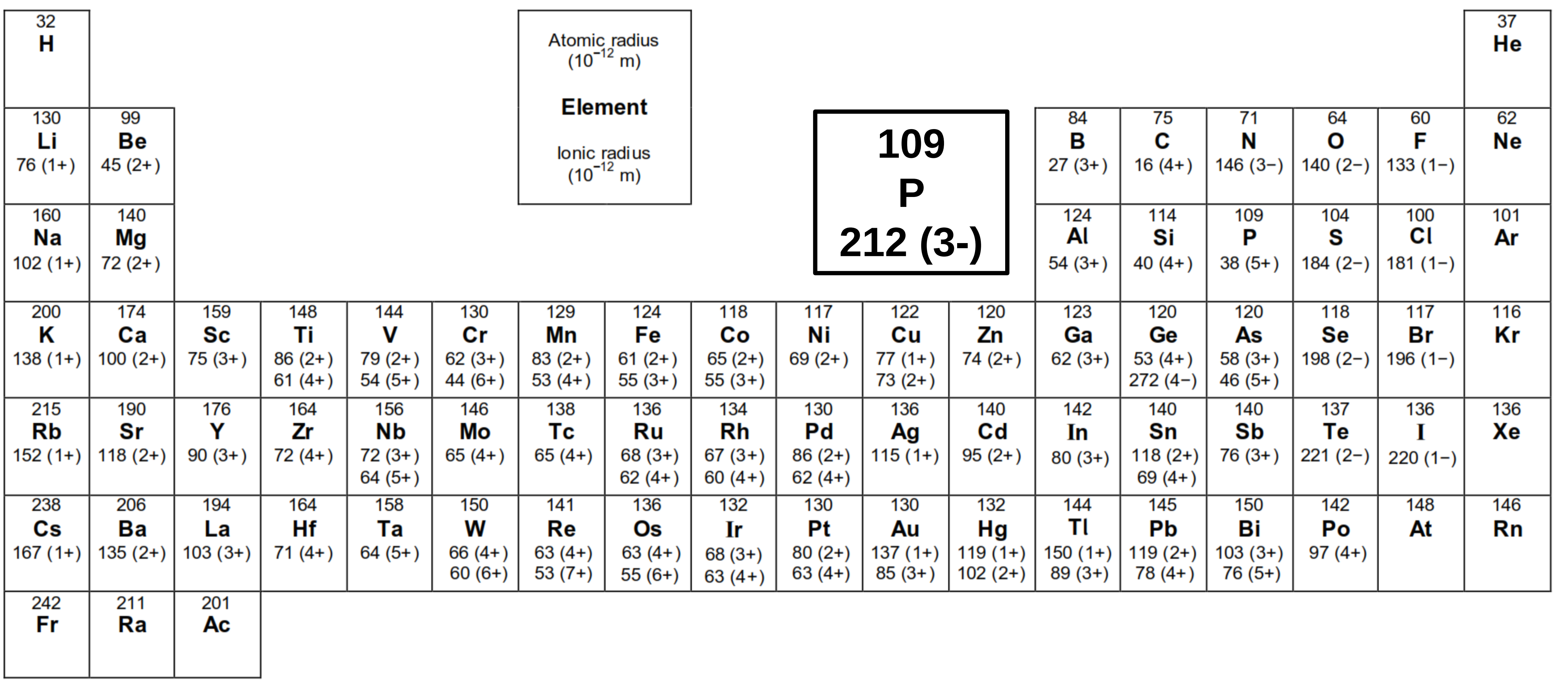
Trends in ionic radii

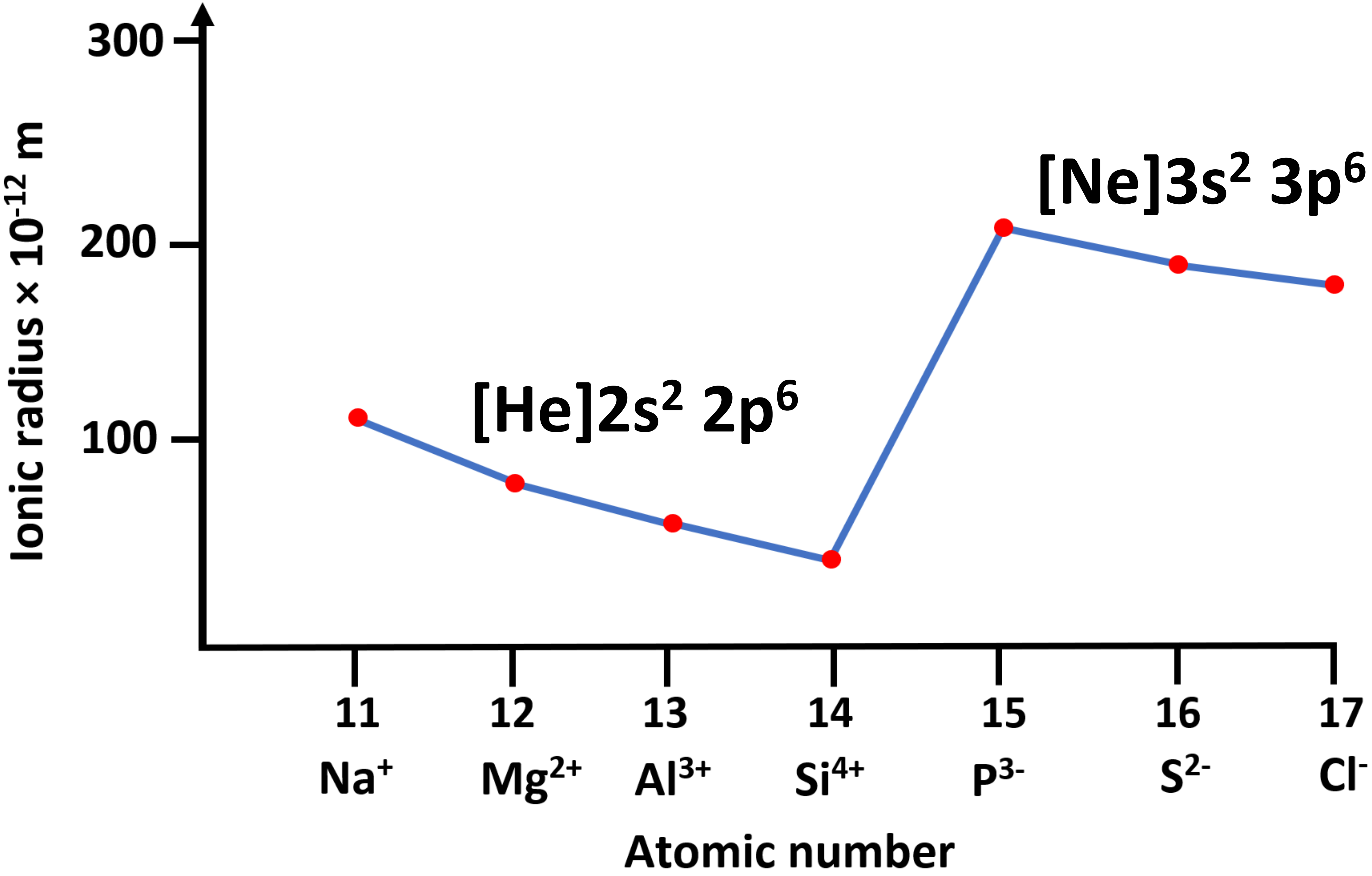
Trends in ionic radii



32 **H**						Atomic radius (10^{-12} m) **Element** Ionic radius (10^{-12} m)											37 **He**
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**															

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 (× 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).

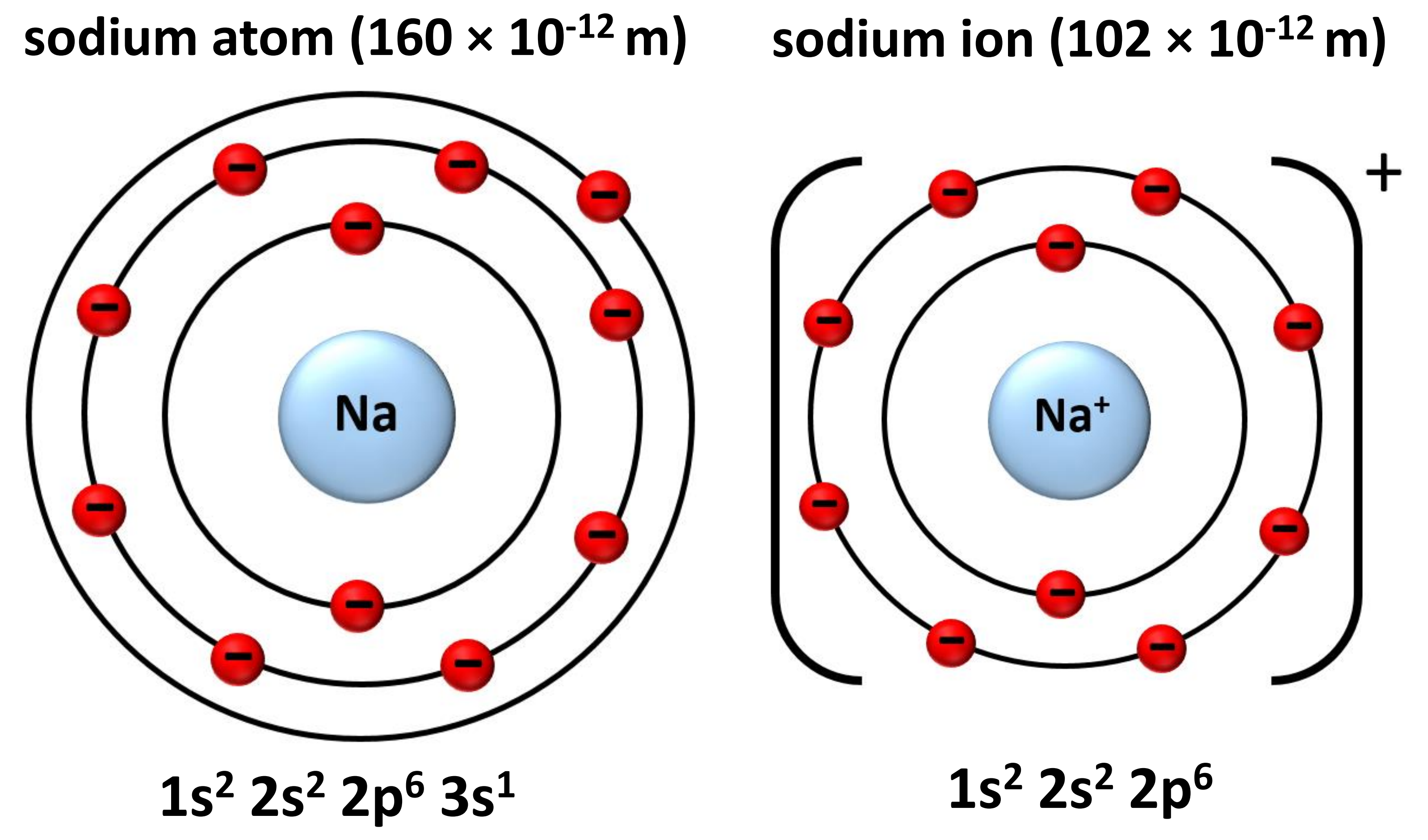
The number of protons increases but the number of electrons remains the same.

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
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

Trends in ionic radii



sodium atom (160×10^{-12} m)

sodium ion (102×10^{-12} m)

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

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

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.

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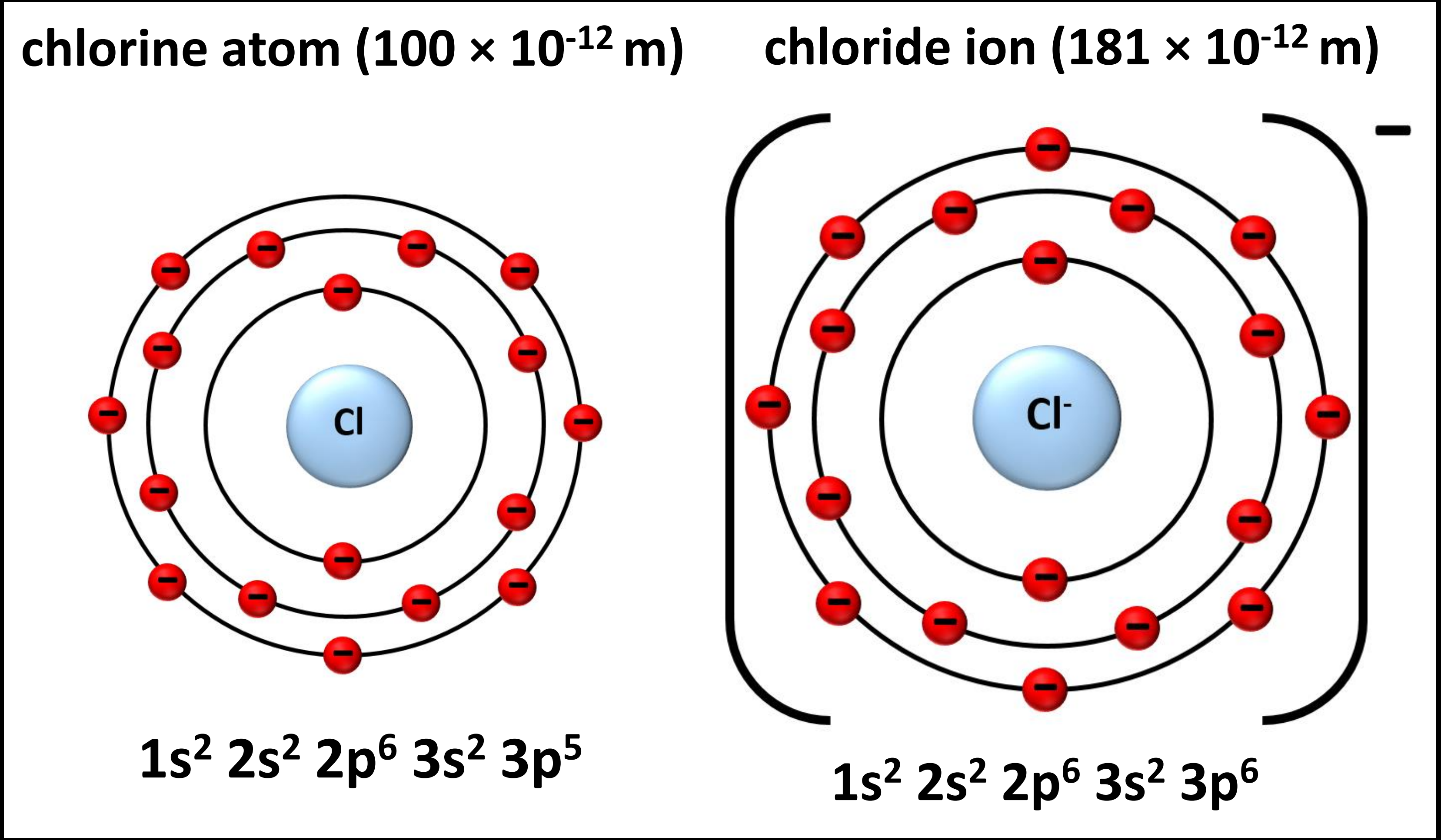
Trends in ionic radii



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.

$$Na_{(g)} \rightarrow Na^{+}_{(g)} + e^{-}$$

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

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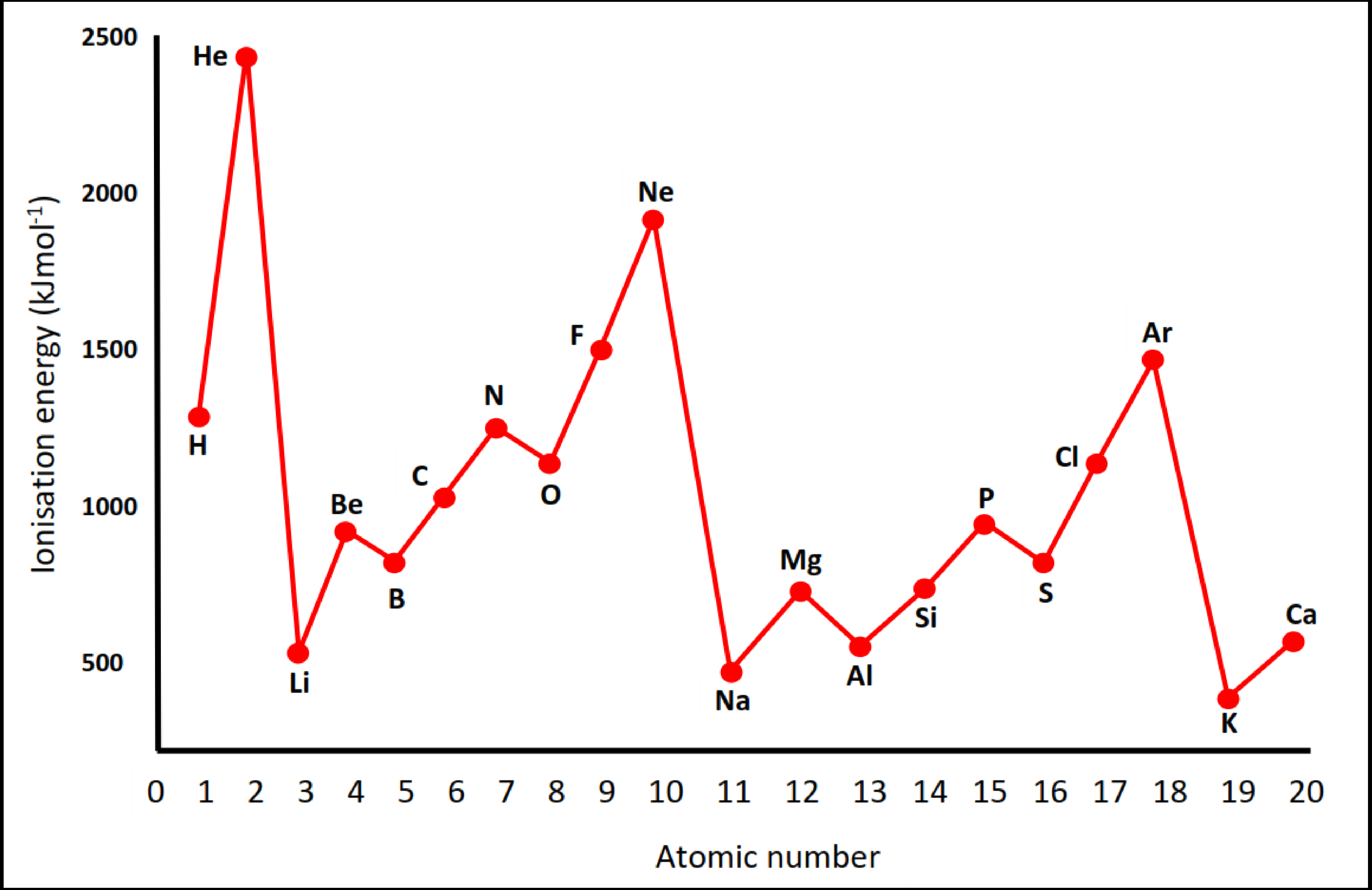
Ionisation energy



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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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.

$$X_{(g)} + e^{-} \rightarrow X^{-}_{(g)}$$

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

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

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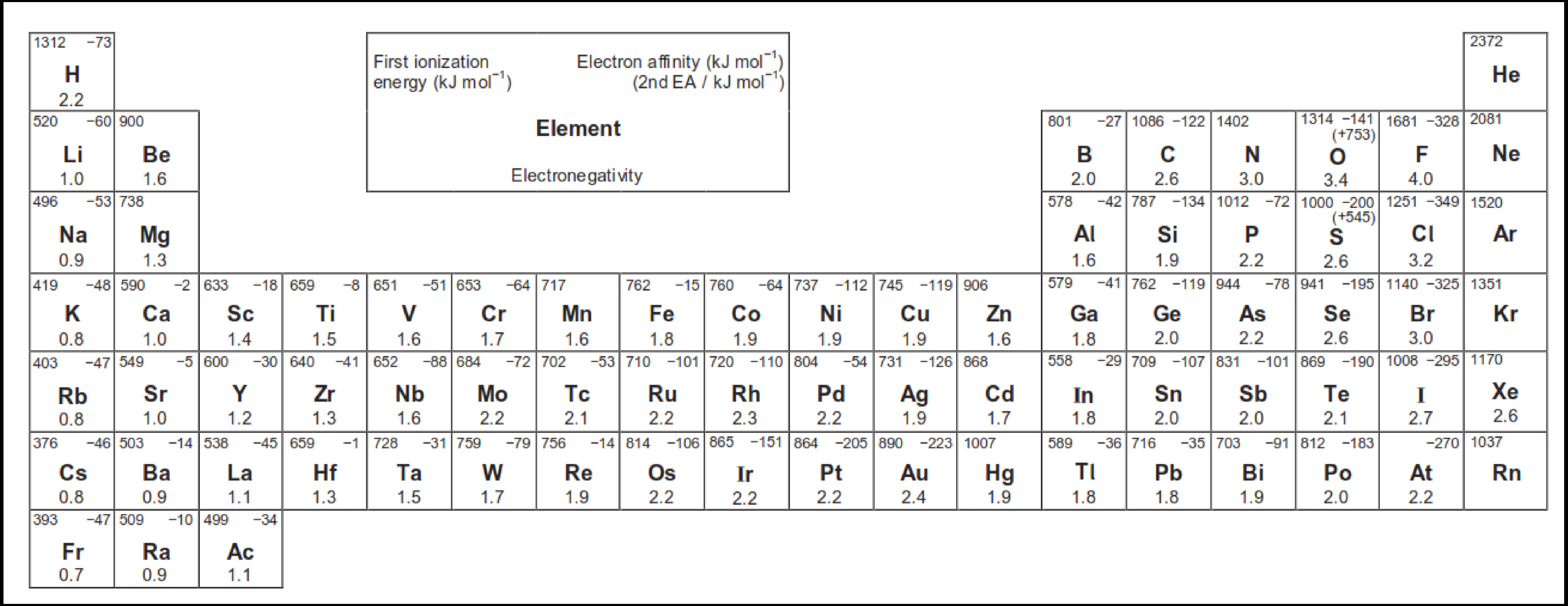
Electron affinity

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



First ionization energy (kJ mol^{-1}) | Electron affinity (kJ mol^{-1}) (2nd EA / kJ mol^{-1})

Element

Electronegativity

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

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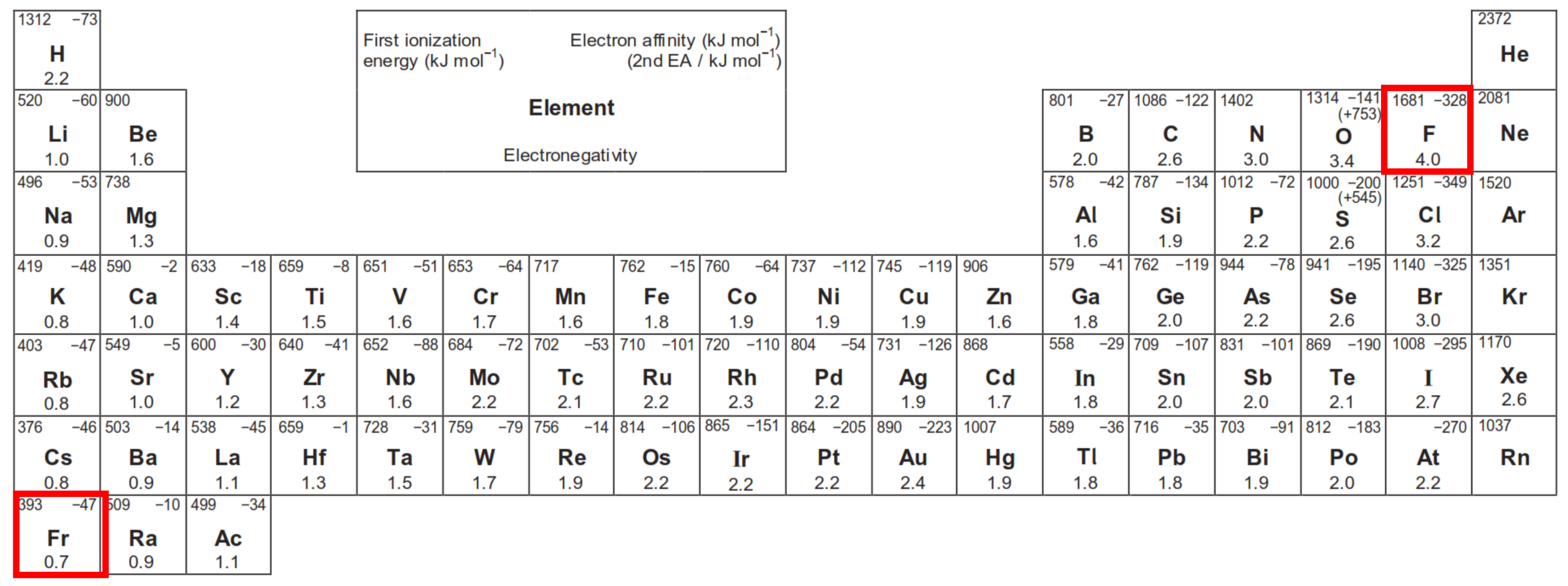
Electronegativity

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



First ionization energy (kJ mol^{-1})
Electron affinity (kJ mol^{-1})
(2nd EA / kJ mol^{-1})
Element
Electronegativity

1312 −73 **H** 2.2																	2372 **He**
520 −60 **Li** 1.0	900 **Be** 1.6											801 −27 **B** 2.0	1086 −122 **C** 2.6	1402 **N** 3.0	1314 −141 (+753) **O** 3.4	1681 −328 **F** 4.0	2081 **Ne**
496 −53 **Na** 0.9	738 **Mg** 1.3											578 −42 **Al** 1.6	787 −134 **Si** 1.9	1012 −72 **P** 2.2	1000 −200 (+545) **S** 2.6	1251 −349 **Cl** 3.2	1520 **Ar**
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376 −46 **Cs** 0.8	503 −14 **Ba** 0.9	538 −45 **La** 1.1	659 −1 **Hf** 1.3	728 −31 **Ta** 1.5	759 −79 **W** 1.7	756 −14 **Re** 1.9	814 −106 **Os** 2.2	865 −151 **Ir** 2.2	864 −205 **Pt** 2.2	890 −223 **Au** 2.4	1007 **Hg** 1.9	589 −36 **Tl** 1.8	716 −35 **Pb** 1.8	703 −91 **Bi** 1.9	812 −183 **Po** 2.0	−270 **At** 2.2	1037 **Rn**
393 −47 **Fr** 0.7	509 −10 **Ra** 0.9	499 −34 **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

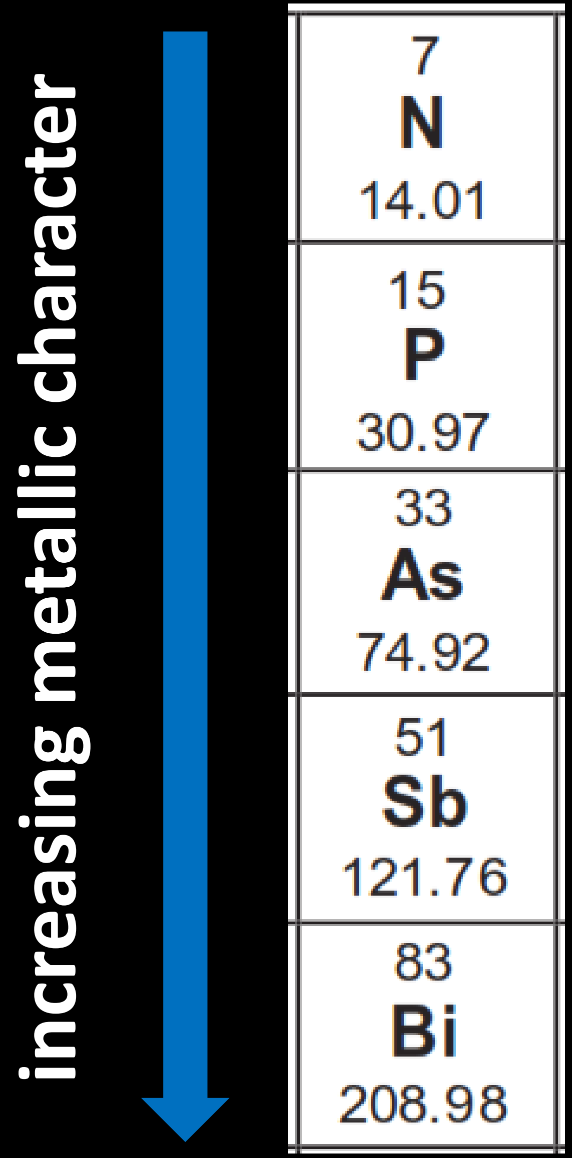
Metallic character

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



nitrogen – gaseous non-metal (3- ions)

phosphorus – solid non-metal (3- ions)

arsenic – metalloid

antimony – metalloid (more metallic)

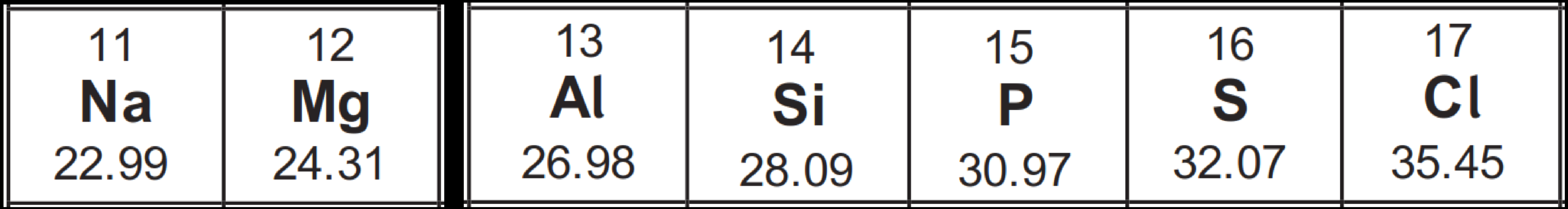
bismuth – metal (2+ ions)

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

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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Non-metallic character

Non-metallic character is the tendency of an element to accept electrons and form negative ions. Non-metallic character increases from left to right across a period and decreases from top to bottom in the periodic table.

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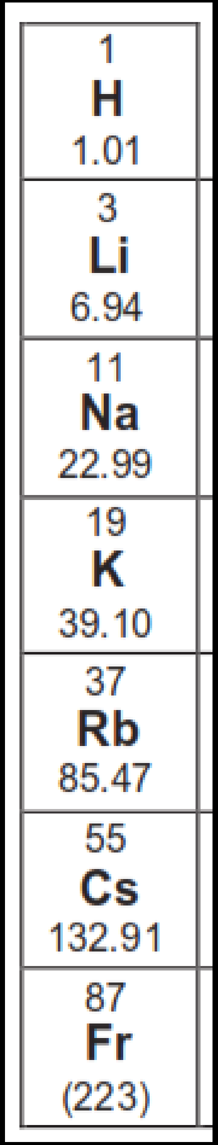
Group 1 metals

Group 1 metals



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)

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

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)

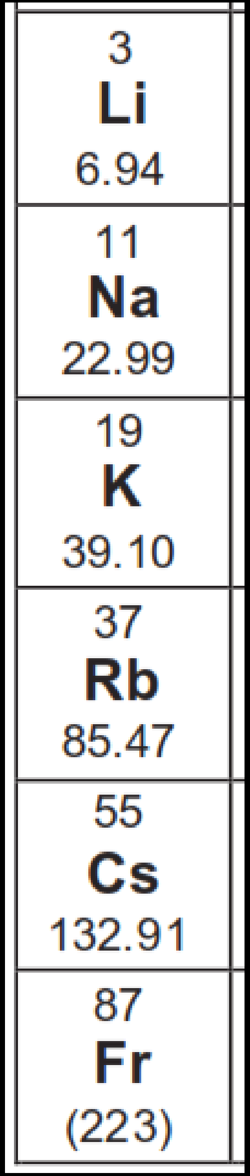
Group 1 metals are soft shiny metals that can be easily cut.

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 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)

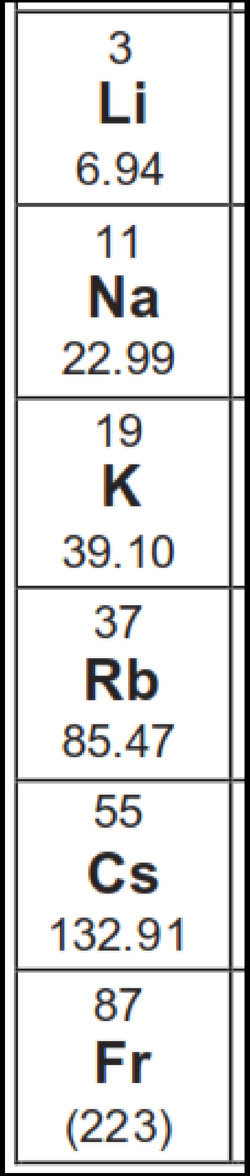
Group 1 metals are stored in oil to prevent the reaction with oxygen in the air.

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

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

$$2Li_{(s)} + 2H_2O_{(l)} \rightarrow 2LiOH_{(aq)} + H_{2(g)}$$

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.

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Group 1 metals

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

$$2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$$

$$2K_{(s)} + F_{2(g)} \rightarrow 2KF_{(s)}$$

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

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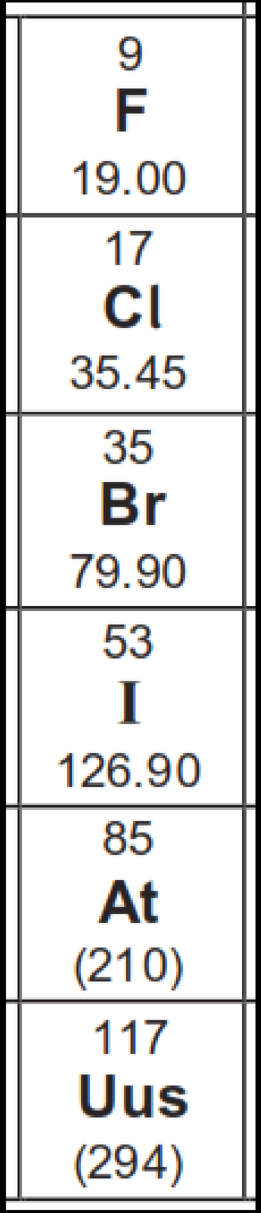
Group 17 elements

Group 17 elements



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

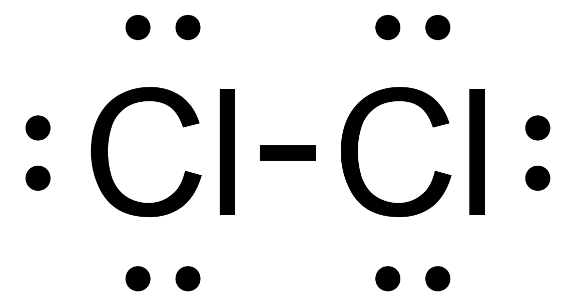
Group 17 elements (the halogens) are located in group 17 of the periodic table.

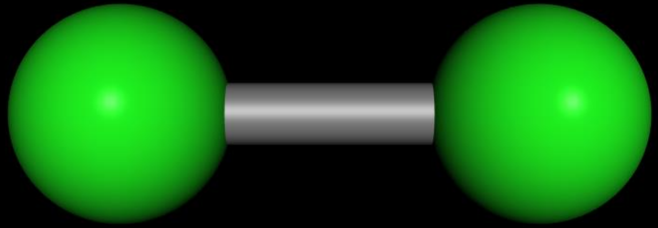
Element	Symbol	Electron configuration
Fluorine	F	$[He]\ 2s^2\ 2p^5$
Chorine	Cl	$[Ne]\ 3s^2\ 3p^5$
Bromine	Br	$[Ar]\ 4s^2\ 4p^5$
Iodine	I	$[Kr]\ 5s^2\ 5p^5$
Astatine	As	$[Xe]\ 6s^2\ 6p^5$

Group 17 elements

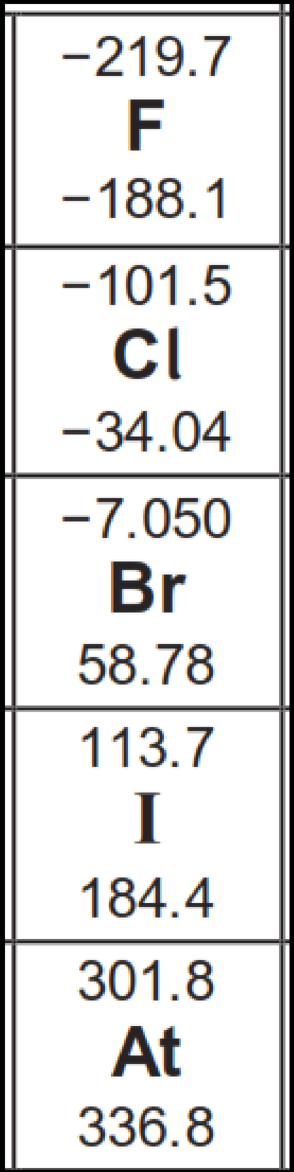
The group 17 elements are coloured; fluorine is a pale yellow gas, chlorine a greenish-yellow gas, bromine a red liquid and iodine a purple solid.

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





Group 17 elements

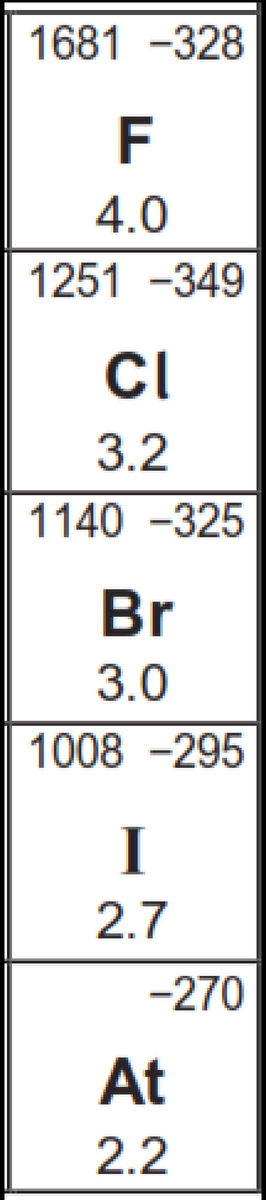


−219.7 F −188.1
−101.5 Cl −34.04
−7.050 Br 58.78
113.7 I 184.4
301.8 At 336.8

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.

Element	Molar mass (g mol^{-1})	Melting point (^{o}C)	Boiling point (^{o}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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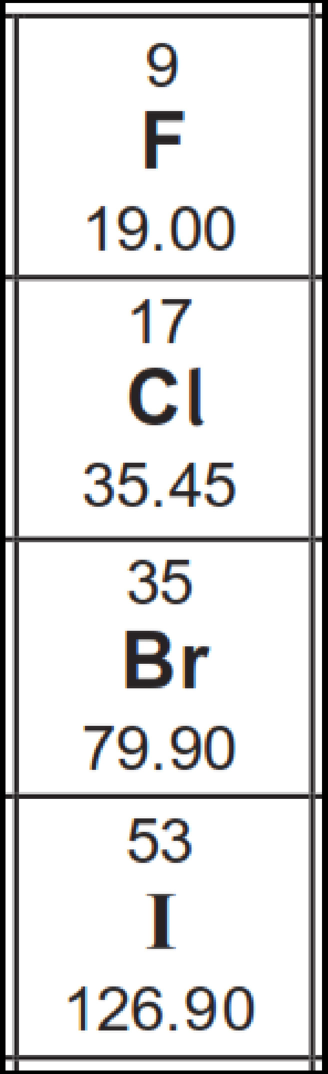
Displacement reactions of the group 17 elements

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

$Cl_{2(aq)} + 2Br^{-}_{(aq)} \rightarrow 2Cl^{-}_{(aq)} + Br_{2(aq)}$

$Cl_{2(aq)} + 2I^{-}_{(aq)} \rightarrow 2Cl^{-}_{(aq)} + I_{2(aq)}$

$Cl_{2(aq)} + 2F^{-}_{(aq)} \rightarrow$ no reaction

$Br_{2(aq)} + 2I^{-}_{(aq)} \rightarrow 2Br^{-}_{(aq)} + I_{2(aq)}$

$Br_{2(aq)} + 2Cl^{-}_{(aq)} \rightarrow$ no reaction

$I_{2(aq)} + 2Br^{-}_{(aq)} \rightarrow$ no reaction

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Bonding and acid-base properties of the period 3 oxides

Bonding

Formula of oxide	$Na_2O_{(s)}$	$MgO_{(s)}$	$Al_2O_{3(s)}$	$SiO_{2(s)}$	$P_4O_{10(s)}$ $P_4O_{6(s)}$	$SO_{3(l)}$ $SO_{2(g)}$	$Cl_2O_{7(l)}$ $Cl_2O_{(g)}$
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	$Na_2O_{(s)}$	$MgO_{(s)}$	$Al_2O_{3(s)}$	$SiO_{2(s)}$	$P_4O_{10(s)}$ $P_4O_{6(s)}$	$SO_{3(l)}$ $SO_{2(g)}$	$Cl_2O_{7(l)}$ $Cl_2O_{(g)}$
Acid-base properties	basic		amphoteric	acidic			

Sodium oxide reacts with water as follows:

$$Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$$

Magnesium oxide reacts with water as follows:

$$MgO(s) + H_2O(l) \rightarrow Mg(OH)_2(aq)$$

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Acid deposition

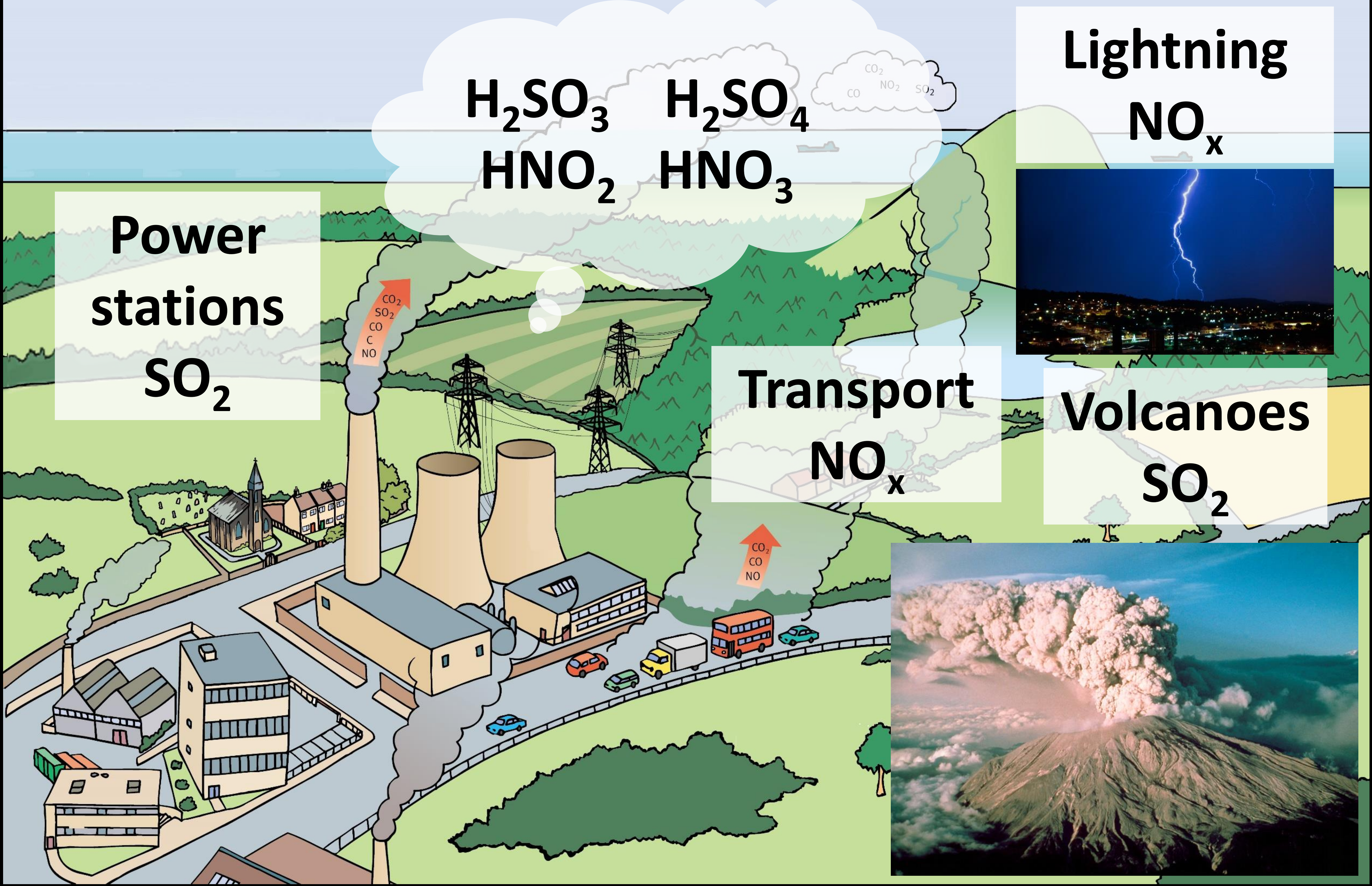
Acid deposition

Rainwater is naturally acidic with a pH of 5.6

$$CO_{2(g)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$$

Acid deposition has a pH of less than 5.0

Dry deposition	Wet deposition
Acidic gases and particles	Acid rain, fog and snow



H_2SO_3 H_2SO_4

HNO_2 HNO_3

Lightning

NO_x

Power stations

SO_2

Transport

NO_x

Volcanoes

SO_2

Acid deposition

$$S_{(s)} + O_{2(g)} \rightarrow SO_{2(g)}$$

$$SO_{2(g)} + H_2O_{(l)} \rightarrow H_2SO_{3(aq)}$$

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

$$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$$

Acid deposition

$$N_{2(g)} + O_{2(g)} \rightarrow 2NO_{(g)}$$

$$2NO_{(g)} + O_{2(g)} \rightarrow 2NO_{2(g)}$$

$$2NO_{2(g)} + H_2O_{(l)} \rightarrow HNO_{3(aq)} + HNO_{2(aq)}$$

$$4NO_{2(g)} + O_{2(g)} + 2H_2O_{(l)} \rightarrow 4HNO_{3(aq)}$$

Acid deposition

Acidic gas	Sources	Acids formed
SO_2	Combustion of coal that contains sulfur Volcanic eruptions	H_2SO_3 H_2SO_4
NO_x	Internal combustion engines Lightning	HNO_2 HNO_3

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Ocean acidification

Ocean acidification

Approximately 30% of anthropogenic carbon dioxide is absorbed by the oceans (carbon sink).

$$CO_{2(g)} \rightleftharpoons CO_{2(aq)}$$

A heterogeneous equilibrium exists between concentrations of gaseous carbon dioxide in the atmosphere and aqueous carbon dioxide dissolved in the oceans.

$$CO_{2(aq)} + H_2O_{(l)} \rightleftharpoons H_2CO_{3(aq)}$$

Ocean acidification

Carbonic acid (H_2CO_3) is a weak acid which partially dissociates in solution to produce H^+(aq).

$$H_2CO_{3(aq)} \rightleftharpoons H^+_{(aq)} + HCO_3^-{}_{(aq)}$$

The increasing concentration of H^+(aq) causes the pH of the oceans to decrease.

Since the beginning of the industrial revolution, the pH of the oceans has decreased by 0.1 pH units.

Continued acidification of the oceans could have harmful effects on marine organisms.

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Oxidation states

Oxidation states

The oxidation state is the hypothetical charge an atom would have if the bonds are assumed to be 100% ionic with no covalent character.

Oxidation states are written with the + or – first followed by the number (+2, not 2+).

The term oxidation number is sometimes used interchangeably with oxidation state.

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Oxidation states

Rules for determining oxidation states	
1	Elements are assigned an oxidation state of zero. Examples: Fe(s), Cu(s), Zn(s), O_2(g), Br_2(l), Cl_2(g) and N_2(g) are all elements and have oxidation states of zero.
2	The sum of the oxidation states of the atoms in a compound must be equal to zero. Example: In H_2O, the oxidation state of the O is −2 and the H is +1. The sum of the oxidation states is $(-2 + (2 \times +1)) = 0$
3	The charge on an ion is numerically equal to its oxidation state. Examples: The oxidation state of the Mg^{2+} ion is +2. The oxidation state of the S^{2-} ion is −2.
4	Hydrogen in compounds is assigned an oxidation state of +1 except in certain metal hydrides (e.g. NaH) in which it is −1. Examples: In methane, CH_4, the hydrogen has an oxidation state of +1 and the carbon is −4. In NaH, the Na has an oxidation state of +1 and the H is −1.

Oxidation states

Rules for determining oxidation states	
5	Fluorine in compounds is always assigned an oxidation state of −1.
6	Oxygen in a compound is assigned an oxidation state of −2 unless it is combined with fluorine (for example OF_2) or in a peroxide (H_2O_2). Examples: In OF_2, the F has an oxidation state of −1 and the O is +2. In H_2O_2, the H is +1 and the O is −1.
7	Chlorine in a compound has an oxidation state of −1 unless it is combined with oxygen or fluorine. Example: In Cl_2O, the oxidation state of the Cl is +1.
8	For a polyatomic ion (molecular ion) the sum of the oxidation states must equal the charge on the ion. Example: In the SO_4^{2-} ion, the oxidation state of the S is +6 and the O is −2. The sum of the oxidation states is $(+6 + (2 \times -2)) = -2$.

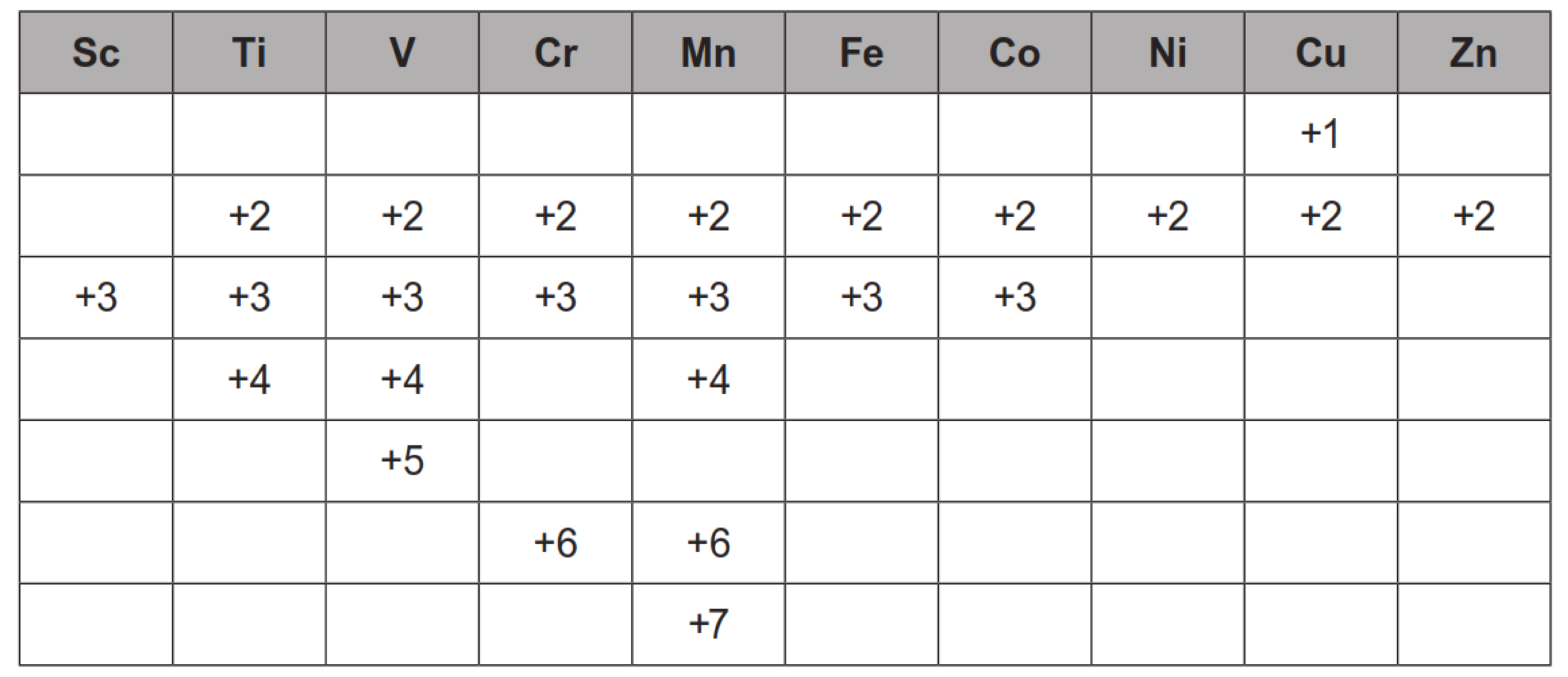
Oxidation states

The group 17 elements can have variable oxidation states.

Halogen	Oxidation state in compounds
Chlorine	-1, +1, +2, +3, +5, +6, +7
Bromine	-1, +1, +3, +5, +7
Iodine	-1, +1, +3, +5, +6, +7
Astatine	-1, +1, +3, +5, +7

Oxidation states

Oxidation states of the first row d-block elements.



Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn
								+1	
	+2	+2	+2	+2	+2	+2	+2	+2	+2
+3	+3	+3	+3	+3	+3	+3			
	+4	+4		+4					
		+5							
			+6	+6					
				+7					

Oxidation numbers

Oxidation numbers are represented by a Roman numeral.

Cu_2O

copper(I) oxide

CuO

copper(II) oxide

$FeCl_2$

Iron(II) chloride

$FeCl_3$

Iron(III) chloride