ATOMIC THEORY

1 Complete the following sentence:

Atoms contain a *positively* charged nucleus composed of *protons* and *neutrons*

2 Complete the table:

Particle	Relative Mass	Relative Charge
PROTON	1	+1
NEUTRON	1	0
ELECTRON	$5x10^{-4}$	-1

3 Complete the following sentence:

Electrons have a *negative* charge and are found *outside the nucleus in orbitals*

4 Virtually all the mass of an atom is due to the *protons and neutrons/nucleus*

5 Explain the terms *mass number* (A), *atomic number* (Z) and *isotope*.

mass number (A), total number of protons+neutrons in the nucleus of an atom

atomic number (Z) number of protons in an atom

isotope atoms of the same element (same number of protons) that have different mass numbers (different number of neutrons)

6 Write the symbol for the element which has A = 108 and Z = 47. $^{108}_{47}Ag$

7 State the number of protons, neutrons and electrons in each of the following:

^{65}Cu ^{73}Ge $^{15}N^{3-}$ $^{137}Ba^{2+}$

	protons	**neutrons**	**electrons**
^{65}Cu	29	36	29
^{73}Ge	32	41	32
$^{15}N^{3-}$	7	8	10
$^{137}Ba^{2+}$	56	81	54

8 Name the instrument which is used to find the isotopic composition of an element so that its relative atomic mass can be determined.

Mass spectrometer

9 Determine the relative atomic mass of copper given the following natural abundances:

^{63}Cu 76.0 % ^{65}Cu 24.0 %

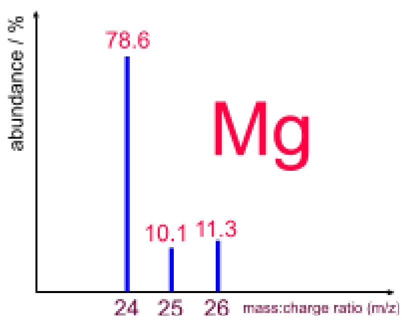
63.48

10 Determine the natural abundance of ^{11}B given that boron consists of two isotopes, ^{10}B and ^{11}B, and the relative atomic mass is 10.80.

80%

11 Use the mass spectrum shown to determine the relative atomic mass of magnesium to 1 decimal place.

24.3



12 Given that the relative atomic mass of iridium is 192.22 and that it has only 2 isotopes – ^{191}Ir and ^{193}Ir. Explain whether ^{191}Ir or ^{193}Ir is the more common isotope.

^{193}Ir is more common as the relative atomic mass is closer to 193 than 191
61% ^{193}Ir and 39% ^{191}Ir

13 State the regions of the electromagnetic spectrum.

radio waves	microwaves	Infrared	Visible light	Ultraviolet	X-rays	γ-rays

14 State the relative frequencies, energies and wavelengths of the regions of the electromagnetic spectrum.

increasing frequency →						
increasing energy →						
radio waves	microwaves	Infrared	Visible light	Ultraviolet	X-rays	γ-rays
decreasing wavelength →						

15 Arrange ***UV radiation*** ***blue light*** ***infrared radiation*** ***red light*** in order of:

(a) increasing frequency (lowest first) (b) decreasing wavelength (longest first)

(c) increasing energy (lowest first)

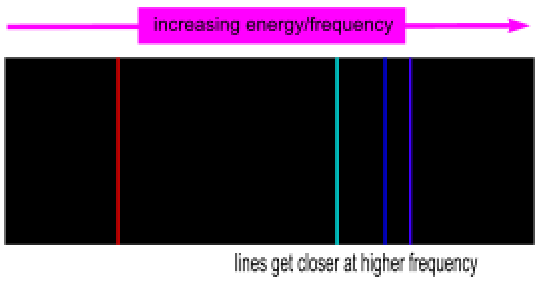
(a)	***infrared radiation***	***red light***	***blue light***	***UV radiation***
(b)	***infrared radiation***	***red light***	***blue light***	***UV radiation***
(c)	***infrared radiation***	***red light***	***blue light***	***UV radiation***

16 Distinguish between a continuous spectrum and a line spectrum

Continuous - all frequencies/wavelengths/colours present

Line – only certain discrete frequencies/wavelengths/colours present

17 Describe the emission spectrum of a hydrogen atom – draw a diagram (include at least 4 lines and label the direction in which frequency increases)



Series of lines of lines on a dark background
Lines get closer at higher frequency

18 Explain how a line in the emission spectrum arises.

Electron promoted to a higher energy level is unstable
Falls down to lower energy level
Energy given out as a photon of light
Energy of photon = energy difference between higher and lower levels

19 Explain how different series of lines arise.

Electrons fall down to different lower energy levels, e.g. to n=1, n=2 etc.

20 State whether each of the following transitions in the hydrogen emission spectrum would produce a line in the visible, infrared or ultra violet region of the electromagnetic spectrum.

$n = 5 \rightarrow n = 1$ $n = 4 \rightarrow n = 3$ $n = 6 \rightarrow n = 2$ $n = 10 \rightarrow n = 3$

$n = 5 \rightarrow n = 1$	$n = 4 \rightarrow n = 3$	$n = 6 \rightarrow n = 2$	$n = 10 \rightarrow n = 3$
UV	IR	Visible	IR

21 Select the highest energy transition in the hydrogen emission spectrum from the following list:

$n = 4 \rightarrow n = 2$ $n = 12 \rightarrow n = 3$ $n = 2 \rightarrow n = 1$ $n = 15 \rightarrow n = 2$

22 Convert each of the following wavelengths to a frequency in Hz.

500 nm 0.450 μm

6.00×10^{14} Hz 6.67×10^{14} Hz

23 Calculate the energy of the photon emitted for each of the following lines in the hydrogen atom spectrum. Planck's constant = 6.63×10^{-34} Js, speed of light = 3.00×10^{8} ms^{-1}

	Frequency /Hz	Wavelength	Energy / J
1	2.46×10^{15}	1.22×10^{-7} m	1.63×10^{-18}
2	4.57×10^{14}	6.56×10^{-7} m	3.03×10^{-19}
3	6.17×10^{14}	486 nm	4.09×10^{-19}
4	2.34×10^{14}	1.28×10^{-6} m	1.55×10^{-19}

24 State in which region of the electromagnetic spectrum each of the lines in 23 occurs (use the data booklet).

1	UV
2	Visible
3	Visible
4	IR

25 Explain what is meant by the *convergence limit* and what important information can be obtained from it.

Frequency/wavelength at which lines in the emission spectrum merge to form a continuum
Ionisation energy can be calculated from this

26 State an equation for the ionisation of hydrogen.

$H(g) \rightarrow H^{+}(g) + e^{-}$

27 Calculate the ionisation energy of hydrogen (in kJ mol^{-1}) given that the frequency of the convergence limit for the series of lines where the electron returns to the ground state (n=1) in the hydrogen atom spectrum is 3.30×10^{15} Hz.

1320 kJ mol^{-1}

28 Explain the term *orbital*.

Atomic orbital: region of space in an atom where there is a high probability of finding an electron

29 State the relative energies of s, p, d and f orbitals within any shell (main energy level).

s < p < d < f

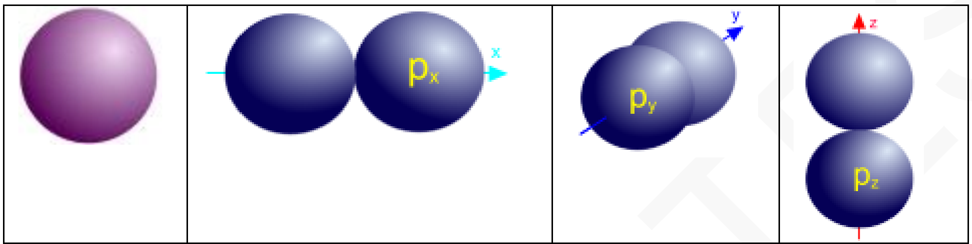
30 State the number of s, p, d, f orbitals within s, p, d, f subshells (sub-levels).

s	p	d	f
1	3	5	7

31 State the number of subshells (sub energy levels) and orbitals in the 4th main energy level (shell).

Subshells: 4 (s,p,d,f) Orbitals: 16 [s(1), p(3), d(5), f(7)]

32 Sketch the shape of an s and p_x, p_y and p_z orbitals.



s	p_x	p_y	p_z

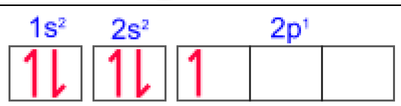
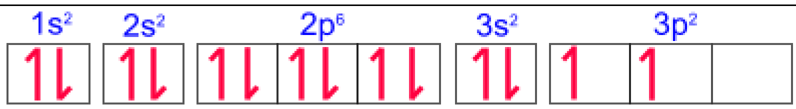
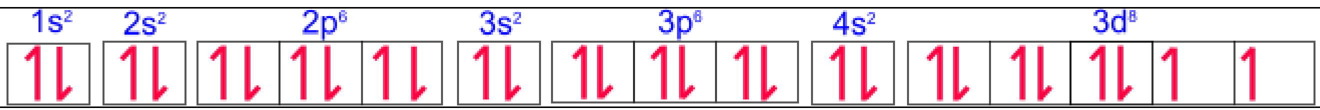
33 State the full electron configurations of: N P Ti Cr Fe Cu Se Kr

N	$1s^2\ 2s^2\ 2p^3$
P	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
Ti	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^2$
Cr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^5$
Fe	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$
Cu	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1\ 3d^{10}$
Se	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$
Kr	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

34 State the condensed electron configurations of: O Cl Mn As

O	$[He]\ 2s^2\ 2p^4$
Cl	$[Ne]\ 3s^2\ 3p^5$
Mn	$[Ar]\ 4s^2\ 3d^5$
As	$[Ar]\ 4s^2\ 3d^{10}\ 4p^3$

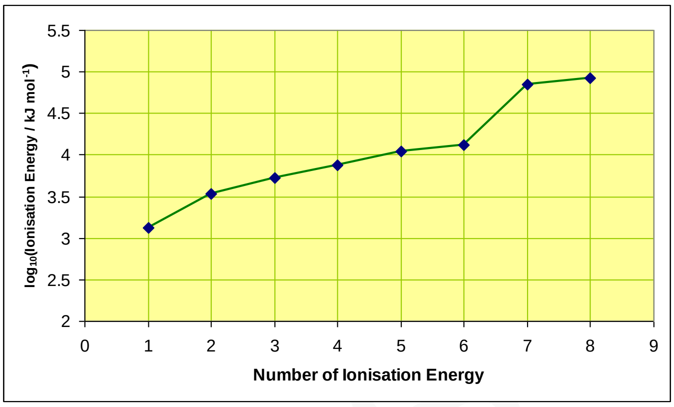
35 Draw orbital diagrams to represent the electron configuration of: B Si Ni

B	
Si	
Ni	

36 Write equations to represent the first and second ionisation energies of potassium.

1st IE $K(g) \rightarrow K^+(g) + e^-$
2nd IE $K^+(g) \rightarrow K^{2+}(g) + e^-$

37 Determine in which group of the Periodic Table an element is from a graph of successive ionisation energies such as



Group 16

First 6 electrons removed from outer shell – furthest from nucleus, more shielded

Big jump between 6th and 7th IEs – 7th electron removed from an inner shell – electron closer to nucleus, less shielded.

38 Determine which group an element is in from ionisation energy data such as

	Ionisation energy / kJ mol^{-1}							
Number of IE	**1**	**2**	**3**	**4**	**5**	**6**	**7**	**8**
X	786	1580	3230	4360	16000	20000	23600	29100
Y	1060	1900	2920	4960	6280	21200	25900	30500

X – Group 14 – 4e$^-$ in outer shell – big jump between 4th and 5th IEs
Y – Group 15 – 5e$^-$ in outer shell – big jump between 5th and 6th IEs

39 Explain how graphs of successive ionisation energies provide evidence for the existence of shells (main energy levels) and subshells (sub-levels).

Large jumps in ionisation energy indicate electrons being removed from different shells
Electrons from inner shells – closer to nucleus, less shielded – attracted more strongly by nucleus

Smaller jumps in ionisation energy indicate electrons being removed from different subshells – electron in a p subshell slightly easier to remove than one in an s subshell – the electron in a p subshell is higher in energy/slightly shielded by inner s electrons.

40 Explain why the second ionisation energy of an element always higher than the first ionisation energy

positive ion (2nd IE) attracts an electron more strongly than a neutral atom (1st IE) does.

OR: once one electron is removed the remaining electrons are pulled in more closely to nucleus (less electron-electron repulsion) – outer electron closer to the nucleus, more strongly attracted and harder to remove.

41 Explain why the second ionisation energy of potassium is substantially higher than the first ionisation energy?

First electron removed from outer shell – further from nucleus, more shielded by inner shells

Second electron removed from an inner shell – electron closer to nucleus, less shielded (fewer inner shells) – electron more strongly attracted to nucleus.

42 Explain why the first ionisation energy of O is higher than the first ionisation energy of Be.

O and Be – electron removed from second shell in each case – same number of inner shells (1) so amount of shielding approximately the same.

O – higher nuclear charge – outer electron attracted more strongly to nucleus, therefore harder to remove.

Also, O smaller, so outer electron closer to nucleus and more strongly attracted.

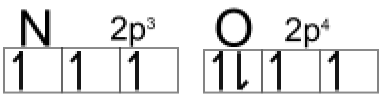
43 Explain why the first ionisation energy of B is lower than that of Be and the first ionisation energy of O is less than that of N.

B $1s^2\ 2s^2\ 2p^1$ Be $1s^2\ 2s^2$

Electron removed from p subshell in B but from s subshell in Be

p subshell in B higher in energy than s subshell in Be, therefore less energy required to remove electron in B

OR: p electron in B slightly more shielded than s electron in Be as the $2s^2$ electrons shield the p electron to a certain extent – more shielding, electron not as strongly attracted – easier to remove.



O – 2 electrons paired in the same p orbital – greater repulsion – easier to remove an electron. N – all p electrons are unpaired.

44 State the full electron configuration of the following ions: Mg^{2+} S^{2-} Fe^{2+} Cu^{2+} Ga^{3+}

Mg^{2+}	$1s^2\ 2s^2\ 2p^6$
S^{2-}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$
Fe^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$
Cu^{2+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^9$
Ga^{3+}	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$