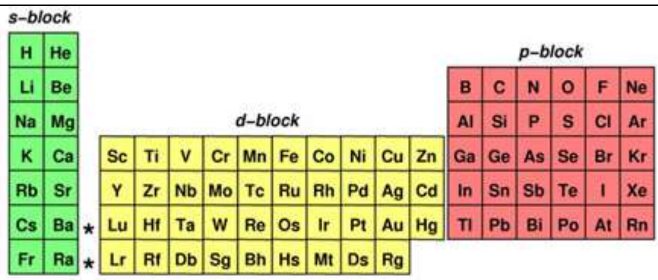
Unit 1 Chemistry Knowledge Organiser

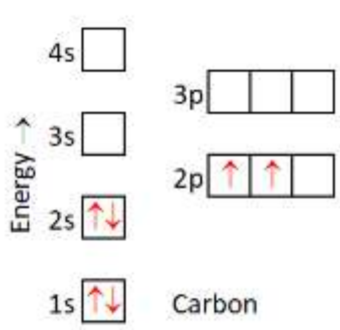


Electronic Structures:

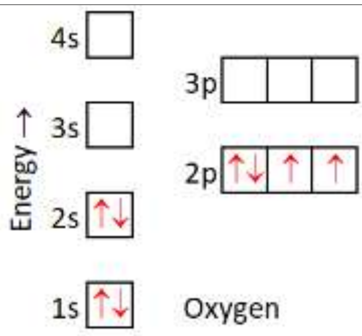
1. Big number = period (row).
2. Letter = electron block.
3. Small numbers = number of electrons.

Group 1 = alkali metals	Group 7 = halogens	Group 0 = noble gases	
(s) = Solid	(l) = liquid	(g) = gas	(aq) = aqueous

	Atom Electronic Structure		Ion Electronic Structure
H	$1s^1$	H^+	No electrons in ion.
He	$1s^2$	He	No ion as stable.
Li	$1s^2\ 2s^1$	Li^+	$1s^2$
Be	$1s^2\ 2s^2$	Be^{2+}	$1s^2$
B	$1s^2\ 2s^2\ 2p^1$	B^{3+}	$1s^2$
C	$1s^2\ 2s^2\ 2p^2$	$C^{4\pm}$	$1s^2\ (2s^2\ 2p^6)$
N	$1s^2\ 2s^2\ 2p^3$	N^{3-}	$1s^2\ 2s^2\ 2p^6$
O	$1s^2\ 2s^2\ 2p^4$	O^{2-}	$1s^2\ 2s^2\ 2p^6$
F	$1s^2\ 2s^2\ 2p^5$	F^-	$1s^2\ 2s^2\ 2p^6$
Ne	$1s^2\ 2s^2\ 2p^6$	Ne	No ion as stable.
Na	$1s^2\ 2s^2\ 2p^6\ 3s^1$	Na^+	$1s^2\ 2s^2\ 2p^6$
Mg	$1s^2\ 2s^2\ 2p^6\ 3s^2$	Mg^{2+}	$1s^2\ 2s^2\ 2p^6$
Al	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$	Al^{3+}	$1s^2\ 2s^2\ 2p^6$
Si	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$	$Si^{4\pm}$	$1s^2\ 2s^2\ 2p^6\ (3s^2\ 3p^6)$
P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$	P^{3-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
S	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$	S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Cl	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$	Cl^-	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Ar	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$	Ar	No ion as stable.
K	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$	K^+	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Ca	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$	Ca^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$



$1s^2\ 2s^2\ 2p_x^1\ 2p_y^1$



$1s^2\ 2s^2\ 2p_x^2\ 2p_y^1\ 2p_z^1$

Orbital Diagrams

1. Fill lowest energy s-orbitals first.
2. Then fill up p-orbitals singularly (with up arrows).
3. Then fill up p-orbitals doubly (with down arrows).

<u>Isotope:</u> Atoms of an element with the same number of protons but different mass number due to different number of neutrons.	If 75% of Cl is 35Cl and 25% is 37Cl what is the relative atomic mass to 1dp? (0.75 x 35) + (0.25 x 37) = 35.5

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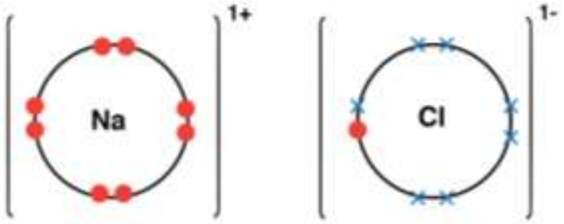
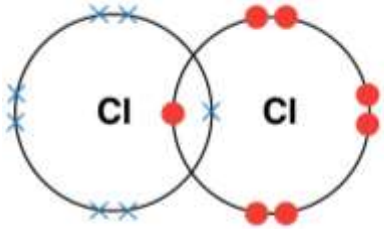
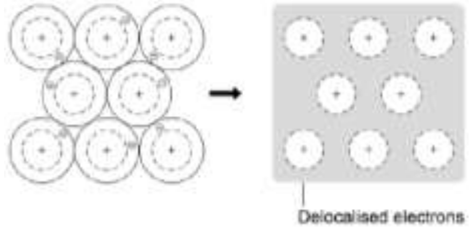
IONIC BOND: Strong electrostatic attraction between oppositely charged ions.	**COVALENT BOND:** A shared pair of electrons with strong electrostatic attraction between nuclei & electrons.	METALLIC BOND: The electrostatic attraction between nuclei of positively charged ions and delocalised electron(s).
		
1. Metal & non-metal ions. 2. Metal ions give electrons to non-metal atoms. 3. Separate circles with square brackets & a charge.	1. Non-metal atoms. 2. Non-metal atoms share pairs of electrons. 3. Overlapping circles with shared electrons in the overlap.	Describe structure & bonding in metals. Regular layers of atoms; In a (giant) lattice; Metallic bonding; Positively charged ions / metal cations; Surrounded by delocalised electrons;
Ionic properties: Solid doesn't conduct; Conducts when molten / in solution; Usually soluble / forms crystals;	Small covalent properties: Doesn't conduct electricity or heat; Some dissolve in water; Low melting / boiling points;	Metal properties: Conductor of electricity & heat; Ductile & malleable; High melting / boiling points;
Describe the structure of ionic compounds. Electrostatic attraction; (between) {positive and negative / oppositely charged} ions; (arranged in a) (giant) lattice structure;	Large covalent properties: Doesn't conduct electricity or heat; Doesn't dissolve in water; High melting / boiling points;	Why are metals ductile / malleable? Atoms are in rows / layers; That slide / slip / move over each other; Ductile: so metal can be drawn into wires without breaking bonding; Malleable: so metal can be hammered into shape without breaking;
Calcium Chloride uses: Mineral supplement, dust control, fertiliser… Sodium Sulfate uses: Paper pulping, drying agents, medication…	Why does methane CH_4 form a tetrahedral structure? All carbon atoms form 4 single bonds.	Why do metals conduct electricity? Metallic structure; delocalised electrons; (electrons) carry (electric) charge;

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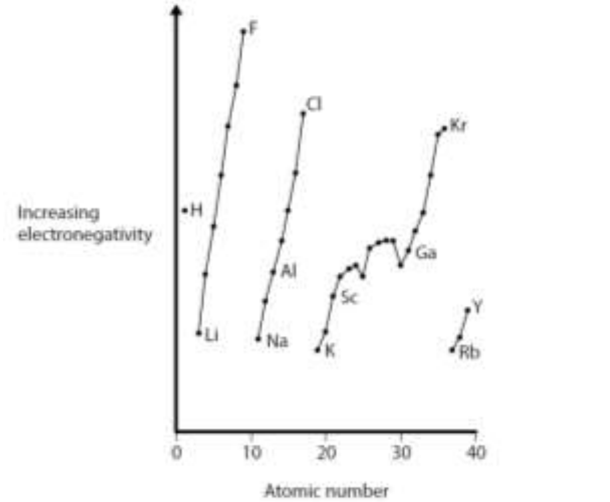
ELECTRONEGATIVITY: The ability of an atom in a molecule to attract a bonding pair of electrons.



Trends:

1. As you go down a group electronegativity decreases.
2. As you go across a period from left to right electronegativity increases.

Factors:

There is a stronger electronegativity …

1. …when the nuclear charge is higher due to more protons.
2. …when the atomic radius is smaller so the electrons are closer to the nucleus.
3. …when there are less shells of electrons between the nucleus & the outer electrons causing shielding.

Intermolecular Force: Forces between different molecules.	Van-der-Waals: Any intermolecular attraction.
Permanent Dipole-Dipole Force: Attraction between oppositely charged ends of polar covalent molecules.	Induced Dipole-Dipole: Attraction between oppositely charged ends of covalent molecules due to proximity to polar covalent molecules.
Instantaneous Dipole: When electrons are not equally spread out inducing a dipole in other molecules.	Hydrogen Bonding: Bonds form between H and F / N / O as large difference in electronegativity.

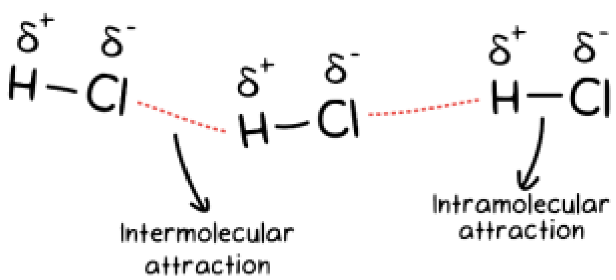
Alkanes e.g. Methane (CH_4)
Induced dipole-dipole forces.

Alcohols e.g. Methanol (CH_3OH)
Induced dipole-dipole forces.
One hydrogen bond per molecule.

Water
Induced dipole-dipole forces.
Permanent dipole-dipole forces.
Two hydrogen bonds per molecule.

Water has highest boiling point as it has the strongest intermolecular forces.

As the molecules get bigger the intermolecular forces get bigger increasing boiling point.



Period 2	Li	Be	B	C	N_2	O_2	F_2	Ne
Period 3	Na	Mg	Al	Si	P_4	S_8	Cl_2	Ar
Structure	giant metallic			giant covalent	simple molecular			
Forces	strong forces between positive ions and negative delocalised electrons			strong forces between atoms	weak intermolecular forces between molecules			
Bonding	metallic bonding			covalent	covalent bonding within molecules intermolecular bonding between molecules			

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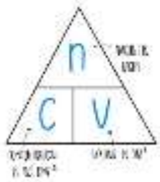
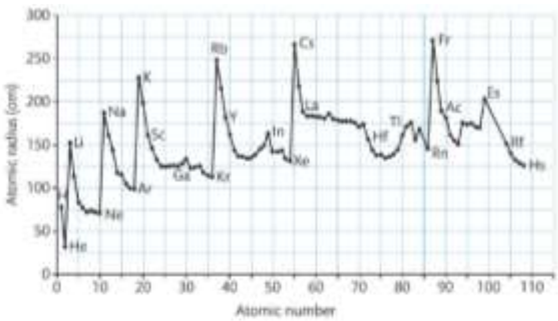
	Calculate the relative formula mass of:	
	Ammonia: NH_3 = 14.0 + (3x1.0) = 17.0	Sulfuric Acid: H_2SO_4 = (2x1.0) + 32.1 + (4x16.0) = 98.1
	Calculate the mass of 3 moles of ammonia. = 3 x 17.0 = 51.0 g	How many moles are in 490.5g of sulfuric acid? n = m ÷ Mr n = 490.5 ÷ 98.1 = 5 moles
	Calculate the number of moles in $100cm^3$ of 0.5M nitric acid [2]. 100 ÷ 1000 = 0.1 dm^3 0.1 x 0.5 = 0.05 mol	Calculate the concentration of NaF (Mr = 46.0) if 21g is dissolved in $400cm^3$ water [3]. 21 ÷ 42 = 0.5 mol 400 ÷ 1000 = 0.4 dm^3 0.5 ÷ 0.4 = 1.25 M

Lithium reacts with chlorine to form lithium chloride: 2Li + Cl2 → 2LiCl 13.8 + 71 → 84.8	Iron reacts with oxygen to make Iron oxide: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$ 223.2 + 96 → 319.2
Calculate the mass of lithium chloride made from 3.45 g lithium. 13.8 ÷ 3.45 = 4 84.8 ÷ 4 = 21.4 g	Calculate the percentage yield if 54 kg is made & 90 kg is the theoretical yield. (54 ÷ 90) x 100 = 60%
Calculate the mass of chlorine used to make 16.96 g lithium chloride. 84.8 ÷ 16.96 = 5 71 ÷ 5 = 14.2 g	Calculate the percentage yield if 288 kg O_2 used & actual yield is 478.8 kg 288 ÷ 96 = 3 3 x 319.2 = 957.6 theoretical yield (478.8 ÷ 957.6) x 100 = 50%
Calculate the mass of lithium used to make 10.6 g lithium chloride. 84.8 ÷ 10.6 = 8 13.8 ÷ 8 = 1.725 g	$\% \text{ Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$

Atomic Radius: 	1. As you go down a group the atomic radius increases as there are extra electron shells. 2. As you go across a period from left to right the atomic radius decreases as there is a greater nuclear charge pulling the electrons closer.

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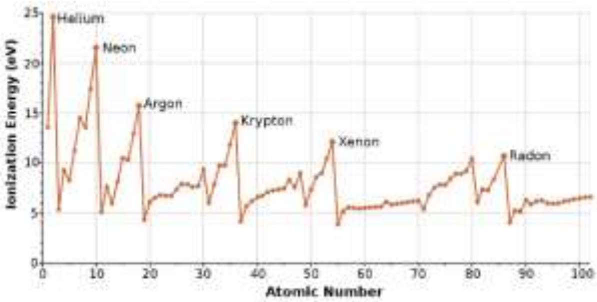
FIRST ELECTRON AFFINITY: Energy released when one mole of atoms gains one mole of electrons in their gaseous state ($kJmol^{-1}$).	**FIRST IONISATION ENERGY:** Energy required to remove one mole of electrons from one mole of atoms in their gaseous state ($kJmol^{-1}$).
$C_{(g)} + e^- \rightarrow C^-_{(g)}$	$B_{(g)} \rightarrow B^+_{(g)} + e^-$
$N_{(g)} + e^- \rightarrow N^-_{(g)}$	$F_{(g)} \rightarrow F^+_{(g)} + e^-$

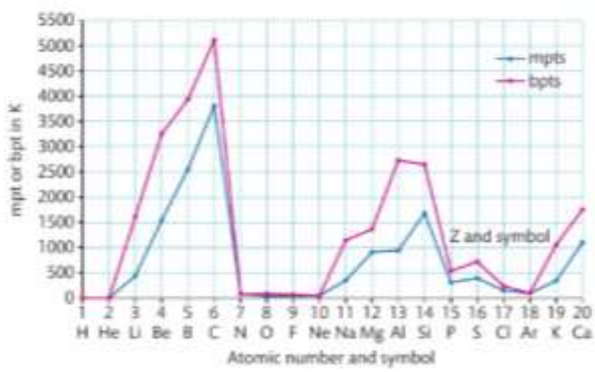
First Ionisation Energy:



1. As you go down a group the first ionisation energy decreases as there are extra electron shells.
2. As you go across a period from left to right the first ionisation energy increases as there is a greater nuclear charge attracting the electrons.

An increase means more energy is required to remove electrons as they are harder to remove.	A decrease means less energy is required to remove electrons as they are easier to remove.

Melting & Boiling Points:



As you go down a group the melting & boiling point…

- …for metals decreases as the metallic bond is easier to break so requires less energy to break.
- …for non-metals increases as there are more electrons so stronger intermolecular forces due to more temporary dipoles.

COMPLETE COMBUSTION: When a fuel burns fully in air to make water and carbon dioxide.	INCOMPLETE COMBUSTION: When a fuel burns without enough oxygen to make water and carbon monoxide or carbon.
Group 1: $4Li + O_2 \rightarrow 2Li_2O$	Group 4: $C + O_2 \rightarrow CO_2$
Group 2: $2Ca + O_2 \rightarrow 2CaO$	Group 5: $N_2 + 2O_2 \rightarrow 2NO_2$
Group 3: $4Al + 3O_2 \rightarrow 2Al_2O_3$	Group 6: $S + O_2 \rightarrow SO_2$

Iron + sulfuric acid → iron sulfate + hydrogen	
Sodium hydroxide + hydrochloric acid → sodium chloride + water	
Calcium oxide + nitric acid → calcium nitrate + water	
Zinc carbonate + hydrochloric acid → zinc chloride + water + carbon dioxide	
$Fe + H_2SO_4 \rightarrow FeSO_4 + H_2$	$CaO + 2HNO_3 \rightarrow Ca(NO_3)_2 + H_2$
$NaOH + HCl \rightarrow NaCl + H_2O$	$ZnCO_3 + 2HCl \rightarrow ZnCl_2 + H_2$

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Lithium + water → lithium hydroxide + hydrogen $2Li + 2H_2O \rightarrow 2LiOH + H_2$	Sodium + water → sodium hydroxide + hydrogen $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Why is K more reactive than Na?
K, Na in group 1;
K in period 4, Na in period 3;
Electronic structure for K: 2,8,8,1, for Na 2,8,1;
One electron in the outer shell to remove to gain;
One electron to remove to gain a stable structure;
Outer electron for K shell 4, Na shell 3;
K has a larger atomic radius, then Na;
Outer electron further from nucleus for K than Na;
K has a larger amount of shielding from inner shells than Na;
K is easier to remove the outer electron than Na;

Why is Na more reactive than Mg?
Na in group 1, Mg in group 2;
Na & Mg both in period 3;
Electronic structure for Na: 2,8,1, Mg: 2,8,2;
Na has one electron in outer shell, Mg has two;
Na needs to lose 1 electron to gain a stable structure, Mg 2 electrons;
Removal of 2 electrons requires more energy than the removal of 1;
Same amount of shielding in Na & Mg;
Mg has an increased nuclear charge over Na whilst being in the same period;
Electrons in same s subshell are more strongly attracted to nucleus in Mg;
More energy needed to remove the electrons;

Write the balanced equation for the reaction of:
Sodium chloride & silver nitrate
$NaCl + AgNO_3 \rightarrow NaNO_3 + AgCl$

<u>DOUBLE DISPLACEMENT REACTION:</u>
A chemical reaction where elements in different compounds exchange places.

metal sulfate solution / metal	magnesium sulfate	zinc sulfate	iron sulfate	copper sulfate
magnesium		✓	✓	✓
zinc	✗		✓	✓
iron	✗	✗		✓
copper	✗	✗	✗	

Order of reactivity:
Most = Mg, Zn, Fe, Cu = least
The more displacement reactions the more reactive the metal is.

Zinc + magnesium sulfate → no reaction
$Zn + CuSO_4 \rightarrow$ no reaction
Zinc + copper sulfate → zinc sulfate + copper
$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$
Zinc + iron sulfate → zinc sulfate + iron
$Zn + FeSO_4 \rightarrow ZnSO_4 + Fe$

Explain in terms of oxidation & reduction what happens in the iron & copper sulfate reaction.
Displacement reaction;
$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$;
Iron is more reactive than copper;
So displaces copper; Copper forms on nail / is the brown coating;
Copper ions remove electrons from iron; $Fe + Cu^{2+} \rightarrow Fe^{2+} + Cu$;
Iron atoms are the reducing agent;
Iron atoms lose electrons;
$Fe \rightarrow Fe^{2+} + 2e^-$;
Copper ions are the oxidising agent;
Copper ions gain electrons;
$Cu^{2+} + 2e^- \rightarrow Cu$;

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Oxidation
1. Gain of oxygen
2. Loss of electrons
3. Decrease in oxidation state

Reduction
1. Loss of oxygen
2. Gain of electrons
3. Increase in oxidation state

Write the half equations for magnesium reacting with oxygen.

Oxidation
$2Mg \rightarrow 2Mg^{2+} + 4e^-$

Reduction
$O_2 + 4e^- \rightarrow 2O^{2-}$

Why is the reaction of iron oxide with aluminium a redox reaction?
(redox reactions) involve oxidation & reduction / increase & decrease in oxidation state;
Oxidation is {loss of electrons / gain of oxygen / increase in oxidation state};
Reduction is {gain of electrons / loss of oxygen / decrease in oxidation state};
Aluminium {gains oxygen / goes from oxidation state 0 to +3 / is oxidised / loses electrons};
Iron oxide loses oxygen; OR
Iron (in iron oxide) {goes from oxidation state +3 to 0 / is reduced / gains electrons};

What is the oxidation state of…	
Sodium: +1 as it loses 1 electron	Nitrogen: -3 as it gains 3 electrons.
Calcium: +2 as it loses 2 electrons	Oxygen: -2 as it gains 2 electrons.
Aluminium: +3 as it loses 3 electrons	Fluorine: -1 as it gains 1 electrons.
Carbon: +4 as it gains 4 electrons	Neon: 0 as it is electronically stable.

Ammonium = $NH4^+$ Sulphate = SO_4^{2-}	N has oxidation state = N^{3-} S has oxidation state = S^{6+}

Order of reactivity:
Most = chlorine, bromine, iodine = least

Chlorine displaces bromine & iodine;
Bromine displaces iodine but not chlorine;
Iodine doesn't displace bromine or chlorine;

halogen \ salt (aq)	potassium chloride	potassium bromide	potassium iodide
chlorine		$2KCl + Br_2$	$2KCl + I_2$
bromine	no reaction		$2KBr + I_2$
iodine	no reaction	no reaction	

Where are the transition metals?
d block because the last electron to fill the atom is placed into a d orbital

What oxidation states do they have?
Form more than one stable ion with an incomplete d subshell.

Why do they act as catalysts?
They have the ability to absorb small molecules onto their surface.

Why do they have different colours?
Have variable oxidation states that give different coloured solutions.

Why is zinc not a transition metal?
Does not have an incomplete d subshell;
Only forms one oxidation state +2;
Only forms one ion;
Only loses electrons from 4s subshell to form an ion;
+2 is [Ar] $3d^{10}$ / an ion with a complete d subshell;
Doesn't form an ion with an incomplete d subshell;
Is not a catalyst & forms white compounds;

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The Periodic Table of Elements

Key

relative atomic mass
atomic symbol
name
atomic (proton) number

1 (1)	2 (2)	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(12)	3 (13)	4 (14)	5 (15)	6 (16)	7 (17)	0 (8) (18)
							1.0 **H** hydrogen 1										4.0 **He** helium 2
6.9 **Li** lithium 3	9.0 **Be** beryllium 4											10.8 **B** boron 5	12.0 **C** carbon 6	14.0 **N** nitrogen 7	16.0 **O** oxygen 8	19.0 **F** fluorine 9	20.2 **Ne** neon 10
23.0 **Na** sodium 11	24.3 **Mg** magnesium 12											27.0 **Al** aluminium 13	28.1 **Si** silicon 14	31.0 **P** phosphorus 15	32.1 **S** sulfur 16	35.5 **Cl** chlorine 17	39.9 **Ar** argon 18
39.1 **K** potassium 19	40.1 **Ca** calcium 20	45.0 **Sc** scandium 21	47.9 **Ti** titanium 22	50.9 **V** vanadium 23	52.0 **Cr** chromium 24	54.9 **Mn** manganese 25	55.8 **Fe** iron 26	58.9 **Co** cobalt 27	58.7 **Ni** nickel 28	63.5 **Cu** copper 29	65.4 **Zn** zinc 30	69.7 **Ga** gallium 31	72.6 **Ge** germanium 32	74.9 **As** arsenic 33	79.0 **Se** selenium 34	79.9 **Br** bromine 35	83.8 **Kr** krypton 36
85.5 **Rb** rubidium 37	87.6 **Sr** strontium 38	88.9 **Y** yttrium 39	91.2 **Zr** zirconium 40	92.9 **Nb** niobium 41	95.9 **Mo** molybdenum 42	[98] **Tc** technetium 43	101.1 **Ru** ruthenium 44	102.9 **Rh** rhodium 45	106.4 **Pd** palladium 46	107.9 **Ag** silver 47	112.4 **Cd** cadmium 48	114.8 **In** indium 49	118.7 **Sn** tin 50	121.8 **Sb** antimony 51	127.6 **Te** tellurium 52	126.9 **I** iodine 53	131.3 **Xe** xenon 54
132.9 **Cs** caesium 55	137.3 **Ba** barium 56	138.9 **La*** lanthanum 57	178.5 **Hf** hafnium 72	180.9 **Ta** tantalum 73	183.8 **W** tungsten 74	186.2 **Re** rhenium 75	190.2 **Os** osmium 76	192.2 **Ir** iridium 77	195.1 **Pt** platinum 78	197.0 **Au** gold 79	200.6 **Hg** mercury 80	204.4 **Tl** thallium 81	207.2 **Pb** lead 82	209.0 **Bi** bismuth 83	[209] **Po** polonium 84	[210] **At** astatine 85	[222] **Rn** radon 86
[223] **Fr** francium 87	[226] **Ra** radium 88	[227] **Ac*** actinium 89	[261] **Rf** rutherfordium 104	[262] **Db** dubnium 105	[266] **Sg** seaborgium 106	[264] **Bh** bohrium 107	[277] **Hs** hassium 108	[268] **Mt** meitnerium 109	[271] **Ds** darmstadtium 110	[272] **Rg** roentgenium 111							

Elements with atomic numbers 112-116 have been reported but not fully authenticated

* Lanthanide series	140 **Ce** cerium 58	141 **Pr** praseodymium 59	144 **Nd** neodymium 60	[147] **Pm** promethium 61	150 **Sm** samarium 62	152 **Eu** europium 63	157 **Gd** gadolinium 64	159 **Tb** terbium 65	163 **Dy** dysprosium 66	165 **Ho** holmium 67	167 **Er** erbium 68	169 **Tm** thulium 69	173 **Yb** ytterbium 70	175 **Lu** lutetium 71
* Actinide series	232 **Th** thorium 90	[231] **Pa** protactinium 91	238 **U** uranium 92	[237] **Np** neptunium 93	[242] **Pu** plutonium 94	[243] **Am** americium 95	[247] **Cm** curium 96	[245] **Bk** berkelium 97	[251] **Cf** californium 98	[254] **Es** einsteinium 99	[253] **Fm** fermium 100	[256] **Md** mendelevium 101	[254] **No** nobelium 102	[257] **Lr** lawrencium 103