

P525/1

S.6 Chemistry weekly test 1

Name Marking guide prepared by Do chemical Signature.....

KORO SECONDARY SCHOOL

GROUP TWO ELEMENTS AND THERMOCHEMISTRY

TIME: 2:45 minutes

INSTRUCTIONS:

- Attempt all questions in section A and B

SECTION A (GROUP (II) ELEMENTS)

1. a) Explain what is meant by the term diagonal relationship

(01 mark)

is the similarity in chemical properties between elements in period two to their diagonal neighbours in period 3 and adjacent groups

- (b) State three reasons why beryllium and aluminum exhibit diagonal relationship. (1½mark)

Similar electronegativity ; Similar polarising power Any three
Similar electropositivity ; Similar charge density 0½ mark
Similar electrode potential 1½

- (c) Using equations explain four properties to show diagonal relationship exhibited by beryllium and aluminium. (04 marks)

1- Both oxides are amphoteric and react with conc. alkalis. X
 $\text{BeO(s)} + 2\text{OH(aq)} + \text{H}_2\text{O}_4 \rightarrow \text{Be(OH)}_4^{\text{2-}} \text{(aq)} \text{ X}$

$\text{Al}_2\text{O}_3(\text{s}) + 2\text{OH(aq)} + 3\text{H}_2\text{O}_4 \rightarrow 2\text{Al(OH)}_4^{\text{2-}} \text{(aq)} \text{ X}$ Any 4 1½

2- Both oxides are amphoteric and react with acids X
 $\text{BeO(s)} + 2\text{HCl(aq)} \rightarrow \text{BeCl}_2 + \text{H}_2\text{O}_4 \text{ X}$

$\text{Al}_2\text{O}_3(\text{s}) + 6\text{HCl(aq)} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2\text{O}_4 \text{ X}$ Any 4 1½

3. Both hydroxides are amphoteric and react with alkalis. X
 $\text{Be(OH)}_2(\text{aq}) + 2\text{OH(aq)} \rightarrow \text{Be(OH)}_4^{\text{2-}} \text{(aq)} \text{ X}$

$\text{Al(OH)}_3(\text{aq}) + \text{OH(aq)} \rightarrow \text{Al(OH)}_4^{\text{2-}} \text{(aq)} \text{ X}$ 06

4. Both react with concentrated sodium hydroxide solution X
 $\text{Be(s)} + 2\text{OH(aq)} + 2\text{H}_2\text{O}_4 \rightarrow \text{Be(OH)}_4^{\text{2-}} \text{(aq)} + \text{H}_2\text{g} \text{ X}$

$2\text{Al(s)} + 2\text{OH(aq)} + 6\text{H}_2\text{O}_4 \rightarrow 2\text{Al(OH)}_4^{\text{2-}} \text{(aq)} + 3\text{H}_2\text{g} \text{ X}$

5. Carbides of both yield methane on hydrolysis X
 $\text{Be}_2\text{C(s)} + 4\text{H}_2\text{O}_4 \rightarrow 2\text{Be(OH)}_2(\text{s}) + \text{CH}_4(\text{g}) \text{ X}$

$\text{Al}_2\text{C}_3(\text{s}) + 12\text{H}_2\text{O}_4 \rightarrow 4\text{Al(OH)}_3(\text{s}) + \text{CH}_4(\text{g}) \text{ X}$

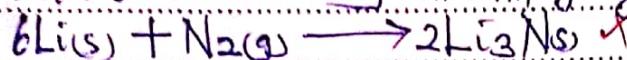
6. Their chlorides are covalent and dimerises in vapour phase. X
 $2\text{AlCl}_3(\text{s}) \rightleftharpoons \text{Al}_2\text{Cl}_6(\text{g}) \text{ X}$

$2\text{BeCl}_2(\text{s}) \rightleftharpoons \text{Be}_2\text{Cl}_4(\text{g}) \text{ X}$

TT = 8½

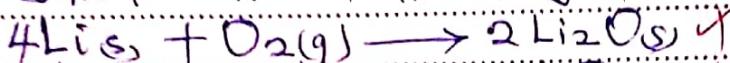
(d) Using equations explain four properties to show diagonal relationship exhibited by lithium and magnesium. (06 marks)

1 - Both combine directly with nitrogen gas forming nitrides \times

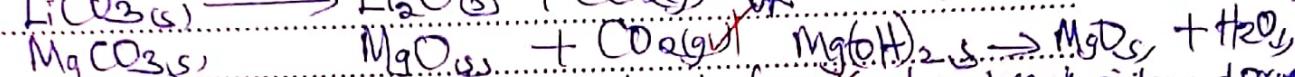
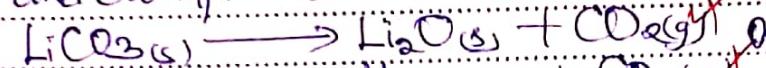


06

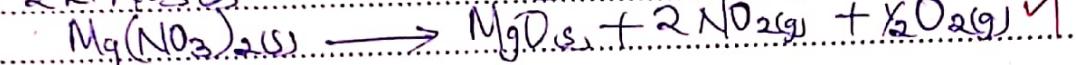
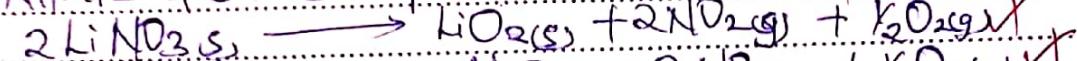
2 - Both form normal oxides only \times



3 - Carbonates, peroxides and hydroxides are sparingly soluble in water and decompose on heating unlike others, which do not decompose on heating

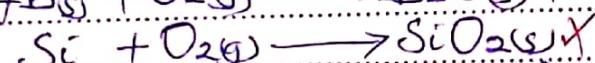
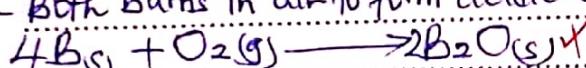


4 - Nitrates of both decompose on heating to form oxides, nitrogen dioxide and oxygen



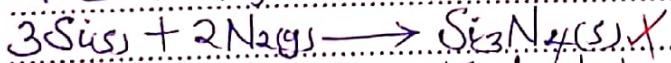
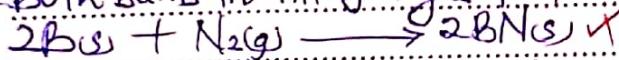
(e) Using equations explain four properties to show diagonal relationship exhibited by boron and silicon. (06 marks)

1 - Both burns in air to form acidic oxides \times

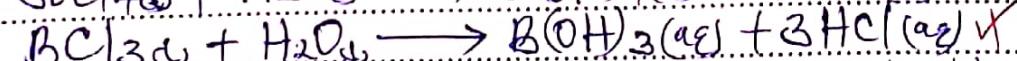
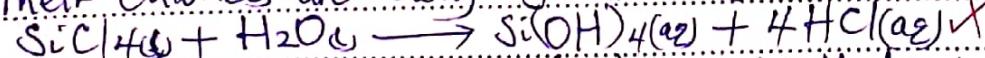


06

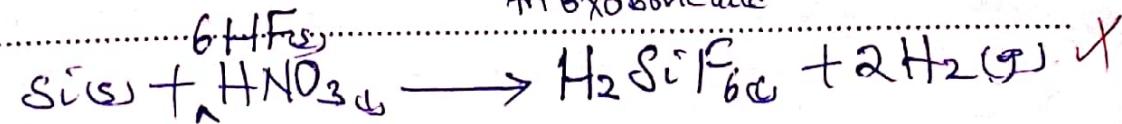
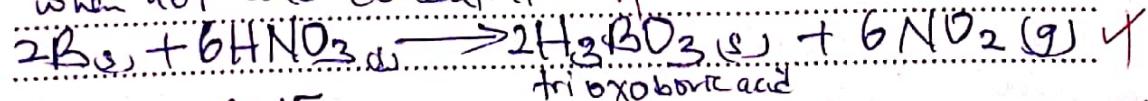
2 - Both burns in nitrogen gas to form nitrides \times



3 - Their chlorides are readily hydrolysed by water to form hydroxides \times



4 - Both react with only oxidising acids i.e. nitric and sulphuric acid, only when hot and concentrated \times



TT = 12

2. The first ionization energy of some elements are as shown below.

Elements	1 st ionization energy	2 nd ionization energy	3 rd ionization energy	4 th ionization energy
A	500	4.2	4600	2
B	740	20	1500	5.1
C	630	2.5	1600	1.8
D	900	2	1800	8.2
E	580	3.1	1800	1.5
				9500
				10500
				4800
				21000
				11600

a) What is meant by the term ionization energy? (01 marks)

IS the minimum amount of energy required to remove an electron from a gaseous atom to form a gaseous cation 01

OR is the minimum amount of energy required to remove an electron completely from a gaseous atom to form a unpositively charged gaseous ion.

b) (i) State the elements that are most likely to form an ion with a unit positive charge. Give reasons for your answer. (02 marks)

Element A, C ✓

They belong to group I, and has one electron in the outermost energy level because there was a big rise in IE while removing the second electron implying that it is being removed from another energy level closer to the nucleus.

(ii). Identify two elements that are in the same group of the periodic table. (01 marks)

A and C or D and E

02

01

3. (a) Compare the following properties of group (II) and group (I) elements. In each case, give a reason for your answer.

(i) First ionization energy ^{1 mark} group II have high first ionization energy than group I, Reason: in group II elements nuclear charge outweighs the shielding thus their electron are strongly held by the attraction of nucleus 02

(ii) Melting point ^{1 mark}.

group II elements have high melting points than group I elements

Reason: group II elements have large charge density which give it high lattice energy in molecules, hence their molecules are thermodynamically stable 02

TT=08

(iii) Electro positivity [02 mark]

Electro positivity is higher in group I than group II ✓

Reason: In group I elements, the long atomic radius, due to smaller nuclear charge results into less attraction for the outermost electrons thus the tendency to lose electrons and become positively charged is high. 02

(b) The decomposition temperatures of the carbonates of group (II) elements are given below

Carbonate	$MgCO_3$	$CaCO_3$	$SrCO_3$	$BaCO_3$
Decomposition temperature °C	404	826	1098	1370

(i) State how the decomposition temperatures vary

(01 mark)

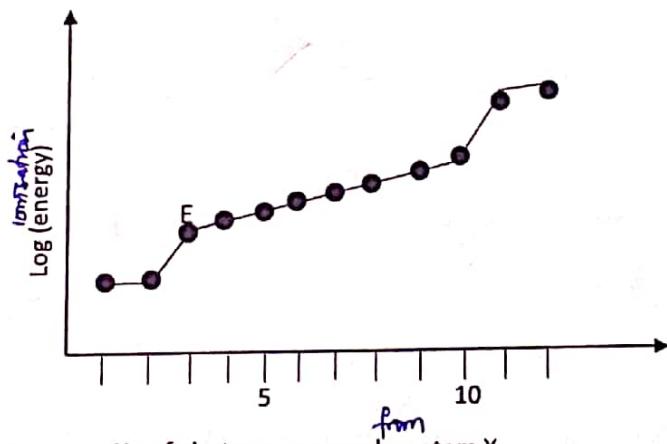
Decomposition temperatures increases down the group. ✓ 01

(ii) Explain your answer in (b) (i)

(3 1/2 marks)

Moving down the group, the nuclear charge decreases, the ionic radius increases, the charge on the carbonates is the same throughout. The charge density/polarising power of the cation decreases, the bonding in the metal carbonates become more ionic and less covalent, hence the compounds become more stable down the group. 3 1/2

4. The diagram below shows successive ionisation energies for an element X, showing removal of all electrons.



(a) Giving reasons state

(i) The group of element X

(1 1/2 marks)

group II ✓

Has two electrons in the outermost energy level.

OR: Removing third electron requires more energy than first and second electron

(1 1/2 marks)

(ii) The period of element X

Period 3. ✓

Has three energy levels/principle Quantum shells.

1 1/2

1 1/2

$$T = 09 \frac{1}{2}$$

(iii) Identify element X
 Magnesium ✓ (01mark) 01

(b) Explain the sudden increase in the energy required to remove electron E
 Electron E is closer to the nucleus and its held strongly by the nuclear attraction (force). (2marks) 02

(c) Explain how the size of X will change as electrons are removed
 Size will decrease, because the remaining electrons will be held more strongly by the attraction of the nucleus hence pull nearer (01mark) 01

(d) (i) explain what would be the sign of change if an electron was added to X to give X⁻
 The first electron affinity is exothermic i.e negative because an electron is being added to a neutral atom (01mark) 01

(ii) How would you expect it to affect the size if X gains an electron.
 Size increase; the outermost electrons will experience shielding effect and would be less attracted toward nucleus hence increase in size. (01mark) 01

(e) Explain giving reasons whether you would expect X to form compound in the +1 oxidation state
 X would not form compound in the +1 oxidation state (01mark) 01

Reason: +1 is less stable, so it loses another electron to gain +2 oxidation state which is more stable

5. (a) Explain what is meant by the term electronegativity. (01mark)

Tendency of an atom to gain electrons to become negatively charged 01

(b) State the factors that affect ~~determine~~ the value of electronegativity of an element. (2marks)
 - Atomic radius; ✓
 - Nuclear charge; ✓ 02

(c) Explain how the following factors affect the value of electronegativity of the element

(i) Atomic radius
 The shorter the atomic radius the higher the value of electronegativity (2marks) 02

(ii) Nuclear charge (2marks)

The greater the nuclear charge the greater the value of electronegativity 02

$$\overline{TT} = 14$$

(iii) The screening effect of the inner electrons

(02marks)

The greater the screening effect the lower the value of electronegativity.

02(d) Explain the difference between *electronegativity* and *electron affinity*

(02marks)

Electronegativity is the tendency of an atom to gain electrons and become negatively charged, while electron affinity is the amount of heat change when electron is gained to form negatively charged ion.

02

6. (a) Explain what is meant by the term first electron affinity.

(01mark)

Is the heat ^{released} when ^{an} electron is gained by one mole by 1 mole of gaseous atom or ion to form ^{uni-} negatively charged ion.

(b) State three factors that can affect electron affinity.

(01 1/2marks)

~~Factors~~
 - Nuclear ~~charge~~ is the minimum heat given out when one mole an electron is gained by one mole of an isolated neutral gaseous atom to form $1\frac{1}{2}$ uninegative ion.

1 1/2

(c) The first electron affinities of some elements of period - 3 are given in the table below

Element	Al	Si	P	S
First electron affinity (KJmol ⁻¹)	-44	-134	-71.7	-200

(i) State the trend in variation of electron affinities

(0 1/2 mark)

Electron affinities increases across the period.

1 1/2

(ii) Explain your answer in c (i) above

(02 marks)

Across the period, there is increase in effective nuclear charge and increase in screening effect, however the increase in nuclear charge outweighs the shielding effect, this leads to increase in the attraction for the incoming electrons hence increase in first electron affinity.

02

$$\overline{TT} = \overline{8\frac{1}{2} + \frac{1}{2}}$$

7. (a) Explain the following observations

- i) Group (II) metal carbonates decompose on heating but group (I) metal carbonates are resistant to decomposition by heat. [3 marks]

group (I) ions has a high charge, small ionic radius than group II ions. It therefore has high polarising power which makes it easy to polarise the carbonate ion making the bonding in group II carbonates to have some covalent character while the bonding in group I carbonate is ionic due to the low polarising power of +1, hence group II carbonates easily decompose. 03

- ii) Group (II) metal sulphates are more soluble in water than group (I) metal sulphates. (03 marks)
 Group II ions has high nuclear charge and high charge density/high polarising power hence they have a higher ability to attract lone pair of electrons in water; afterwards their hydration energy outweighs the lattice energy; Conversely group I metal sulphates are less soluble due to lower nuclear charge; low polarising power/charge density these make the sulphates to have high lattice energy than the hydration energy; Also there is lower ability to attract lone pair of electrons from water molecule due to small charge density hence less soluble

- iii) The decomposition temperature of group (II) metal carbonates increases down the group. (03 marks)

The decomposition temperature of group II metal carbonates increases down the group because down the group, the charge remains the same but the ionic radius increases, the charge on the carbonate ion is also the same throughout. The polarising power of the cations decreases, the bonding in the metal carbonates becomes more ionic down the group, hence the compounds becomes more stable. 03

- iv) Lithium compounds are mainly covalent while the compounds of other group (I) elements are mainly ionic. (03 marks)

Lithium ions has short ionic radius and high polarising power/charge density these makes the bonding electrons tightly pulled towards it hence the bond becomes more covalent, the nature of the bond in other group I compounds is ionic due to the low charge density/polarising power. 03

$\text{FT} = \underline{\underline{12}}$

- v) The solubility of group (II) metal hydroxides increases down the group. (03 marks)

Due to decrease in lattice energy as the size of ionic radius of metal ion increase. Large ionic radius makes less holding force of the ion reducing the lattice energy. 03

- vi) The solubility of sulphates of group (II) metals decreases down the group. (03 marks)

Because both lattice energy and hydration energy decreases but the hydration energy decreases more rapidly than the lattice energy. Since both cation and anions are double charged, this results in high lattice energy hence decrease in solubility down the group. 03

SECTION B THERMOCHEMISTRY

8. (a) States the laws of thermochemistry (04 marks)

First law (Lavoisier - Laplace law) states that the amount of heat used to decompose a compound to its element is equal in magnitude but opposite in sign to the heat given out when the compound is being formed. 02

Second law: (Hess's law of constant heat summation) states that if a reaction is carried out in a series of steps, the enthalpy change for the overall reaction equals the sum of the enthalpy changes for the individual steps. 02

- (b) Explain the factors that affect enthalpy of reaction steps at constant temperature and pressure (05 marks)

- The amounts of reactants or products that are involved e.g. the heat of combustion of 2 moles of hydrogen is double that of combustion of 1 mole
- Physical states of reactants or products; example: when hydrogen combine with O_2 to give liquid H_2O $\Delta H_f^\circ = -286 \text{ kJ mol}^{-1}$ while when it combine to form gas H_2O $\Delta H_f^\circ = -242 \text{ kJ mol}^{-1}$
- The temperature at which the reaction is taking place must be specified and usually the standard temperature of ~~760 mm Hg~~ of 25°C or 298K is used
- The pressure at which the reaction is taking place must be specified and usually the standard pressure of ~~760 mm Hg~~ or 1 atm is used.
- Allotopic modifications e.g. the enthalpy of combustion of diamond is -396 kJ mol^{-1} while heat of graphite is -394 kJ mol^{-1} 05

(c) Define standard enthalpy of formation

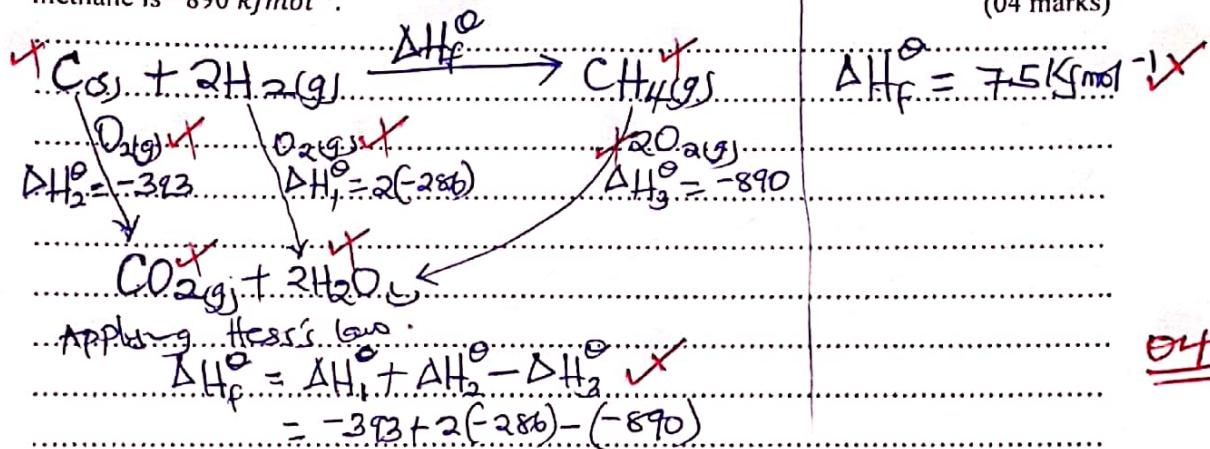
Is the enthalpy change when one mole of a compound is formed from its constituent elements in their normal physical states under standard condition of pressure and temperature

(02 marks)

02

(d) Calculate the enthalpy of formation of methane given that the enthalpy of combustion of carbon is -393 kJ mol^{-1} , that of hydrogen is -286 kJ mol^{-1} and enthalpy of combustion of methane is -890 kJ mol^{-1} .

(04 marks)

04

9. (a) Define standard enthalpy of atomization

(02 marks)

Is the heat required to convert a substance in its normal physical state and under standard conditions into one mole of atoms in gaseous state.

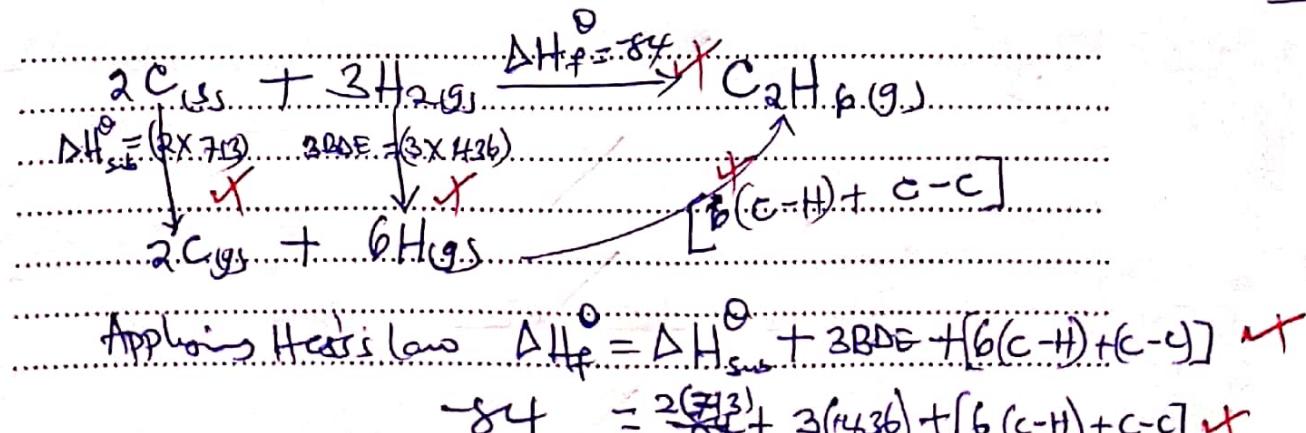
02

(b) Calculate the C-C bond energy in ethane given that;

(04 marks)



, if Average BDE(C-H) = -412



From Example ②, average(C-H) = -412 kJ mol⁻¹

Average bond DE(C-C) = -346 kJ mol⁻¹

04

10. Given that the standard enthalpy of formation of CH_4 is -63 kJ mol^{-1} and that of C_2H_6 is -84 kJ mol^{-1} , calculate the standard enthalpy change for the reaction below
 $C_2H_6 + H_2(g) \rightarrow 2CH_4(g)$ (04 marks)

Standard enthalpy for reaction
 From $\Delta H_r^\circ = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$

$$\Delta H_f = [2(-63)] - [(-84) + 6] \quad \checkmark$$

$$\Delta H_r = -42 \text{ kJ/mol} \checkmark$$

11. Some bond energies, BE are given below; **Bond energy**.

Bond	$C - H$	$C = O$	$O - H$	$O = O$	$C - O$	$C \equiv C$	$C - C$	$H - H$
BE ($kJmol^{-1}$)	412	803	463	496	326	813	346	436

Use the bond energies given above to calculate;

(i) The standard enthalpy of combustion of methanol. (0.5 marks)



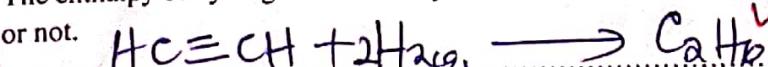
$$\Delta H_r = \sum (\text{BE})_{\text{reactants}} - \sum (\text{BE})_{\text{products}}$$

$$= [3(\text{C-H}) + (\text{C-O}) + (\text{O-H}) + \frac{3}{2}(\text{O=O})] - [2(\text{C=O}) + 2(\text{H-O-H})]$$

$$= [3(412) + (803) + 463 + \frac{3}{2}(496)] - [2(803) + 2(2(463))]$$

$$\begin{array}{r} 2769 - 3458 \\ \hline = 689 \text{ kTm} \end{array} \quad \checkmark$$

(ii) The enthalpy of hydrogenation of ethyne. State whether hydrogenation of ethyne is feasible (02 marks)



\sum BE inputs - \sum BE products

$$2(4_{12}) + 8_{13} + 2(4_{26}) - 1(3_{46}) + 6(4_{12}) = -309 \text{ KJ mol}^{-1}$$

Hence its feasible

12. (a) Define Born-Haber cycle.

define Born Haber cycle (02 marks)
Is the cycle which shows formation of one mole of an ionic compound from its elements in their normal physical state.

$$\pi = \frac{10}{3} + 2$$

(b) Describe the energy changes that take place in a *born* Haber cycle

(04 marks)

1 - Atomization of gas

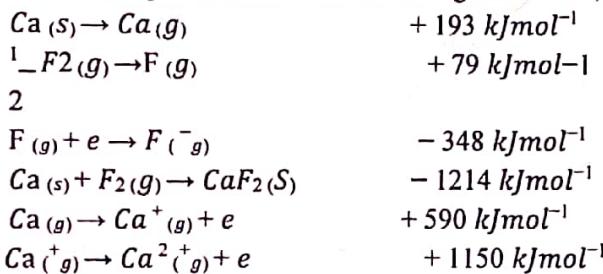
1. Is the energy required to form one mole of gaseous atom ✓

2. Ionisation of a metal: Is the energy required to remove one mole of electrons from gaseous atom to form one mole of positively charged gaseous ion.

3. Electron affinity: Is the energy change when one mole of an electron is added to one mole of a gaseous atom to form one mole of a negatively charged ion.

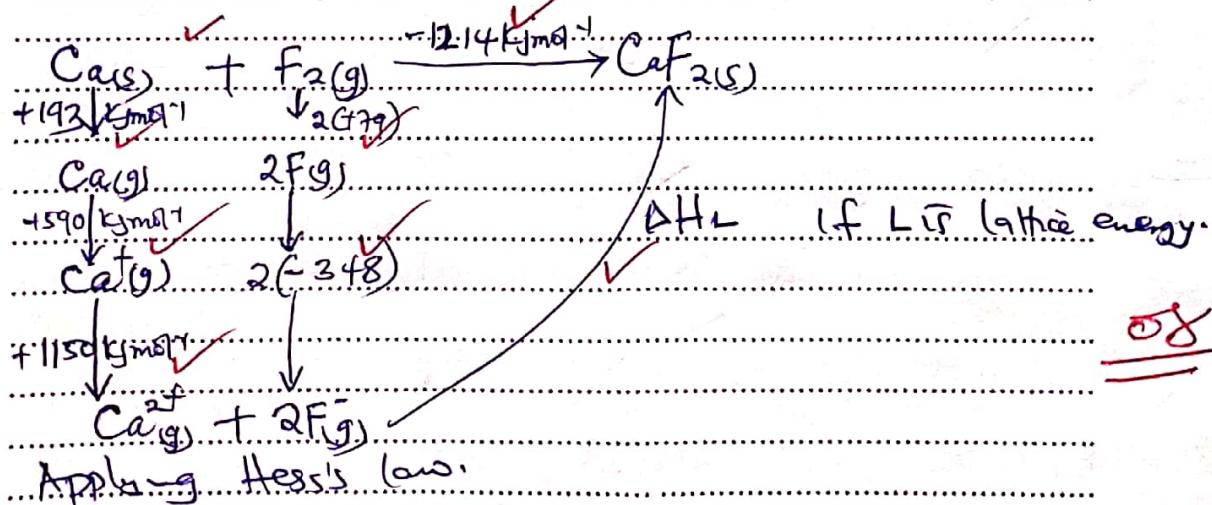
4. Lattice energy: Is enthalpy change that occurs when one mole of a crystal lattice is formed from or broken into its constituent gaseous ions.

(c) Standard enthalpy changes for some reactions are given below;



(i) Construct a Born-Haber cycle for the formation of calcium fluoride.

(8 marks)



$$-1214 = +193 + +590 + 1150 + 2(+77) + 2(-348) + \Delta H_L$$

$$\Delta H_r = -2609 \text{ kJ/mol}$$

$$TT = 12$$

(ii) Use the cycle you have constructed in (i) to calculate the lattice energy of calcium fluoride. (02 mark)

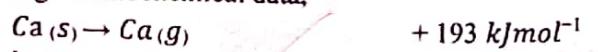
..... Applying Hess's law

$$-1214 = 193 + 590 + 1150 + 2(79) + 2(-348) + \Delta H_L$$

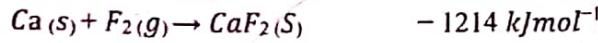
$$\Delta H_L = 2609 \text{ kJmol}^{-1}$$

02

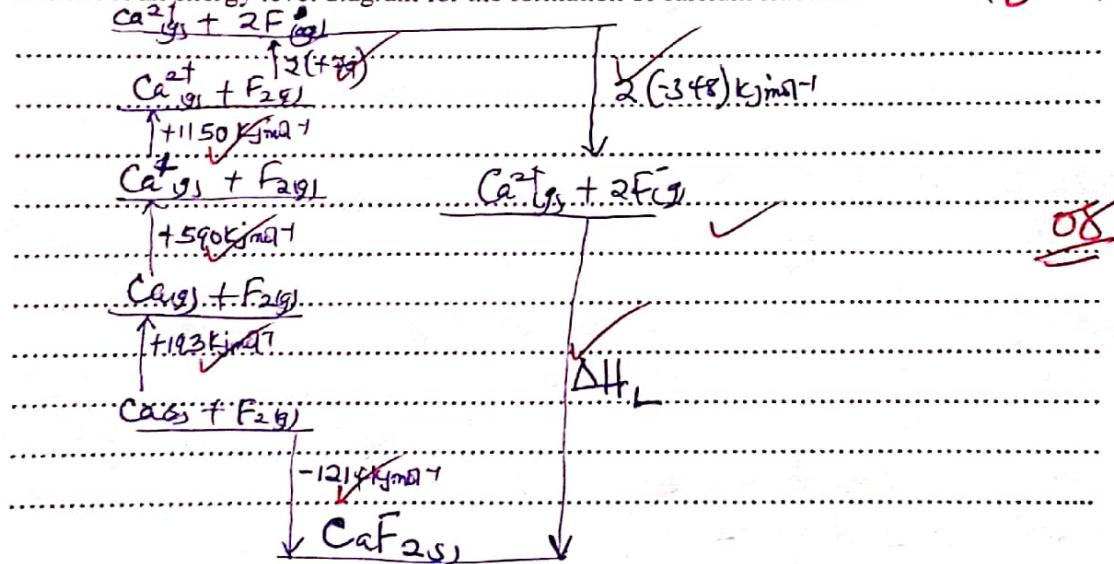
13. Given the following thermochemical data,



2



(a) Construct an energy level diagram for the formation of calcium fluoride. (08 marks)



08

(b) Use the energy level diagram you have constructed in (a) to calculate the lattice energy of calcium fluoride. (02 marks)

..... Applying Hess's law

$$-1214 = 193 + 590 + 1150 + 2(79) + 2(-348) + \Delta H_L$$

$$\Delta H_L = -2609 \text{ kJmol}^{-1}$$

02

TF = 12

14. (a) Define the terms

(i) Lattice energy

Is measure of strength of electrostatic attraction between ~~between~~ 02 marks
the positive and negative ions.

(ii) Solvation energy/~~hydration~~

Is the enthalpy change (heat given out) when one mole of a ~~gas~~ 02 marks
ion is surrounded by water molecules to form an infinite dilute solution

(iii) Enthalpy of solution

Is enthalpy change when one mole of an ionic solute is ~~is~~ 02 marks
dissolved in water to form an infinitely dilute solution so that further
dilution does not give any further heat change

(b) State two factors that can affect hydration energy. (02 marks)

- Charge on the ion ✓

- Ionic radius: ✓

(c) The table below shows the enthalpies of hydration of cations of group (II) elements of the periodic table.

Cation	Mg^{2+}	Ca^{2+}	Sr^{2+}	Ba^{2+}
Enthalpy of hydration ($kJ mol^{-1}$)	-1920	-1640	-1480	-1360

(i) State how hydration energy of the ions vary. (01 marks)

It decreases down the group ✓

01

(ii) Explain your answer in b(i) above. (03 marks)

The ~~decrease in~~ hydration energy ~~outweighs~~ the lattice energy because up the group the ~~ions~~ 03
ions have high nuclear charge and charge density hence they have a high ability to attract lone pairs of electrons in water thus the hydration energy is higher than lattice energy, but down the group there is decrease in charge density making them to have low ability to attract lone pairs of electrons

(iii) Explain why the values of hydration energy are negative. (02 marks)

because it is accompanied by release of energy, ✓

02

TT = 14

(c) The values for some energy changes are given below.

Lattice energy of $\text{CaCl}_2(s)$	-2230 kJmol^{-1}
Enthalpy of hydration of $\text{Cl}^-(g)$	-343 kJmol^{-1}

(i) Calculate the enthalpy of solution of calcium chloride. (02 marks)

$$\text{enthalpy of solution} = \text{lattice energy} + \text{hydration}$$

$$= -2230 + -343$$

$$= +1887 \text{ kJmol}^{-1}$$

02

(ii) State how the solubility of calcium chloride, would be affected when the temperature is increased. Give a reason for your answer. (04 marks)

~~Solubility of calcium chloride would increase because enthalpy of solution is positive or endothermic; the extra heat breaks the lattice/ holding force leading to formation of ions which has high charge density to attract water molecules~~

13. (a) Define enthalpy of neutralisation. (02 marks)

~~Is the enthalpy change of neutralisation change (heat evolved) when one mole of hydrogen ions reacts with one mole of hydroxyl ions to form one mole of water~~

02

(a) Briefly explain why the enthalpy of neutralisation of ethanoic acid by potassium hydroxide is -55.8 kJmol^{-1} while the enthalpy of neutralisation of hydrochloric acid by potassium hydroxide is -57.3 kJmol^{-1} (03 marks)

~~The enthalpy of neutralisation of ethanoic acid by potassium hydroxide is low i.e. -55.8 kJmol^{-1} because ethanoic acid is a weak acid thus some energy is first used in its ionisation while the enthalpy of neutralisation of hydrochloric acid is high i.e. -57.3 kJmol^{-1} because HCl is a strong acid and it is completely ionised to give $\text{H}^+(\text{aq})$~~

03

TT = 11

- (b) Explain why the enthalpy of neutralisation of sodium hydroxide by nitric acid is 57.1 kJ/mol^{-1} whereas the enthalpy of neutralisation of sodium hydroxide by hydrocyanic acid is 12.0 kJ/mol^{-1} . (02 marks)

Hydrocyanic acid is a weak acid thus requires some of the energy for it to ionise to form H^+ hence enthalpy of neutralisation is low i.e. 12.0 kJ/mol^{-1} while enthalpy of neutralisation of NaOH by Nitric acid is high (57.1) because HNO_3 is a strong acid and ionises completely to form H^+ .

- (c) 250cm^3 of 0.5M KOH at 12°C were mixed in a plastic beaker of negligible heat capacity with an equal volume of 0.5M HCl at the same temperature. The final temperature was 15.4°C . Calculate the enthalpy of neutralisation assuming that the specific heat capacity of the solution is $4.2\text{J/g}^{-1}\text{K}^{-1}$. (06 marks)

using, Heat change = (Total mass of solution \times SHC \times Temp change)

$$\text{Total mass of solution} = (250 + 250) = 500\text{g} \quad \text{X}$$

$$\text{SHC} = 4.2\text{J/g}^{-1}\text{K}^{-1}$$

$$\text{Temperature change} = (15.4 - 12)^\circ\text{C} = 3.4^\circ\text{C} \quad \text{X}$$

$$= 500\text{g} \times 4.2\text{J/g}^{-1}\text{K}^{-1} \times 3.4^\circ\text{C}$$

$$= 7140 \quad \text{X}$$

But moles of KOH equals moles of HCl reacted.

These implies that moles of water formed $\frac{250}{1000} \times 0.5 = 0.125 \text{ moles}$

0.125 moles of water are formed after liberation of 7140J

1 mole of water are formed after liberation of $1 \times 7140\text{J}$

$$\frac{0.125 \text{ moles}}{1 \times 7140\text{J}} = -57.12 \text{ kJ/mol}^{-1} \quad \text{X}$$

Therefore enthalpy of neutralisation = $-57.12 \text{ kJ/mol}^{-1}$

06