Please check the examination details below before entering your candidate information

Candidate surname	Other names
StemUp Model Solution	

Pearson Edexcel Level 3 GCE

Centre Number	Candidate Number

Time 1 hour 30 minutes

Paper reference **8CH0/01**

Chemistry

Advanced Subsidiary

PAPER 1: Core Inorganic and Physical Chemistry

Candidates must have: Scientific calculator
Data Booklet

Total Marks

Instructions

- Use **black** ink or **black** ball-point pen.
- **Fill in the boxes** at the top of this page with your name, centre number and candidate number.
- Answer **all** questions.
- Answer the questions in the spaces provided
  – *there may be more space than you need.*

Information

- The total mark for this paper is 80.
- The marks for **each** question are shown in brackets
  – *use this as a guide as to how much time to spend on each question.*
- For the question marked with an **asterisk** (*), marks will be awarded for your ability to structure your answer logically, showing the points that you make are related or follow on from each other where appropriate.
- A Periodic Table is printed on the back cover of this paper.

Advice

- Read each question carefully before you start to answer it.
- Show all your working in calculations and include units where appropriate.
- Check your answers if you have time at the end.
- Good luck with your examination.

Turn over ▶

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P 6 7 0 8 3 A 0 1 2 4

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Answer ALL questions.

Some questions must be answered with a cross in a box ☒. If you change your mind about an answer, put a line through the box ☒ and then mark your new answer with a cross ☒.

1 Bromine exists as two stable isotopes. The two isotopes are represented by the symbols $^{79}_{35}Br$ and $^{81}_{35}Br$.

(a) Give one similarity and one difference between these two isotopes by referring to the **number of particles** in the nuclei of the two isotopes.

(2)

Similarity: both isotopes have the same number of protons, 35 protons in each nucleus.

Difference: they have different numbers of neutrons: bromine-79 has 44 neutrons and bromine-81 has 46 neutrons (two more neutrons).

(b) The relative abundance of the two isotopes in a sample cannot be found in a chemical test.

(i) Give the reason why, despite the difference in atomic structure, the isotopes have the same chemical reactions.

(1)

They have the same electronic configuration (both have 35 electrons arranged in the same way), and chemical reactions depend on the electrons.

(ii) State how the relative abundance of the two isotopes can be found.

(2)

Use a mass spectrometer.

Compare (measure) the intensity of the signal, i.e. the relative number of particles of each isotope detected, at m/z = 79 and m/z = 81.

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(c) (i) Complete the electronic configuration of a bromine atom.

(1)

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

(ii) What is the number of electrons in the fourth quantum shell of bromine?

(1)

- ☒ **A** 2
- ☒ **B** 7
- ☒ **C** 17
- ☒ **D** 18

(Total for Question 1 = 7 marks)

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2 This question is about sodium carbonate.

(a) Sodium carbonate forms a number of hydrates with the general formula $Na_2CO_3.xH_2O$.

A 250 cm^3 standard solution of one of these hydrates contained 10.0 g of the compound.

Describe, including the names of any relevant apparatus, how to make this standard solution when provided with 10.0 g of the hydrate in a beaker.

(5)

Dissolve the 10.0 g of hydrate in the beaker in a small volume (much less than 250 cm^3) of distilled (deionised) water, stirring with a glass rod.

Pour the solution into a 250 cm^3 volumetric flask using a funnel.

Rinse the beaker, glass rod and funnel with distilled water and transfer all the washings into the volumetric flask.

Add distilled water until the bottom of the meniscus is on the mark (line) on the neck of the flask, so the volume is exactly 250 cm^3; add the last few drops with a dropping pipette.

Stopper the flask and invert / shake it several times to mix the solution thoroughly.

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(b) 25.0 cm^3 portions of the standard solution described in (a) are titrated with hydrochloric acid solution of concentration 0.300 mol dm^{-3}, using methyl orange as an indicator.

The table shows the results for this titration.

	Titration **1**	Titration **2**	Titration **3**
Final volume / cm^3	30.25	29.75	31.25
Initial volume / cm^3	0.30	0.90	2.60
Total titre / cm^3	29.95	28.85	28.65

(i) What is the colour change at the end-point of the reaction?

(1)

	From	To
☒ **A**	red	orange
☒ **B**	red	yellow
☒ **C**	yellow	orange
☒ **D**	yellow	red

(ii) State why the value for the total titre in Titration **1** should not be used to calculate the mean titre.

(1)

It is not concordant with the other two results: it is not within 0.20 cm^3 of titrations 2 and 3 (it is over 1 cm^3 larger).

(iii) Calculate the mean titre.

(1)

mean titre = (28.85 + 28.65) ÷ 2

= 28.75 cm^3

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(iv) Calculate the relative formula mass, M_r, of the hydrated sodium carbonate, $Na_2CO_3.xH_2O$.

The equation for the reaction in the titration is

$$Na_2CO_3 + 2HCl \rightarrow 2NaCl + H_2O + CO_2$$

(4)

$n(HCl) = 28.75 \times 0.300 \div 1000 = 8.625 \times 10^{-3}$ mol

$n(Na_2CO_3)$ in 25.0 $cm^3 = 8.625 \times 10^{-3} \div 2 = 4.3125 \times 10^{-3}$ mol

$n(Na_2CO_3)$ in 250 $cm^3 = 4.3125 \times 10^{-3} \times 10 = 0.043125$ mol

M_r = mass ÷ moles = 10.0 ÷ 0.043125

M_r of $Na_2CO_3.xH_2O$ = 232 (g mol^{-1})

(c) In an experiment, the M_r of a **different** hydrated sodium carbonate was found to be 286 g mol^{-1}.

(i) Calculate the relative formula mass of anhydrous sodium carbonate, Na_2CO_3.

(1)

$M_r(Na_2CO_3) = (2 \times 23.0) + 12.0 + (3 \times 16.0) = 106$

(ii) Calculate the number of molecules of water of crystallisation, x, for this hydrated sodium carbonate, $Na_2CO_3.xH_2O$.

(1)

$x = (286 - 106) \div 18.0 = 180 \div 18.0$

$x = 10$

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(d) Sodium carbonate is manufactured from sodium chloride in a two-stage process.

$$NaCl + NH_3 + CO_2 + H_2O \rightarrow NaHCO_3 + NH_4Cl$$
$$2NaHCO_3 \rightarrow Na_2CO_3 + H_2O + CO_2$$

Calculate the maximum mass of sodium carbonate, Na_2CO_3, which could be obtained from 500 kg of sodium chloride.

(3)

$n(NaCl) = 500\,000\ g \div 58.5 = 8547\ mol$

$2\ mol\ NaCl \rightarrow 2\ mol\ NaHCO_3 \rightarrow 1\ mol\ Na_2CO_3$

$n(Na_2CO_3) = 8547 \div 2 = 4273.5\ mol$

$mass\ of\ Na_2CO_3 = 4273.5 \times 106 = 452\,991\ g$

$maximum\ mass = 453\ kg$

(Total for Question 2 = 17 marks)

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3 Ammonia reacts with sodium to form sodium amide, $NaNH_2$, and hydrogen.

(a) (i) Write the equation for this reaction.
State symbols are not required.

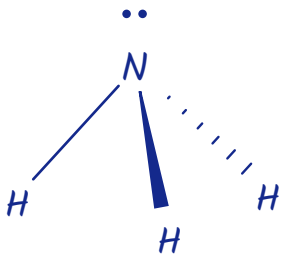
(1)

$2Na + 2NH_3 \rightarrow 2NaNH_2 + H_2$

(ii) Draw diagrams showing the 3-dimensional shape of an ammonia molecule and of an amide ion, NH_2^-.
Include any lone pairs of electrons in each species.

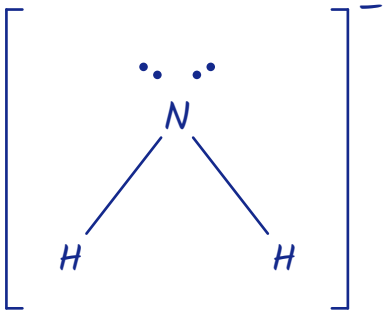
(3)

ammonia molecule



107°

amide ion



104.5°

(lone pairs shown as pairs of dots)

(iii) What is the H—N—H bond angle in an ammonia molecule?

(1)

☒ **A** 104.5°

☒ **B** 107°

☒ **C** 109.5°

☒ **D** 120°

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(iv) Explain the difference between the H—N—H bond angle in ammonia and in the amide ion.

(2)

The H-N-H bond angle is smaller in the amide ion (about 104.5°) than in ammonia (107°).

The amide ion has two lone pairs of electrons on nitrogen but ammonia has only one.

Lone pairs repel more strongly than bonding pairs (lone pair-lone pair repulsion is greater than lone pair-bonding pair repulsion), so the two lone pairs push the bonding pairs closer together.

(b) Give a possible reason why samples of sodium amide are stored in oil.

(1)

Sodium amide reacts vigorously with water (moisture) and oxygen in the air, so the oil keeps air and water away from it.

(Total for Question 3 = 8 marks)

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4 Sulfur is a bright yellow crystalline solid at room temperature.

Sulfur forms rings of 8 sulfur atoms so the formula of the yellow solid is S_8.

(a) A section of a periodic table showing values of first ionisation energy in kJ mol^{-1} is shown.

N 1400	O 1310	F 1680
P 1010	S 1000	Cl 1250
As 950	Se 940	Br 1140

(i) Which equation represents the first ionisation energy of sulfur?

(1)

☒ **A** $S(s) \rightarrow S^+(g) + e^-$

☒ **B** $S_8(s) \rightarrow S_8^+(g) + e^-$

☒ **C** $S(g) \rightarrow S^+(g) + e^-$

☒ **D** $S_8(g) \rightarrow S_8^+(g) + e^-$

(ii) Explain the trend in the values of the first ionisation energies for the group containing sulfur.

(3)

The first ionisation energy decreases down the group from oxygen to selenium, even though the number of protons (nuclear charge) increases.

This is because the electron being removed is in a shell further from the nucleus (one extra shell each time).

There is also more shielding by the extra inner shell of electrons, so the outer electron is less strongly attracted to the nucleus and less energy is needed to remove it.

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(iii) Explain why the first ionisation energy of sulfur is lower than that of chlorine.

(2)

Sulfur has a smaller nuclear charge than chlorine: it has one fewer proton (16 compared with 17).

The electron removed is from the same 3p sub-shell in both atoms, so it has similar shielding and distance, and is attracted less strongly in sulfur.

(iv) Explain why the first ionisation energy of sulfur is lower than that of phosphorus.

(2)

In sulfur the electron is removed from a 3p orbital containing two electrons: spin pairing has occurred for the first time in the 3p sub-shell ($3p^4$).

The repulsion between the two paired electrons in that orbital means one of them is lost more easily than an unpaired 3p electron in phosphorus ($3p^3$).

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(b) Compound **X** is an oxide of sulfur. A gaseous sample of 0.318 g of **X** occupied a volume of 132 cm^3 at a temperature of 420 K and pressure of 105 kPa.

The number of moles of a gas and the volume occupied by it can be found using the ideal gas equation

$$pV = nRT$$

Calculate the relative molecular mass of **X** and hence its molecular formula.
You must show **all** your working.

[$R = 8.31\ J\ mol^{-1}\ K^{-1}$]

(5)

$pV = nRT$ so $n = pV \div RT$

$p = 105\ kPa = 105\ 000\ Pa;\ V = 132\ cm^3 = 1.32 \times 10^{-4}\ m^3$

$n = (105\ 000 \times 1.32 \times 10^{-4}) \div (8.31 \times 420)$

$n = 13.86 \div 3490.2 = 3.9711 \times 10^{-3}\ mol$

$M_r = mass \div n = 0.318 \div 3.9711 \times 10^{-3} = 80.1$

SO_3: $32.1 + (3 \times 16.0) = 80.1$

relative molecular mass of X = 80.1; molecular formula = SO_3

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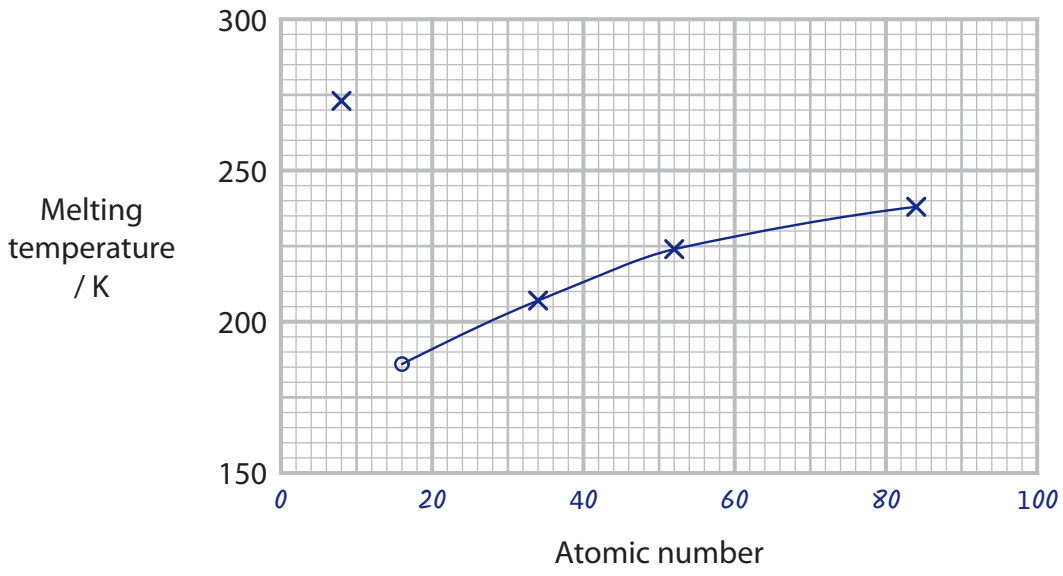
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(c) Sulfur and the other elements in Group 6 form dihydrogen compounds.

Compound	Atomic number of Group 6 element	Melting temperature / K
H_2O	8	273
H_2S	16	To be estimated
H_2Se	34	207
H_2Te	52	224
H_2Po	84	238

(i) Plot a graph of atomic number of the Group 6 element on the *x*-axis against melting temperature of the dihydrogen compound on the *y*-axis.

(2)



(ii) Give an estimate of the melting temperature of H_2S.

(1)

About 186 K (between 170 K and 200 K)

(Total for Question 4 = 16 marks)

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5 This question is about crystalline solids.

(a) Iodine and diamond are crystalline solids at room temperature.

Explain why diamond has a much higher melting temperature than iodine.

(5)

Iodine has a simple molecular structure made up of I_2 molecules. Within each molecule the I-I covalent bond is strong, but the molecules are held together only by weak London forces (instantaneous dipole-induced dipole attractions).

Diamond has a giant covalent (lattice) structure: each carbon atom is joined to four other carbon atoms by strong covalent bonds in a three-dimensional network.

On melting iodine, only the weak intermolecular forces between the molecules have to be overcome, and the covalent bonds are not broken, so little energy is needed.

On melting diamond, many strong covalent bonds must be broken. Strong covalent bonds require much more energy to break than weak intermolecular forces, so diamond has a much higher melting temperature.

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(b) Graphite is also a crystalline solid at room temperature.
Unlike diamond, graphite conducts electricity.

Describe the key feature of the bonding of the carbon atoms in graphite that results in it being an electrical conductor.

(2)

Each carbon atom is covalently bonded to only three other carbon atoms, in flat hexagonal layers, using three of its four outer electrons.

The fourth electron from each carbon is delocalised (free to move) within the layers, so it can carry the current along the layers.

(Total for Question 5 = 7 marks)

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6 This question is about the reactions of the halogens and their salts.

(a) The potassium halides react with concentrated sulfuric acid to form hydrogen halides.

(i) The equation for this reaction for potassium chloride can be written

$$KCl + H_2SO_4 \rightarrow HCl + KHSO_4$$

The hydrogen chloride does not react further.

State why this reaction is not a redox reaction.

(1)

The oxidation number of every element stays the same (K +1, Cl −1, H +1, S +6, O −2), so there is no transfer of electrons.

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(ii) On descending Group 7, the hydrogen halides become better reducing agents.

Explain how the reactions of potassium chloride, potassium bromide and potassium iodide with concentrated sulfuric acid provide evidence for this statement.

No explanation of the trend is required.

(3)

With potassium chloride, only HCl (steamy fumes) is formed: hydrogen chloride does not reduce the sulfuric acid, so sulfur stays in the +6 oxidation state.

With potassium bromide, HBr is formed and it reduces the sulfuric acid to sulfur dioxide, SO_2, in which sulfur is +4. Brown fumes of Br_2 are also seen.

$2HBr + H_2SO_4 \rightarrow Br_2 + SO_2 + 2H_2O$

With potassium iodide, HI reduces the sulfur in sulfuric acid further, to sulfur (S, 0) and hydrogen sulfide, H_2S, in which sulfur is −2 (bad egg smell, yellow solid, purple I_2).

$8HI + H_2SO_4 \rightarrow 4I_2 + H_2S + 4H_2O$

So the sulfur is reduced further by HI than by HBr (and HCl), showing reducing power increases down the group.

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(b) The reaction that occurs between chlorine and sodium hydroxide depends on the temperature.

(i) At room temperature the reaction that occurs is

$$Cl_2 + NaOH \rightarrow NaClO + NaCl$$

Explain, with reference to oxidation numbers, why this is a disproportionation reaction.

(2)

Chlorine is simultaneously oxidised and reduced in the same reaction.

The oxidation number of chlorine changes from 0 in Cl_2 to +1 in NaClO (oxidation, increase of 1) and from 0 to −1 in NaCl (reduction, decrease of 1).

(ii) With hot sodium hydroxide solution, a different disproportionation reaction occurs. Sodium chlorate(V) is one of the products.

Complete the equation for this reaction. State symbols are not required.

(2)

3 Cl_2 + *6* NaOH → *$NaClO_3$ + 5NaCl + $3H_2O$*

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(c) Chlorine is used as a bleach in the textiles industry. Any excess chlorine can be removed by reduction to chloride ions.

The half-equation for the reaction of chlorine is

$$Cl_2 + 2e^- \rightarrow 2Cl^-$$

In one reaction, 768 cm^3 of chlorine gas was reduced.

(i) Calculate the number of moles of electrons gained by chlorine molecules during this reaction.
[Under these conditions one mole of gas occupies 24 dm^3]

(2)

$n(Cl_2) = 768 \div 24\,000 = 0.032$ mol

$n(\text{electrons}) = 0.032 \times 2$

$= 0.064$ mol (6.4×10^{-2} mol)

(ii) The reducing agent was a solution containing thiosulfate ions, $S_2O_3^{2-}$.
The chlorine reacted with 40 cm^3 of a 0.20 mol dm^{-3} solution of these ions.

Deduce the number of moles of electrons lost by each atom of sulfur in the thiosulfate ion, and hence the final oxidation state of the sulfur in the product.

(3)

$n(S_2O_3^{2-}) = 40 \times 0.20 \div 1000 = 0.0080$ mol

$n(\text{S atoms}) = 0.0080 \times 2 = 0.016$ mol

electrons lost per S atom $= 0.064 \div 0.016 = 4$

oxidation state of S in $S_2O_3^{2-} = +2$

each S loses 4 electrons: $+2 \rightarrow +6$

final oxidation state of sulfur = +6 (e.g. in SO_4^{2-})

(Total for Question 6 = 13 marks)

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7 The nitrates of lithium, rubidium and strontium are all white solids. The compounds are held together by ionic bonds.

(a) State the meaning of the term 'ionic bond'.

(2)

A (strong) electrostatic attraction between oppositely charged ions (positive cations and negative anions).

(b) What is the percentage by mass of strontium in strontium nitrate?

(1)

- ☒ **A** 38.0 %
- ☒ **B** 41.4 %
- ☒ **C** 58.6 %
- ☒ **D** 74.5 %

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(c) These three compounds cannot be identified with certainty from a flame test as the colours seen are similar.

Concentrated hydrochloric acid is used in a flame test procedure.

(i) Which of the following is a reason for dipping the flame test wire in concentrated hydrochloric acid during a flame test procedure?

(1)

- ☒ **A** it dissolves metal ions from the wire
- ☒ **B** it neutralises hydroxide ions that might colour the flame
- ☒ **C** it reduces the metal ions to metal atoms
- ☒ **D** it reacts with the compounds to form volatile chlorides

(ii) The flame colour given by these three solids in the flame test are shades of

(1)

- ☒ **A** green
- ☒ **B** lilac
- ☒ **C** red
- ☒ **D** yellow

(iii) What is the best explanation for why metal ions produce different flame colours?

(1)

Different wavelengths of light energy are

- ☒ **A** required to promote electrons to higher energy levels
- ☒ **B** released because electrons move from lower to higher energy levels
- ☒ **C** released due to different gaps between energy levels
- ☒ **D** required for electron transfer from non-metal ions to metal ions

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*(d) Devise a procedure to **identify** the nitrates of lithium, rubidium and strontium using the effect of heat on the three solids and any precipitation reactions of the compounds.

Practical details are not required, but you should give the observations expected in each case.

(6)

Label the three white solids P, Q and R. First heat a sample of each solid strongly in a test tube.

Lithium nitrate and strontium nitrate both decompose to give a brown gas (nitrogen dioxide, NO_2) and oxygen, which relights a glowing splint:

$4LiNO_3 \rightarrow 2Li_2O + 4NO_2 + O_2$ and $2Sr(NO_3)_2 \rightarrow 2SrO + 4NO_2 + O_2$

Rubidium nitrate gives no brown gas: it only gives off oxygen ($2RbNO_3 \rightarrow 2RbNO_2 + O_2$), so the solid giving no brown gas is identified as rubidium nitrate.

The two solids that gave brown gas must be lithium nitrate and strontium nitrate, so they need to be told apart with a precipitation test.

Dissolve a sample of each of these two solids in distilled (deionised) water to make solutions.

Add a few drops of dilute sulfuric acid (or sodium sulfate solution) to each solution.

Strontium nitrate gives a white precipitate of strontium sulfate: $Sr^{2+}(aq) + SO_4^{2-}(aq) \rightarrow SrSO_4(s)$, so this solid is strontium nitrate.

Lithium nitrate gives no precipitate (the solution stays colourless) because lithium sulfate is soluble, so the remaining solid is lithium nitrate.

(Rubidium nitrate would also give no precipitate, which confirms the result of the heating test.)

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(Total for Question 7 = 12 marks)

TOTAL FOR PAPER = 80 MARKS

The Periodic Table of Elements

1	2											3	4	5	6	7	0 (8) (18)
							1.0 H hydrogen 1										4.0 He helium 2
(1)	(2)											(13)	(14)	(15)	(16)	(17)	
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	(3)	(4)	(5)	(6)	(7)	(8)	(9)	(10)	(11)	(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							

Key

relative atomic mass
atomic symbol
name
atomic (proton) number

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

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
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

* Lanthanide series

* Actinide series

P 6 7 0 8 3 A 0 2 4 2 2 4

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