Assignment brief
Learning Outcomes and Assessment Criteria 2020-21

Sample number: 1

Unit Title: Chemical Science – RA1/3/AA/05G

Overview of assessment task

Electron configuration, bonding, metals and non-metals, chemical formulae and equations, periodicity

Students should complete the task to the best of their ability within the deadline set.

Learning outcomes and assessment criteria

Learning Outcomes	**Assessment Criteria**
1. To understand electron configuration within the atom, bonding, and main properties of metals and non-metals	1.1. Explain: (a) electronic configuration (b) chemical bonding (c) metals and non-metals (d) chemical formulae and equations
2. Understand the concepts of periodicity with reference to electronic configuration	2.1. Explain patterns associated with increasing atomic number, including: (a) first ionisation energies (b) atomic and ionic radii (c) boiling point (d) reaction with oxygen and chlorine (e) oxidation numbers

Grading Descriptors and Components 2020-21

Sample number: 1

Unit title: Chemical Science – RA1/3/AA/05G

GD 1: Understanding of the subject

Merit	**Distinction**
The student, student's work or performance: a. demonstrates a **very good** grasp of the relevant knowledge base	The student, student's work or performance: a. demonstrates an **excellent** grasp of the relevant knowledge base

GD 7: Quality

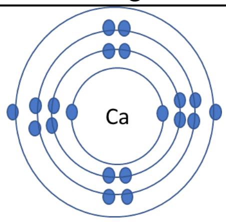
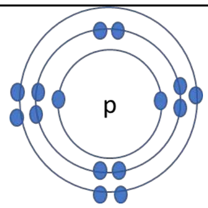
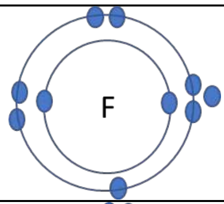
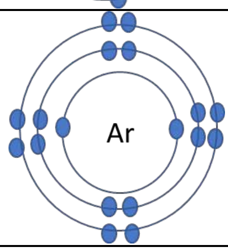
Merit	**Distinction**
The student, student's work or performance: a. is structured in a way that is **generally** logical and fluent	The student, student's work or performance: a. is structured in a way that is **consistently** logical and fluent

Chemical Science – RA1/3/AA/05G

You are working for a chemical company and have been asked to prepare some notes to help explain atomic structure and bonding to some apprentice technicians. Complete the following questions:

1. **For each of the following elements, write down the number of protons, neutrons and electrons; draw a Bohr diagram to show the electron configuration; and name the group that the element is in:**
   a. **Calcium**
   b. **Phosphorus**
   c. **Fluorine**
   d. **Argon**

A. Calcium $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
B. Phosphorus $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$
C. Fluorine $1s^2\ 2s^2\ 2p^5$
D. Argon $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

Element ^{6}	Electron	Proton	Neutron	Bohr diagram
Calcium	20	20	20	
Phosphorus	15	15	16	
Fluorine	9	9	10	
Argon	18	18	22	

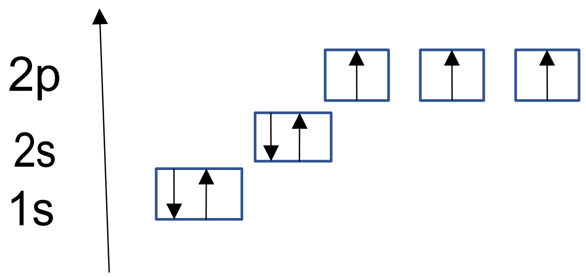
2. **Write the electronic subshell configuration (s, p, d) for the following:**
    a. **Sodium**
    b. **Chlorine**
    c. **Helium**
    d. **Potassium**

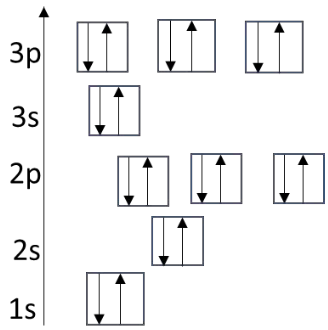
a. Sodium $1s^2\ 2s^2\ 2p^6\ 3s^1$
b. Chlorine $1s^2 2s^2 2p^6 3s^2$
c. Helium $1s^2$
d. Potassium $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$

3. **Draw energy level diagrams for nitrogen and Ca^{2+}.**

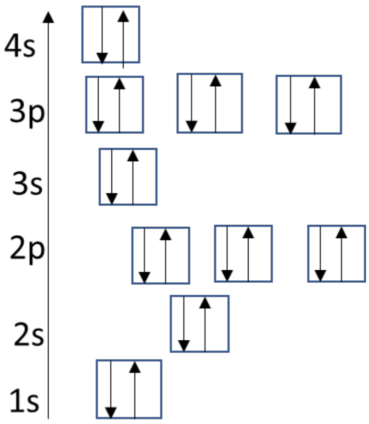
Nitrogen $1s^2\ 2s^2\ 2p^3$

Ca^{2+} $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$





4. The electron configuration of potassium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$. Using energy level diagram, explain why the 4s orbitals fills before 3d.



The 3s subshell is filled after 4s because 4s has a lower energy. As the 3d sublevel becomes filled with electron, the relative energies of the 4s and 3d vary relative to one another and the 4s results in higher energy as the 3d sublevel fills.

5. Describe at least five properties of metals. You must state the property and describe what that property is. Compare and contrast to some properties of non-metals.

A metal is a material that, when freshly prepared, polished, or fractured, shows a lustrous appearance, and conducts electricity and heat relatively well.

- Malleable (can be hammered)
- Ductile (can be drawn into wires)
- Usually solid at room temperature (an exception)
- High density (heavy for their size)
- High melting point
- Good conductors of heat and electricity.

Comparison between metals and non-metals

Metals	Non- metals
Shiny	Dull
High melting points	Low melting points
Good conductors of electricity	Poor conductors of electricity
High density	Low density
Malleable and ductile	Brittle

6. Conduction of electricity is a property of all metals. Explain how the bonding in metals allows them to have this property.

Metals conduct electricity because they have delocalized electrons. These delocalized electrons carry electrical charge through the metals. Metals form giant structure in which electrons move freely in the outer shell of the metal's atom. That is why metals have high melting and high boiling points and conduct electricity. The metallic bond is the force of attraction between these free electrons and positively charged metals ions.

7. Metals react with oxygen in the air when burned:

a. Write a word equation for the reaction between copper and oxygen.

b. Write a symbol equation using chemical formulae to show the reaction between zinc and oxygen, include state symbols:

a. Copper + oxygen → Copper oxide

b. Zn (s) + O_2 (g) → 2 ZnO (s)

8. Metals can also reacts with acids.

a. Write a word equation for the reaction between magnesium and nitric acid:

b. Write a symbol equation for the reaction between sodium hydroxide and hydrochloric acid, include state symbols:

c. Below is the chemical reaction between magnesium and sulphuric acid. One of the products is magnesium sulphate. Explain how many atoms and elements there are in magnesium sulphate and explain how you can tell this:

Magnesium + sulphuric acid → magnesium sulphate + hydrogen

$Mg_{(s)} + H_2SO_{4(aq)} \rightarrow MgSO_{4(aq)} + H_{2(g)}$

a. Magnesium+ nitric acid → magnesium nitrate + hydrogen

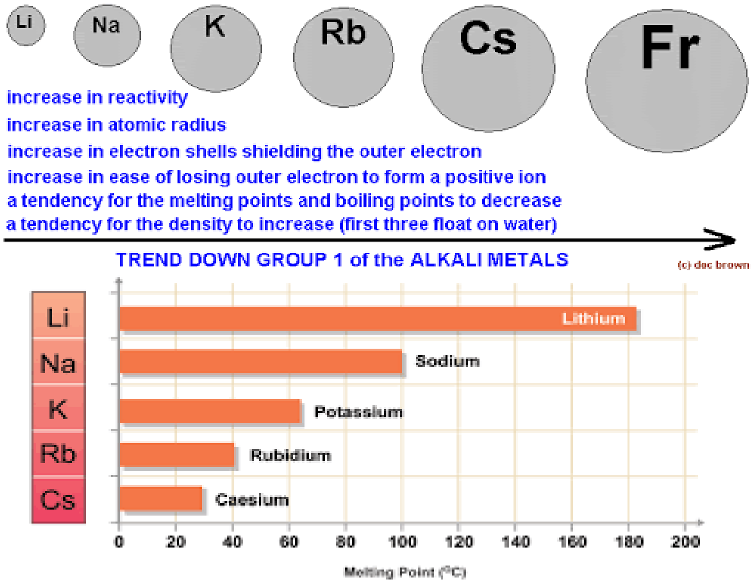
b. NaOH (aq) + HCl (aq) → NaCl (aq) + H_2O (l)

c. Magnesium sulphate has the chemical formulae $MgSO_4$. There are three different elements present, as represented my Mg, S and O. There is also a total of 6 atoms – 1 x Mg, 1 x S and 4 x O, as represented the subscript number after each capital letter that represents an element (O_4).

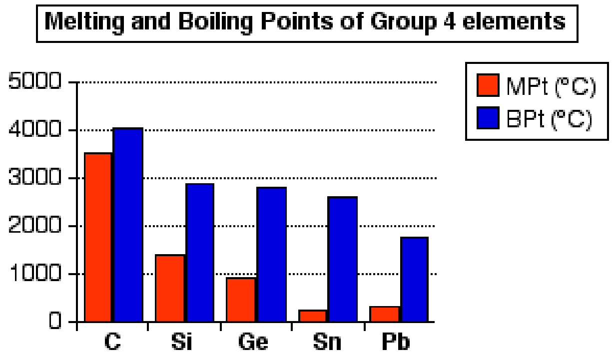
9. Using the melting and boiling points of the elements in Groups 1, 4 and 7, explain how the type of bonding in these elements affects these properties. Use diagrams and graphs in your explanation, and reference using the Harvard system (500 words maximum).

Group one elements bind by metallic bounding, as they are alkali metals. When a positive nucleus attracts an electron from a neighbouring atom in the same element, they bond metallically. This type of bonding only occurs between metals.

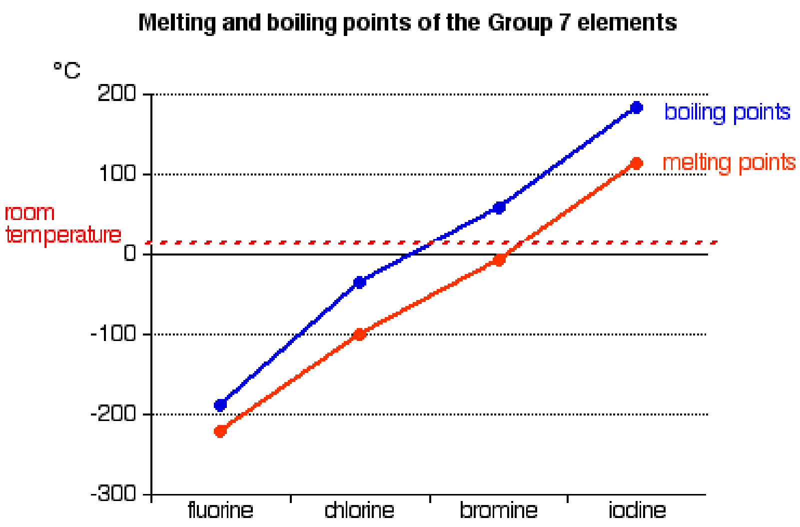
The melting and boiling points of group one elements decrease as you travel down the group. This is because each individual atom increases in size. The increase in atomic radius means that the electrostatic attractive force is weakened. From Lithium, to sodium and potassium etc, the atoms increase in atomic number and size. The bigger they get, the further the delocalised electron from nucleus. This results in weaker force that is easier to break by heat energy.



Group seven elements are halogens that bond covalently. Their melting and boiling point increase as you travel down the group. The intermolecular forces present increase in size when two atoms sharing the electron pair becomes bigger in size because of their increase in atomic mass. Van der Waal's are stronger and more present as you go down group. Fluorine has the weakest van der Waal's, where elements such as Bromine have them stronger. This means that Bromine's dispersion forces are harder to break with heat energy than Florine, resulting in higher melting and boiling point.



Group four elements are alloy metals- metalloids that show characteristics of metals and non-metal simultaneously. The bounding is a mixture of and lies between metallic and covalent bonding. As a result, the melting and boiling point down group 4 can not properly be determined.



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9. **(2.1a) look at the graph below. Describe and explain the trend in first ionisation energy down group 2.**

Ionisation energy down 2 decreases. The number of protons and electron shells increases, but because of three reasons the first ionisation energy does not increase:

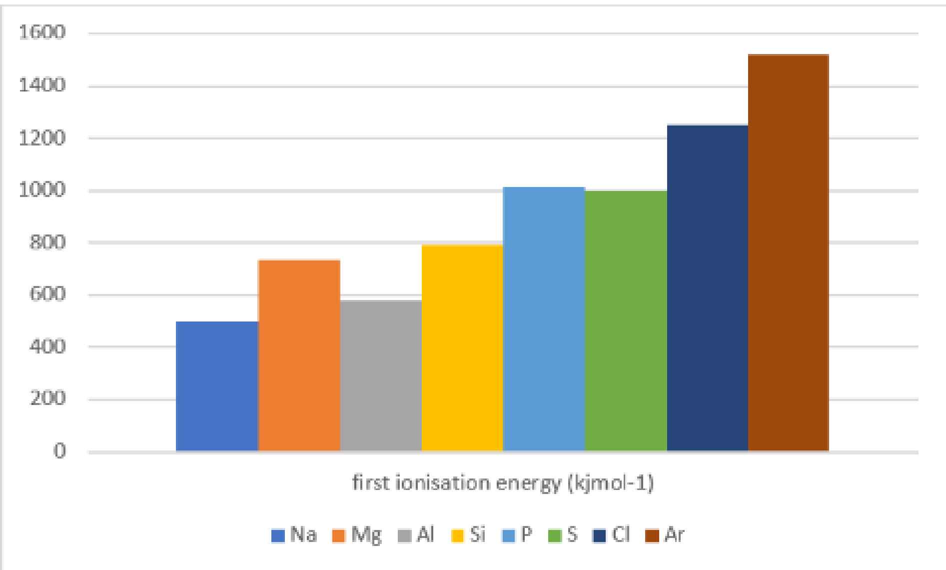
Atomic radius- distance between nucleus and outermost electron increases, weakening its attraction to valence electron. The atom can let go of the electron easier.

Electron shielding – extra electrons added as you travel down the group means that the force of attraction between nucleus and valence electron is weakened.

10. **Ionization energy generally increases across period 3 because the nuclear charge increases but the shielding of the outer electrons remains relatively the same. This means that the electrostatic force of attraction between the outer electrons and the nucleus is becoming greater, so more energy is needed to remove the electrons. Since the elements are in period 3, this means that all of them have an outer electron in energy level 3. However, the trend has two anomalies.**

The first is between Mg and Al, because the outer electron of Mg is in the orbital 3s, whereas that of Al is in 3p. The 3p electron has more energy than the 3s electron, so the ionization energy of Al is less than that of Mg. This makes sense because the 3p electron requires less energy to be removed from the atom. Remember that if an electron has more energy, it needs less external energy to be removed from the atom because having more energy makes it more unstable.

The second anomaly is between p and s. Although both have outer electrons in the 3p orbital, that of p is unpair ($3p^{3)}$ but that of S is paired ($3p^4$). 3p orbitals can fit up to 6 electrons, but if there are 3 or less electrons, they are unpaired and take up a "space" just for themselves. Since the fourth electron in $3p^4$ is paired, it will experience a repulsive force from the third electron. This makes it easier to remove, and therefore the ionization energy of S is less than that of P.



11. Explain the pattern in atomic radii across periods and down groups.

The atomic radius of a chemical element is the distance from the centre of the nucleus to the outermost shell of the electron.

Atomic radius trends in periodic table

- Atomic radius increases moving from top to down across a group. This is because down the group atomic number increases with increase in atomic radius, a new shell is added that means the distance between nucleus and valence shell increases down the group.
- Atomic radius decreases moving left to right across a period. This is because decrease in effective nuclear charge with increase in atomic number. Moving from one element to the next, electrons are being added to the same energy level and protons are being added to the nucleus. One proton has greater effect than one electron, thus, the electrons in the outer energy level are attracted more strongly, and so are pulled closer to the nucleus. This gives in the atomic radius decreasing.

12. Explain why atomic radii of positive ions (cations) are smaller than those of their parent atoms, while atomic radii of negative ions (anions) are larger than those of their parent atoms.

Cations are neutral atoms that have lost an electron and shows a positive charge. They are smaller than their respective atoms, because when an electron is lost, electron-electron repulsion, (result shielding) decreases and the protons are better able to pull the remaining electron towards the nucleus. A second lost electron further reduces the radius of the ion.

Anion are neutral atoms that have gained an electron. They are much larger than their parent atoms. When an added electron occupies an outer orbital, there is increased electron- electron repulsion (result increased shielding) which pushes the electron further apart. Because the electron now outnumbers the protons in the ion, the protons cannot pull the extra electrons as tightly towards the nucleus, and this results in decreased.

13. Define the term oxidation state.

Another name of oxidation state is oxidation number, the total number of electron that an atom either gains or losses in order to form a chemical bond with another atom.
Some rules for finding oxidation state:

- Oxidation state of simple monatomic ion is the same as its charge.
- The oxidation state of free elements is always 0.
- The usual oxidation number of hydrogen is +1, except in metal hydrides for instance, NaH where it is -1.
- The oxidation number of oxygen in compound is usually -2, except in peroxides, where is -1.
- Fluorine is -1.
- The oxidation number of any group one metal is +1, and any group 2 metal is +2.
- The sum of all oxidation states for neutral compound is 0.

14. Oxidation States of the following:

- in H_2O: -2
- O_2 in H_2O_2: -1 per Oxygen molecule
- Cl in HCl: -1
- Cl in $HClO_4$: +7

15. discuss the reaction of the following period 3 elements with oxygen and chlorine.

  a. Sodium
  b. Aluminum
  c. Phosphorous

a. sodium

reaction with oxygen:

Sodium burns in oxygen with an orange flame to produce a white solid mixture of Sodium oxide and sodium peroxide.

For the simple oxide:

$4Na + o_2 \rightarrow Na_2O$

For the peroxide:

$2Na + O_2 \rightarrow Na_2O_2$

Reaction with chlorine

Sodium

Sodium burns in chlorine with an orange flame. White solid Sodium Chloride is produced.

$2Na + Cl_2 \rightarrow 2\ NaCl$

b. Aluminium

Aluminium with oxygen

Aluminium burns in oxygen if it is powdered, otherwise the strong oxide layer on the aluminium incline to inhibit the reaction. If you sprinkle aluminium powder into a Bunsen flame, you get white sparkles. White aluminium oxide is formed.

$4Al + 3O_2 \rightarrow 2Al_2O_3$

Aluminium with Chlorine

Aluminium is often reacted with Chlorine by passing dry Chlorine over aluminium foil heated in a long tube. The aluminium burns in the stream of Chlorine to produce very pale-yellow Aluminium chloride. This results straight from solid to vapour and back again and collects further down the tube where it is cooler.

$2Al + 3Cl_2 \rightarrow 2AlCl_3$

c. Phosphorous

Phosphorus with Oxygen

White phosphorus catches fire spontaneously in air, burning with a white flame and producing clouds of white smoke- a mixture of phosphorus (III) oxide and phosphorus (v) oxide. The proportion of these depend on the amount of oxygen available. In an excess of oxygen, the product will be almost entirely phosphorus (v) oxide. For the phosphorous (III) oxide:

$P_4 + 3O_2 \rightarrow P_4O_6$

For the phosphorous (v) oxide.

$P4 + 5O_2 \rightarrow P_4O_2$

Phosphorus with Chlorine

White phosphorous burns spontaneously in Chlorine to produce a mixture of two Chlorine, phosphorous (III) Chloride and phosphorus (v) Chloride (phosphorus trichloride and phosphorus pentachloride).

Phosphorus (III) chloride is a colourless fuming liquid.

$P_4 + 6Cl_2 \rightarrow P_4Cl_3$

Phosphorus (v) Chloride is an off- white (turns to yellow) solid.

$P_4 + 10Cl_2 \rightarrow P_4Cl_5$

d. Sulphur

Reaction with Oxygen

Sulphur burns in air or oxygen on gentle heating with a pale blue flame. It produces colourless sulphur dioxide gas.

$S + O_2 \rightarrow SO_2$

Reaction with Chlorine

If a stream of Chlorine is passed over some heated sulphur, it reacts to form an orange, evil-smelling, disulphur dichloride, S_2Cl_2.

16. The reactivity, structure and bonding in the chloride or oxide formed from question 15.

The oxides above are known as highest oxides of the different elements. They have highest oxidation states. All the outer electrons in the period 3 elements are being involved in the bonding- from just the one with Sodium, to all seven of Chlorine's outer electrons.

Except Chlorine and Argon, all elements in period 3 combine directly with Oxygen to form oxides. MgO is an ionic oxide. This reaction is full of energy with a sparkling white flame forming a white ash of Magnesium oxide. Al_2O_3 is mostly ionic, but there is important covalent character.

The trend in structure of these oxides is from the metallic oxides, containing giant structures of ions on the left of the period via a giant covalent oxide (silicon dioxide) in the middle to molecular oxides on the right.

References

1. Chemistry LibreTexts. (2020). *Group 1: Properties of Alkali Metals*. [online] Available at: https://chem.libretexts.org/Bookshelves/Inorganic_Chemistry/Modules_and_Websites_(Inorganic_Chemistry)/Descriptive_Chemistry/Elements_Organized_by_Group/Group_01%3A_Hydrogen_and_the_Alkali_Metals/1Group_1%3A_Properties_of_Alkali_Metals [Accessed 14 Feb. 2020].
2. Google.co.uk. (2020). [online] Available at: https://www.google.co.uk/search?q=melting+and+boiling+point+in+group+1&safe=strict&source=lnms&tbm=isch&sa=X&ved=2ahUKEwjAxdWFrdjnAhX9SBUIHd3_A90Q_AUoAXoECBEQAw&biw=1536&bih=722#imgrc=158_6poU6peZ9M [Accessed 17 Feb. 2020].
3. Anon, (2020). [online] Available at: https://chem.libretexts.org/Bookshelves/Inorganic_Chemistry/Modules_and_Websites_(Inorganic_Chemistry)/Descriptive_Chemistry/Elements_Organized_by_Block/2_p-Block_Elements/Group_17%3A_The_Halogens/0Group_17 [Accessed 17 Feb. 2020].
4. Mytutor.co.uk. (2020). *What is the trend in ionization energy across period 3 in the periodic table? | MyTutor*. [online] Available at: https://www.mytutor.co.uk/answers/12676/A-Level/Chemistry/What-is-the-trend-in-ionization-energy-across-period-3-in-the-periodic-table/ [Accessed 17 Feb. 2020].
5. Chemguide.co.uk. (2020). *chemical reactions of period 3 elements*. [online] Available at: https://www.chemguide.co.uk/inorganic/period3/elementsreact.html [Accessed 21 Feb. 2020].
6. En.wikibooks.org. (2020). *A-level Chemistry/AQA/Module 5/Periodicity - Wikibooks, open books for an open world*. [online] Available at: https://en.wikibooks.org/wiki/A-level_Chemistry/AQA/Module_5/Periodicity [Accessed 28 Feb. 2020].
7. ThoughtCo. (2020). *Know the Rules for Assigning Oxidation Numbers*. [online] Available at: https://www.thoughtco.com/rules-for-assigning-oxidation-numbers-607567 [Accessed 28 Feb. 2020].