WJEC Chemistry AS-level

1.2: Basic Ideas About Atoms

Practice Questions

England Specification

1.(a) The vast majority of motor vehicles worldwide are powered by petrol or diesel which come from crude oil. Give **two** reasons why we cannot rely indefinitely on oil as a source of transport fuel.

[2]

(b) Many vehicle manufacturers around the world have made the development of alternative fuels a priority. One such fuel being studied is hydrogen.
Its main advantage is that the only waste product is water, however hydrogen does not occur naturally on Earth. It is produced by passing an electric current through water.

(i)A leading car manufacturer said,

"Cars powered by hydrogen will be pollution-free".

Give **two** reasons why this is not necessarily true.

[2] QWC [1]

(ii) A spokesperson for a safety group said,

"Hydrogen can burn explosively. It must not be used in cars unless it is 100 % safe". State, giving a reason, whether you agree with this.

[1]

(c) The first line in the visible atomic emission spectrum for hydrogen has a wavelength of 656 nm, while that for helium has a wavelength of 707 nm.
State, giving a reason, which line has

(i) the higher frequency,

[1]

(ii) the higher energy.

[1]

(d) The first ionisation energy of helium is 2370 kJ mol-¹ while that of neon is 2080 kJ mol-1. Explain why neon has a lower first ionisation energy than helium.

[2]

(e) Another noble gas is radon. Its more stable isotope ^{222}Rn has a half-life of 3.8 days, decays by α-emission and is responsible for the majority of the public exposure to ionising radiation.

(i) Give the symbol and mass number of the atom formed by the loss of one α-particle from an atom of ^{222}Rn.

[1]

(ii) Explain why doctors are concerned that an over-exposure to radon may cause lung cancer.

[1]

(Total 12)

2.

The metal lead was one of the first in common use and even as far back as two thousand years ago, tens of thousands of tonnes of the metal were being produced every year in the Roman Empire. It is still in common use today, although many of its former uses have declined due to the toxic nature of the element.

(a) Lead is commonly extracted from lead(II) sulfide, PbS. Initially this ore is heated in a limited supply of air to produce lead(II) oxide, PbO, giving off sulfur dioxide gas, SO_2.

$$2PbS + 3O_2 \longrightarrow 2PbO + 2SO_2$$

If 20 kg of lead(II) sulfide were heated in air, calculate the mass of lead(II) oxide formed. [3]

.....

.....

.....

.....

Mass of lead(II) oxide formed = ... kg

(b) Metallic lead can then be obtained from lead(II) oxide by one of two methods:

Method 1: Reduction with a fresh supply of lead(II) sulfide in the absence of air

$$2PbO + PbS \longrightarrow 3Pb + SO_2$$

Method 2: Reduction by carbon monoxide in a blast furnace

$$PbO + CO \longrightarrow Pb + CO_2$$

(i) Both methods for producing lead release waste gases. Give an environmental problem associated with each of these gases. [2]

Sulfur dioxide, SO_2 .....

.....

Carbon dioxide, CO_2 .....

.....

(ii) The atom economy for producing lead by method 1 is 90.7 %.

I. Calculate the atom economy for producing lead by method 2. [2]

.....

.....

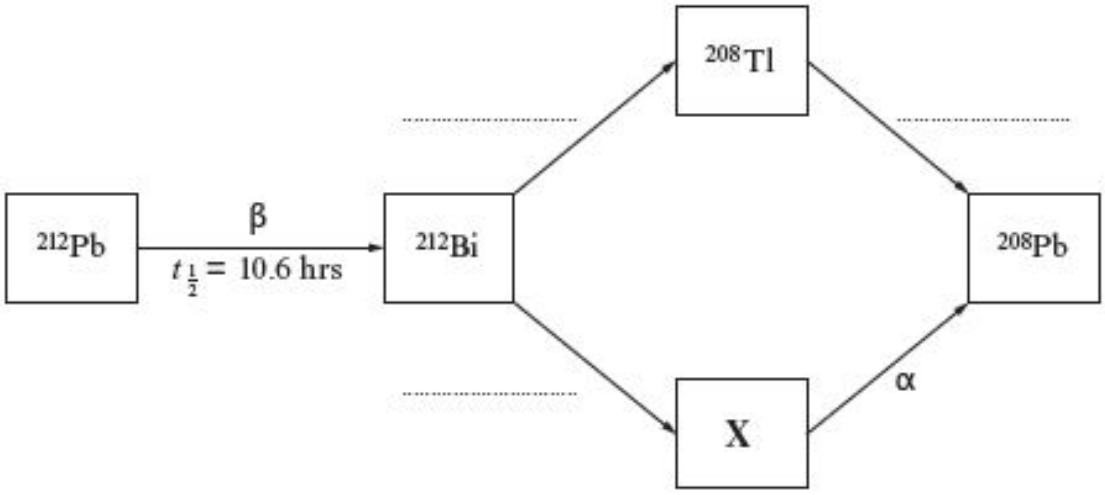
.....

II. Atom economy is one factor used in 'Green Chemistry' to assess the advantages and disadvantages of different routes to produce the same product. State, giving a reason, which of the two alternative methods would be considered to have the more advantageous atom economy. [1]

.....

.....

(c) Lead has a wide range of isotopes, some of which are stable and others that are radioactive. Radioactive lead-212 decays to eventually form the stable isotope ^{208}Pb. This process involves the decay of ^{212}Pb into ^{212}Bi followed by two alternative routes that both lead to ^{208}Pb, as shown in the scheme below.



(i) Give the correct symbol and mass number of the isotope indicated by **X** on the scheme above. [2]

Symbol ..... *Mass Number* .....

(ii) Two arrows have been labelled with α and β.
Label the remaining **three** arrows to indicate the nature of the radioactive decay occurring in each step. [2]

(iii) It is not possible to identify whether γ-radiation is also produced during any of the radioactive decay processes from the information given in the scheme.

State what is meant by γ-radiation and why it cannot be identified from the information given in the scheme. [2]

.....

.....

.....

(iv) A sample of 24 mg of ^{212}Pb was allowed to stand for 31.8 hours. Calculate the mass of ^{212}Pb that would remain after this time. [2]

.....

.....

..... mg

(d) Naturally-occurring lead consists of a mixture of stable isotopes which include ^{206}Pb, ^{207}Pb and ^{208}Pb. The relative amounts of these isotopes can vary between different sources. The abundance of each isotope in a sample is given below.

Isotope	Relative isotopic mass	Percentage abundance
^{206}Pb	206.0	25.48%
^{207}Pb	207.0	22.12%
^{208}Pb	208.0	52.40%

Calculate the relative atomic mass (A_r) for this sample of lead. Give your answer to **four significant figures**. [3]

.....

.....

.....

.....

Total [19]

3. This question is about atomic structure.

(a) Give the full electronic configuration of a nitrogen atom and use this to describe the way in which electrons are arranged in atoms.

[4] QWC [1]

(b) Describe the main features of the atomic emission spectrum of hydrogen in the visible region. Explain how these features arise and how their interpretation provides evidence for energy levels in the atom.

[6]

(c)(i) Hydrogen has a first ionisation energy of 1312 kJ mol^{-1}.Explain why helium has a higher first ionisation energy than hydrogen

[2]

(ii) Beryllium and magnesium are both in Group 2 of the Periodic Table.Explain why beryllium has a higher first ionisation energy than magnesium

[2]

(iii) The table below gives the first three ionisation energies for boron and potassium.

Element	Ionisation energy/kJ mol^{-1}		
	1st	2nd	3rd
B	800	2420	3660
K	419	3051	4412

(I) Suggest why compounds containing B^{3+} ions are unlikely to exist.

[1]

(II) Write an equation to represent the **second** ionisation energy of potassium

[1]

(III) State how the first three ionisation energies of calcium would differ from those of potassium.

[2]

(Total 19)

4. Magnesium is best known for burning with a characteristic brilliant white light, however in industry it is the third most commonly used structural metal. The metal itself was first produced by Sir Humphry Davy in 1808 by the electrolysis of a mixture of magnesia and mercury oxide.

(a) Magnesium has three stable isotopes ^{24}Mg, ^{25}Mg and ^{26}Mg.

(i) State the number of protons present in an atom of ^{24}Mg. [1]

(ii) Deduce the number of neutrons present in an atom of ^{26}Mg. [1]

(iii) In order to calculate the relative atomic mass of magnesium, what would you need to know in addition to the relative mass of each isotope? [1]

(b) Magnesium also has a radioactive isotope ^{28}Mg which has a half-life of 21 hours.

(i) If you started with 2.0 g of ^{28}Mg, calculate the mass of this isotope remaining after 84 hours. [1]

(ii) Name **one** useful radioactive isotope and briefly describe how it is used in medicine, industry or analysis. [2]

(c) In order to obtain a mass spectrum of a gaseous sample of magnesium, the sample must be ionised.

(i) State how the magnesium atoms are ionised in the sample.

[1]

(ii) Give a reason why it is necessary to ionise the magnesium atoms in the sample.

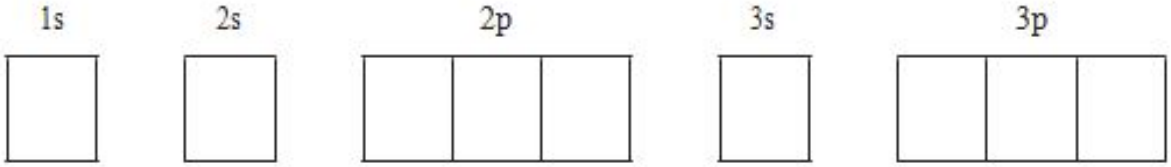
[1]

(iii) State how the ions of magnesium are separated.

[1]

(d) Magnesium reacts with nitrogen forming magnesium nitride, which is an ionic compound.

By inserting arrows to represent electrons, complete the boxes below to show the electronic configuration of a nitride ion, N^{3-}. [1]



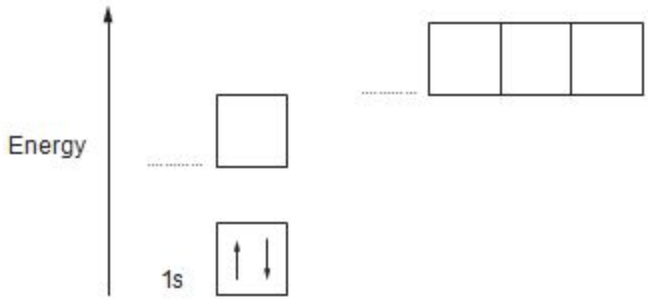
(e) Magnesium nitride reacts with water to form magnesium hydroxide and ammonia.

(i) Balance the equation above. [1]

(ii) Calculate the minimum mass of magnesium nitride required to form 1.75 g of magnesium hydroxide, giving your answer to **three** significant figures. [3]

(Total 14)

5. (a) Electrons are arranged in energy levels. The diagram below shows two electrons in the 1s level in a nitrogen atom.



Complete the diagram for the electrons in a nitrogen atom by labelling the sub-shell levels and showing how the electrons are arranged.

[2]

(b)Nitrogen forms several oxides.

(i) An oxide of nitrogen contains 25.9 % by mass of nitrogen. Calculate the empirical formula of this oxide.

[2]

Empirical formula . . . . . .

(ii) Dinitrogen oxide is formed when ammonia is oxidised

$$\ldots\ldots NH_3 + \ldots\ldots O_2 \longrightarrow \ldots\ldots N_2O + \ldots\ldots H_2O$$

Balance the equation above

[1]

(iii) Nitrogen dioxide is formed when calcium nitrate decomposes.

$$2Ca(NO_3)_2(s) \longrightarrow 2CaO(s) + 4NO_2(g) + O_2(g)$$

Calculate the total volume of gas, measured at room temperature and pressure, which would be produced when 0.886 g of calcium nitrate decomposes. [3]

[1 mol of gas occupies 24.0 dm^3 at room temperature and pressure]

Volume = ... dm^3

(c) Hydrated calcium nitrate can be represented by the formula $Ca(NO_3)_2.xH_2O$.

A 6.04 g sample of $Ca(NO_3)_2.xH_2O$ contains 1.84 g of water of crystallisation.

Calculate the value of x in $Ca(NO_3)_2.xH_2O$. You **must** show your working. [3]

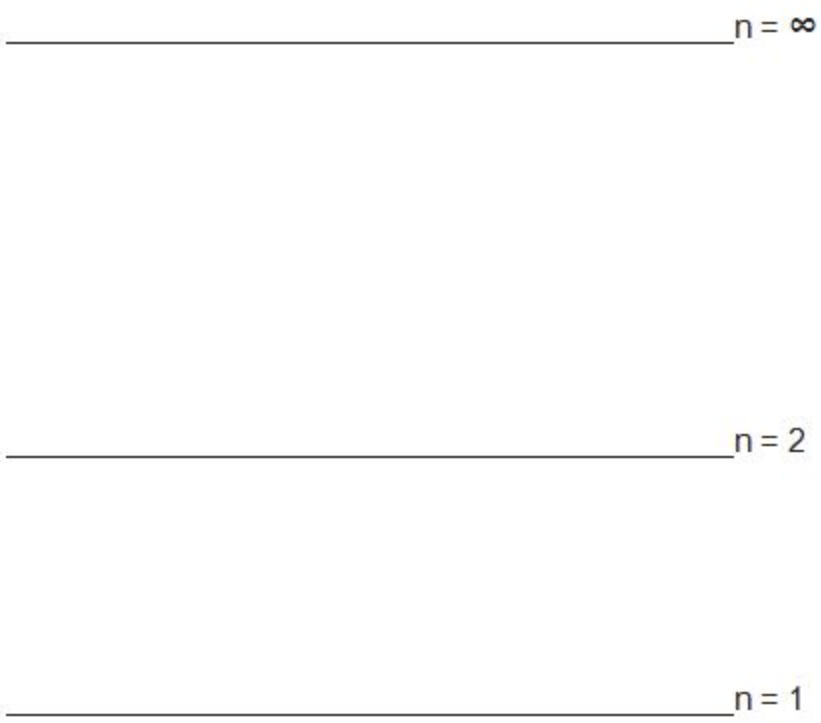
x = ...

Total [11]

6. (a) Hydrogen exists as three isotopes with relative masses of 1, 2 and 3.

State the similarities and differences in the composition of these specific isotopes. [2]

(b) The first two electronic energy levels in a hydrogen atom are shown on the diagram.



(i) Complete the diagram to show energy levels n = 3, n = 4 and n = 5.

[1]

(ii) Mark with an arrow the energy change corresponding to the ionisation energy of hydrogen.

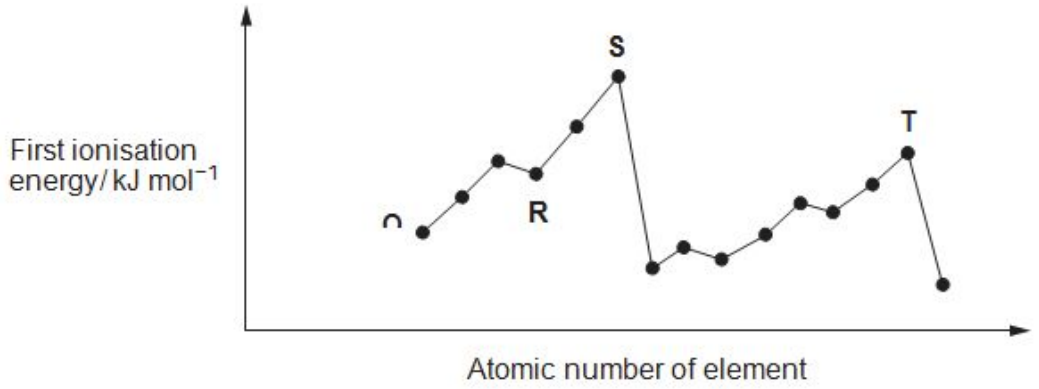
[2]

(c) A student said that the ionisation energy of hydrogen could be calculated using the Balmer Series of lines.

(i) In which part of the electromagnetic spectrum does the Balmer Series appear? [1]

(ii) Explain whether or not this student was correct. [2]

(d) The diagram shows part of a plot of the first ionisation energy of elements against their atomic numbers. Letters **Q**-**T** do **not** represent the symbols of the elements.



(i) Write the equation for the change occurring for the first ionisation energy of element **Q**.

[1]

(ii) In which group of the Periodic Table is element **R** found?

[1]

(iii) Explain why the first ionisation energy of **S** is greater than that of **T**.

[3] QWC [1]

(Total 14)

7. (*a*) The electronic structures of five atoms, **A** to **E**, are listed below. Arrange these atoms in order of increasing molar first ionisation energy.

[2]

Atom	A	B	C	D	E
Electronic structure	$1s^2$	$1s^2 2s^2$	$1s^2 2s^2 2p^1$	$1s^2 2s^2 2p^3$	$1s^2 2s^2 2p^6$

lowest ..... *highest*

(*b*) State, giving a reason for your choice, which **one** of the following gives the first four ionisation energies for silicon, Si.

[2]

	Ionisation energy / kJ mol^{-1}			
	1st	2nd	3rd	4th
W	496	4563	6913	9544
X	578	1817	2745	11578
Y	738	1451	7733	10541
Z	789	1577	3232	4356

Letter .....

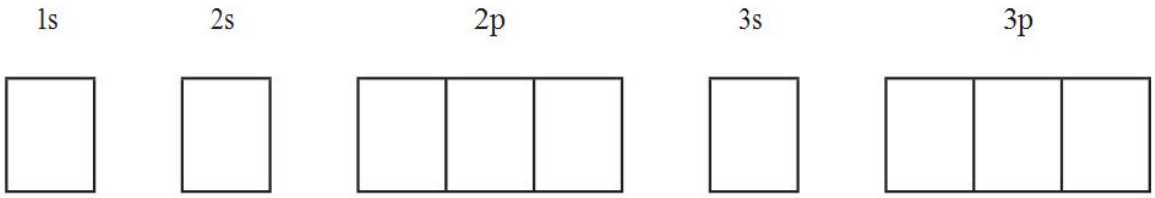
Reason .....

.....

.....

(c) By inserting arrows to represent electrons, complete the boxes below to show the electronic configuration of a sulfur atom.

[1]



(Total 5)

8.

(a) Write the letter corresponding to the correct electronic structure of an atom that is a member of the *d*-block in the box below. [1]

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

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

C $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

D $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

(b) Write the letter corresponding to the electronic structure of the atom with the highest first ionisation energy in the box below. [1]

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

B $1s^2 2s^2 2p^6$

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

D $1s^2 2s^2 2p^6 3s^2 3p^4$

(Total 2)

9.

(a) Potassium hydroxide contains potassium ions, K^+.
Give the electron configuration of a potassium **atom** and use this to explain why most potassium compounds contain the potassium ion. [2]

(b) Michael was asked to make 250 cm^3 of a solution of potassium hydroxide and to record the maximum rise in temperature that occurred as it dissolved.
He measured 250 cm^3 of water in a glass beaker and then added 7.01 g (0.125 mol) of solid potassium hydroxide to this, with stirring.
He noticed that the temperature rose from 20.2 °C to a maximum of 25.0 °C.

(i) Calculate the molar enthalpy change of solution of potassium hydroxide by use of the formula

$$\Delta H = -\frac{mc\Delta T}{n}$$

where m = mass of the solvent in grams (assume 1 cm^3 has a mass of 1 g)
c = 4.2 $J\,g^{-1}\,°C^{-1}$
ΔT = change in temperature of the solution
n = number of moles of the solute
ΔH = molar enthalpy change of solution

You should show the **units** in your answer. [3]

ΔH = ...

(ii) Michael's measurements produced a value for the enthalpy of solution of potassium hydroxide that was different to the literature value.

Use the information given to suggest and explain **two** factors that might produce a different result.

[2]

1. .....
.....

2. .....
.....

(c) Solid potassium hydroxide can be used in analysis to find the percentage of carbon dioxide present in a mixture of gases. The equation for the reaction that occurs is given below.

$$2KOH + CO_2 \longrightarrow K_2CO_3 + H_2O$$

2.0 m^3 of a gas mixture was passed through potassium hydroxide. Analysis showed that 1.50 mol of potassium carbonate had been formed.

(i) State the number of moles of carbon dioxide necessary to produce 0.050 mol of potassium carbonate. [1]

(ii) Calculate the volume of carbon dioxide that produced 0.050 mol of potassium carbonate. [1]

[1 mol of carbon dioxide has a volume of 24.0 dm^3 under these conditions]

Volume of carbon dioxide = ... dm^3

(iii) Calculate the percentage of carbon dioxide in the gas mixture, in terms of volume. [2]

[1 dm^3 = 0.001 m^3]

Percentage of carbon dioxide = ... %

(d) Scientists have commented that 'an increase in the amount of carbon dioxide dissolved in seawater will cause problems for animals whose shells are composed of calcium carbonate'.

$$CO_2(aq) + H_2O(l) + CaCO_3(s) \rightleftharpoons Ca^{2+}(aq) + 2HCO_3^-(aq)$$

Use the equation above to help you discuss the problem that is caused for these animals by this increase in carbon dioxide concentration.

[3] *QWC* [1]

(Total 15)

10. This question is about atomic structure.

(a) Give the full electronic configuration of a nitrogen atom and use this to describe the way in which electrons are arranged in atoms.

[4] QWC [1]

(b) Describe the main features of the atomic emission spectrum of hydrogen in the visible region. Explain how these features arise and how their interpretation provides evidence for energy levels in the atom.

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(c)(i) Hydrogen has a first ionisation energy of 1312 kJ mol^{-1}.Explain why helium has a higher first ionisation energy than hydrogen

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(I) Suggest why compounds containing B^{3+} ions are unlikely to exist. [1]

(II) Write an equation to represent the **second** ionisation energy of potassium. [1]

(III) State how the first three ionisation energies of calcium would differ from those of potassium. [2]

(Total 19)

11. In 2011 a man was detained at Moscow Airport when he tried to smuggle samples containing a radioactive isotope of sodium, ^{22}Na, onto an aircraft.

(i) This isotope is made from an aluminium isotope by loss of an α-particle.
State what is meant by an α-*particle*.

[1]

(ii) ^{22}Na decays by the loss of a positron. This may occur by the breakdown of a proton into a neutron and a positron, giving the product, ^{b}X.

Deduce the mass number (b) and the chemical symbol (X) of this product. [2]

b .....

X .....

(iii) The half-life of the isotope ^{22}Na is 2.6 years. The mass of a sample of this isotope is 48 mg.

Calculate the time taken for the mass of ^{22}Na to fall to 3 mg

[1]

Time taken = ...years

(b) The visible emission spectrum of sodium shows a strong yellow-orange line at a wavelength of 589 nm and a weaker green line at 569 nm.

Complete the sentences below by using the words **higher** or **lower** as appropriate.

The frequency of the green line at 569 nm is .....than the frequency of the yellow-

orange line at 589 nm. Another line is seen at 424 nm. This is caused by an electronic transition of .....energy than the line at 569 nm.

(c) Trona is a naturally-occurring 'sodium carbonate' mineral. It has the formula $Na_2CO_3.NaHCO_3.2H_2O$.

(i) Show that the relative molecular mass of trona is 226. [1]

(ii) On heating, trona loses water and carbon dioxide giving sodium carbonate.

$$2[Na_2CO_3.NaHCO_3.2H_2O](s) \longrightarrow 3Na_2CO_3(s) + CO_2(g) + 5H_2O(l)$$

Calculate the atom economy of this reaction, assuming that sodium carbonate is the only required product. [2]

Atom economy = ... %

(iii) The above reaction is used commercially to obtain sodium carbonate.

Suggest **one** environmental disadvantage of this reaction as indicated by the equation, and state what could be done to overcome this problem. [2]

(d) When sodium carbonate is added to water, some of the carbonate ions react with the water to give an alkaline solution.

$$CO_3^{2-}(aq) + H_2O(l) \rightleftharpoons HCO_3^{-}(aq) + OH^{-}(aq)$$

(i) Explain why this reaction is considered to be an acid-base reaction. [2]

(ii) The pH of a sodium carbonate solution is 11.4.
How would you explain the meaning of the pH scale to a member of the public? [3]

Total [15]