

present starts to disintegrate the amount of carbon-14 in the dead object after t years is determined as the N_t . The amount of C-14 in a similar live object is taken as N_0 . Using the $t_{1/2}$ of Carbon 14 which is known (5570 years), these values are used to calculate the approximate age of the object using equations $t_{1/2} = \frac{\ln 2}{\lambda}$ and $\ln\left(\frac{N_0}{N_t}\right) = \lambda t$



$$t_{1/2} = 1590$$

t for 40% 4/06/2025

THERMO-CHEMISTRY

log mass

This is the study of energy changes accompanying a chemical process or reaction. It investigates/determines the feasibility of a reaction and how far a reaction

time can go.

λ = decay constant

$-\lambda$ = slope Enthalpy changes & energy profiles

Heat energy leaves a system if the system is at a higher temperature than the surroundings and vice versa.

$$t_{1/2} = \frac{\ln 2}{\lambda} \quad \ln\left(\frac{N_0}{N_t}\right) = \lambda t$$

Enthalpy (H) is the heat content or total energy of the system.



Enthalpy change (ΔH) is the change in the system's heat content. It is denoted as ΔH .

Standard enthalpy change / heat of the reaction $\rightarrow (\Delta H^\ominus)$ - This is the heat change that occurs when moles of the reactants shown in a balanced equation for a given reaction in their standard states react completely & at standard conditions.

The standard / normal / natural state of a substance is the phase in which a pure substance exists at one atmosphere and 298 K / 25°C

Hydroxyl

Sent you points

Types of reactions.

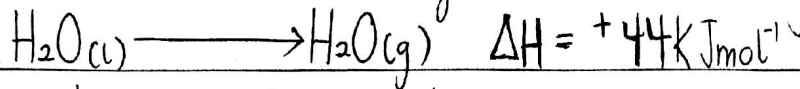
Endo - +ve
Exo - -ve

Chemical reactions are accompanied by some energy changes which involve absorption or evolution/giving off of heat. The heat change involved also referred to as enthalpy change can be positive or negative depending on the heat content of products and reactants. $\Delta H_{\text{reaction}} = \Delta H_{\text{products}} - \Delta H_{\text{reactants}}$ ($\Delta H_r = H_p - H_r$)

There are two types of reactions depending on the heat changes involved:

(i) Endothermic reactions

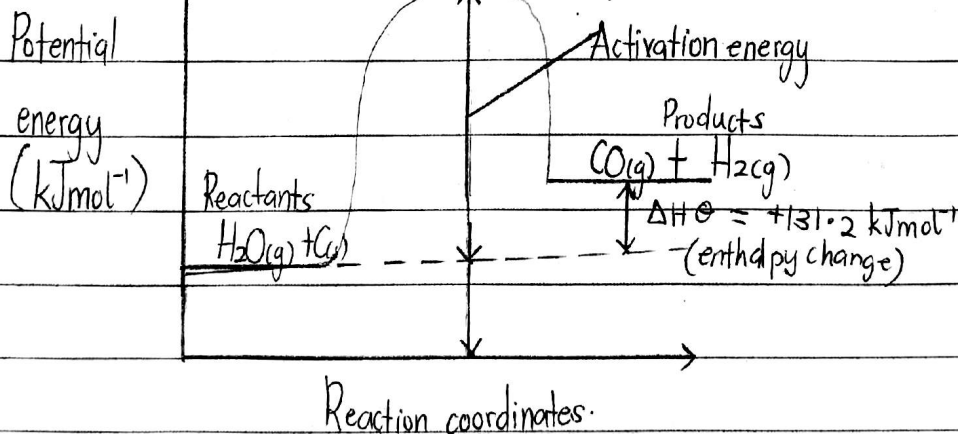
This is a chemical reaction which proceeds with absorption of heat from the surrounding. Temperatures of the surrounding decrease as the heat content is absorbed from the surrounding to the reaction. The heat change is denoted $+\Delta H$ e.g.



Potential diagram/profile diagram for endothermic reactions.

Consider a reaction $\text{C}(\text{s}) + \text{H}_2\text{O}(\text{g}) \longrightarrow \text{CO}(\text{g}) + \text{H}_2(\text{g}) \quad \Delta H^\circ = +131.2 \text{ kJ mol}^{-1}$

(Activated complex) where bonds are broken and formed.

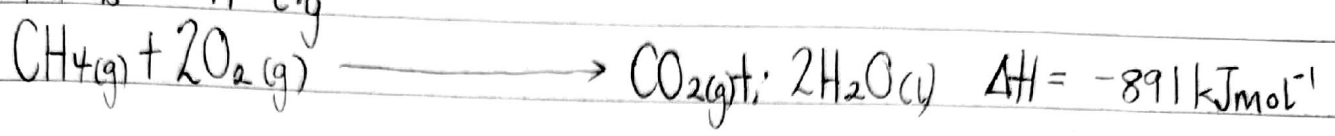


06/2025 Examples of endothermic reactions in daily life.

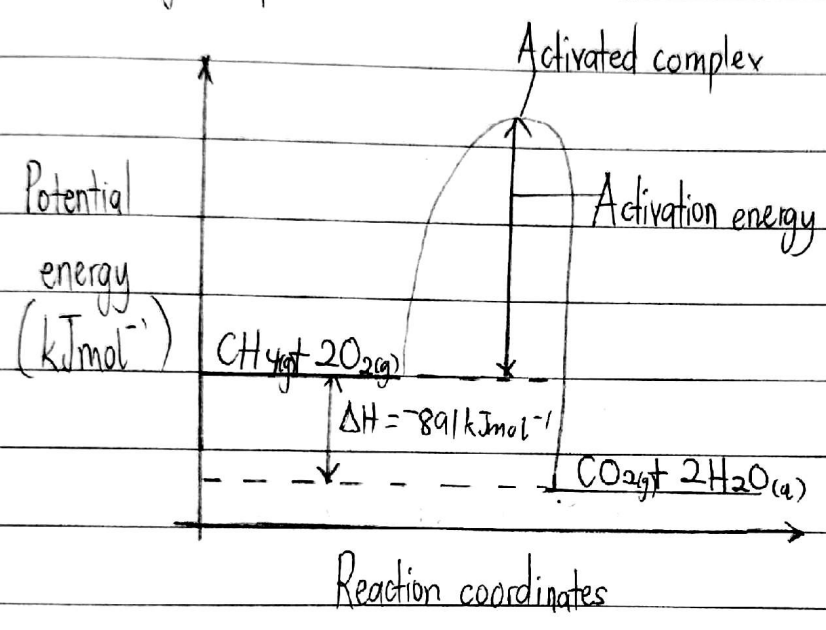
- Evaporation of liquid water
- Photosynthesis (manufacture of food by plants)
- Cold packs (ammonium nitrate dissolving in water)
- Melting of ice
- Tepid sponging
- Cooling of the engine

(ii) Exothermic reactions.

This is a chemical reaction in which heat is released to the surrounding. The system loses heat and the temperature of the surrounding increases. This heat change is denoted as $-\Delta H$ e.g.



Potential diagram for exothermic reactions.



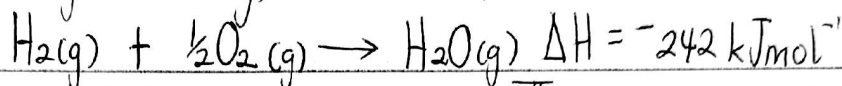
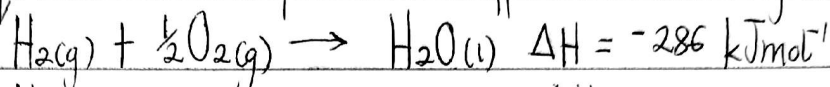
Examples of exothermic reactions in daily life.

- Respiration
- Condensation
- Freezing of water
- Biogas production
- Fermentation
- ~~Exo~~ Rusting
- Combustion
- Neutralisation

Factors that affect enthalpy of a reaction.

Amount of products or reactants that are involved in a reaction. That is why heat changes are quoted for specific molar quantities of either reactants or products. e.g. the heat of combustion of two moles of carbon is double the heat of combustion of one mole.

The physical state of reactants and products e.g. when hydrogen & oxygen gases combine² give liquid water the heat of reaction is different than when they combine to form steam



Temperature at which the reaction is taking place. This must be specified and usually the standard temperature is 25°C / 298K

The pressure at which the reaction is taking place and this is usually specified as one atmosphere.

The allotropic form of the reactant used in the reaction for example the enthalpy of combustion of carbon in form of diamond is -396 kJmol^{-1} while that of graphite is -394 kJmol^{-1} .

Types of enthalpy changes & Hess's law.

Laws of thermochemistry.

There are two laws of thermochemistry i.e.;

(i) Lavoisier-Laplace's law

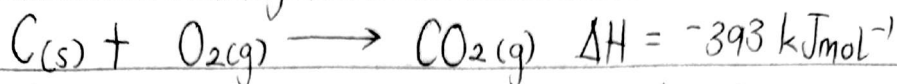
(ii) Hess's law

(i) Lavoisier-Laplace's law

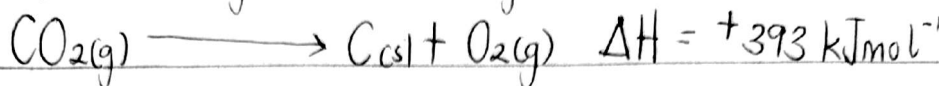
It states that the amount of heat used to decompose a compound into its constituent elements is equal in magnitude but opposite in sign to the heat given off when the compound is being formed. This means that when a chemical reaction is reversed, the sign of the

enthalpy of the reaction is changed i.e.

exo



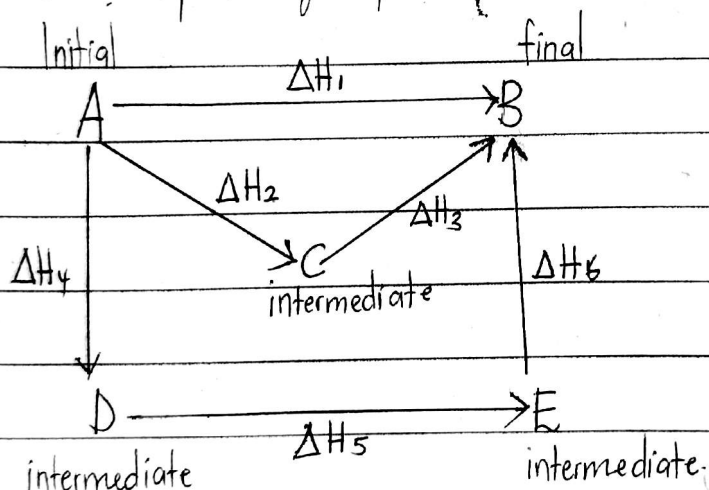
endo



(ii) Hess's law of constant heat summation

It states that if a reaction is carried out in a series of steps, the enthalpy change for the chemical reaction (overall) equals to the sum of the enthalpy changes for the individual steps at constant temperature and pressure. **OR**

At constant temperature and pressure, enthalpy change for a chemical reaction is the same irrespective of the number of steps passed through but will only depend on the initial and final stages of the reaction. Illustration:

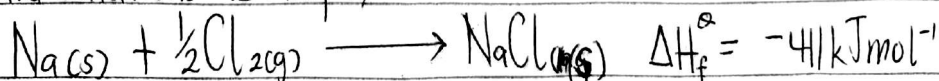


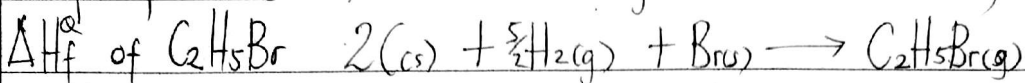
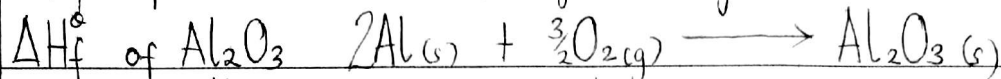
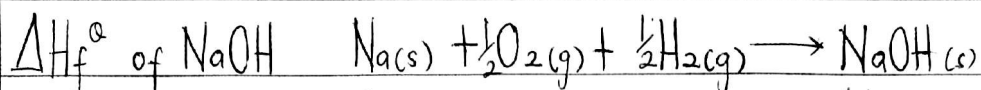
$$\Delta H_1 = \Delta H_2 + \Delta H_3 = \Delta H_4 + \Delta H_5 + \Delta H_6$$

Types of enthalpy changes

1. Standard enthalpy of formation (ΔH_f°) -ve

This is the enthalpy change that occurs when one mole of a compound is formed from its constituent elements in their natural/normal/standard physical states at standard conditions. Example;





All elements in their standard states have a value of zero as their enthalpy of formation.

Enthalpies of formation of compounds can be used to calculate the heat change of any chemical reaction.

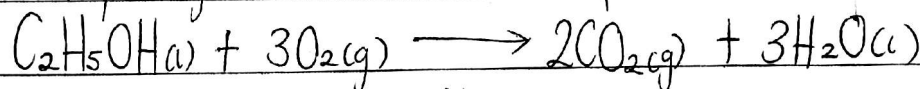
The stability of a chemical substance can be determined using the enthalpy of formation of that substance. A negative value shows that the compound is stable while a positive value shows that the compound is not stable.

Standard heat of a reaction can be determined using the enthalpies of formation of the compounds in the reaction by using the formula below.

$$\Delta H_{\text{reaction}}^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

Sample

1. Standard enthalpy of formation of ethanol, carbon dioxide and water are -277 , -393 and -285 kJmol^{-1} respectively. Calculate the standard heat of the reaction below:



$$\Delta H_{\text{reaction}}^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\sum \Delta H_f^\circ \text{ products} = 2 \times -393 + 3 \times -285$$

$$= -786 + -855$$

$$= -1641 \text{ kJmol}^{-1}$$

$$\sum \Delta H_f^\circ \text{ reactants} = 3 \times 0 + 1 \times -277$$

$$= -277 \text{ kJmol}^{-1}$$

$$\Delta H_{\text{reaction}}^\circ = -1641 - (-277)$$

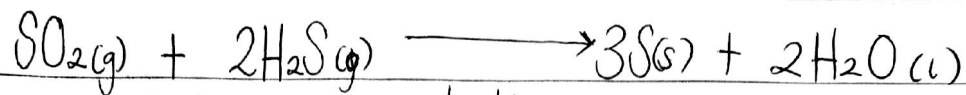
$$= -1364 \text{ kJmol}^{-1}$$

2. Calculate the standard enthalpy change for the reaction:

$C(s)$ $O_2(g)$

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Given that the standard enthalpy of ^{combustion} conversion of sulphur is -297 kJmol^{-1} and the standard enthalpy of formation of H_2S and H_2O are -20.2 kJmol^{-1} and -286 kJmol^{-1}

$$\Delta H_{\text{reaction}}^{\circ} = \sum (\Delta H_f^{\circ} \text{ products}) - \sum (\Delta H_f^{\circ} \text{ reactants})$$

$$\Delta H_f^{\circ} \text{ products} = 3 \Delta H_f^{\circ} (S) + 2 \Delta H_f^{\circ} (H_2O)$$

$$= 3 \times 0 + 2 \times -286$$

$$= -572 \text{ kJmol}^{-1}$$

$$\Delta H_f^{\circ} \text{ reactants} = \Delta H_f^{\circ} (SO_2) + 2 \Delta H_f^{\circ} (H_2S)$$

$$= -297 + 2 \times -20.2$$

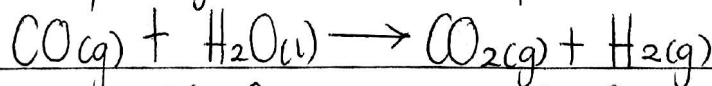
$$= -337.4 \text{ kJmol}^{-1}$$

$$\Delta H_{\text{reaction}}^{\circ} = -572 - (-337.4)$$

$$= -234.6 \text{ kJmol}^{-1} \approx -235 \text{ kJmol}^{-1}$$

Exercise.

1. Given that the heat of ~~com~~ formation of CO_2 , CO & H_2O are -394 , -110.5 , and -286 respectively. Calculate the heat of the reaction below at the same temperature



$$\Delta H_{\text{reaction}}^{\circ} = \sum (\Delta H_f^{\circ} \text{ products}) - \sum (\Delta H_f^{\circ} \text{ reactants})$$

$$= (-394 + 0) - (-110.5 + -286)$$

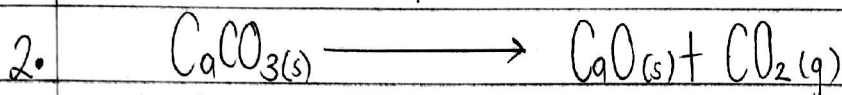
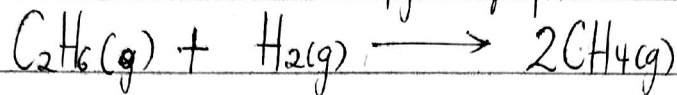
$$= (-394) - (-396.5)$$

$$= -394 + 396.5$$

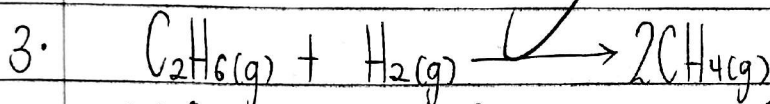
$$= \underline{\underline{-61.7 \text{ kJmol}^{-1}}}$$

2. Calculate the enthalpy of decomposition of CaCO_3 given that the enthalpy of formation of CaCO_3 , CaO and CO_2 are -1207 , -636 and -394 respectively.

3. Given that the enthalpy of formation of CH_4 is -63 kJ mol^{-1} and that of ethane C_2H_6 is -84 kJ mol^{-1} . Determine the enthalpy change for the reaction below;



$$\begin{aligned}\Delta H_{\text{reaction}}^{\circ} &= \sum (\Delta H_f^{\circ} \text{ products}) - \sum \Delta H_f^{\circ} (\text{reactants}) \\ &= (-636 + -394) - (-1207) \\ &= -1030 + 1207 \\ &= +177 \text{ kJ mol}^{-1}\end{aligned}$$



$$\begin{aligned}\Delta H_{\text{reaction}}^{\circ} &= \sum \Delta H_f^{\circ} \text{ products} - \sum \Delta H_f^{\circ} \text{ reactants} \\ &= (2 \times -63) - (0 + -84) \\ &= -126 - -84 \\ &= -42 \text{ kJ mol}^{-1}\end{aligned}$$

12th/06/2025. Standard enthalpy of combustion.

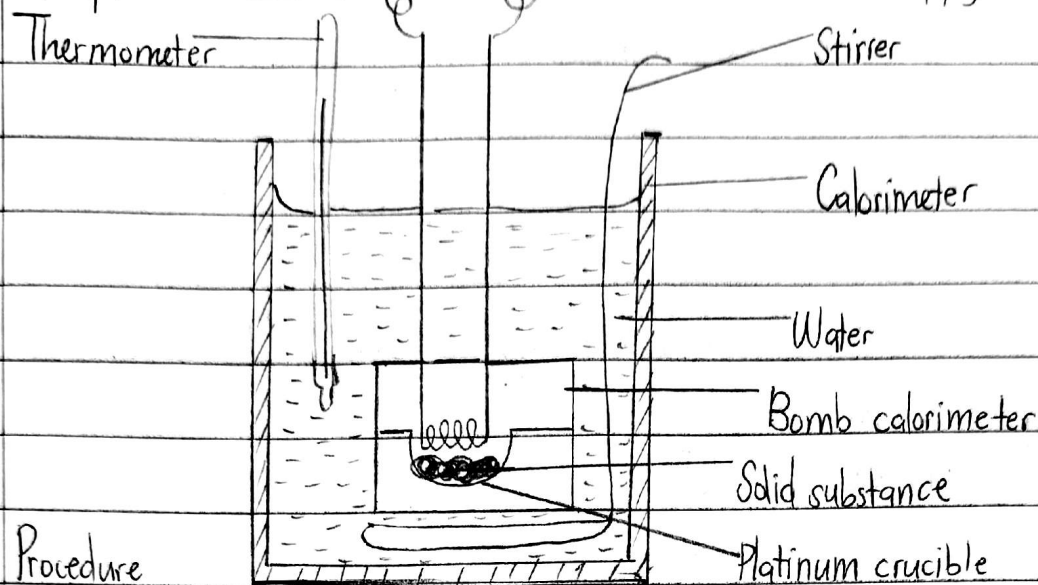
This is the heat involved / given off / liberated when one mole of a substance is completely burnt in ^(excess) oxygen under standard conditions.

Enthalpy of combustion of a pure element is equivalent to the enthalpy of formation of their respective oxides formed during the combustion process.

The enthalpy of combustion is always a negative value indicating that the enthalpy change is always exothermic.

Experiment to determine the enthalpy of combustion of solid substances.

Setup;



Procedure

A known mass of substance (M_g) under test is placed in platinum crucible in the bomb calorimeter. The bomb calorimeter is filled with oxygen at a pressure of 25 atmosphere, to ensure rapid and complete combustion.

The calorimeter is then immersed in a known volume of water ($V \text{ cm}^3$) containing a thermometer, stirrer.

The initial temperature (T_1) of water is recorded using the thermometer.

Substance is ignited by passing an electric current through a thin wire in contact with a substance. The maximum temperature of water is determined using the thermometer and recorded (T_2).

Assumptions.

- i) The heat evolved by burning of the substance is equal to the heat gained by water.
- ii) The heat gained by the calorimeter is neglected.

Treatment of results.

Heat gained by water = Heat liberated by the substance

$$\text{Heat gained by water} = \text{mass of water} \times \text{specific heat capacity of water} \times \text{Temperature change.}$$

Enthalpy

Enthalpy

(Temperature change $\Delta T = T_2 - T_1$)

$$\text{Mass of water} = \text{Volume of water} \times \rho = V \times 1 = V \text{ g}$$

$$= (V \times 4.2 \times \Delta T) \text{ J}$$

If the Rmm of the substance is M_r

M_g of the substance liberates $(V \times 4.2 \times \Delta T) \text{ J}$

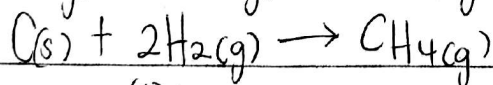
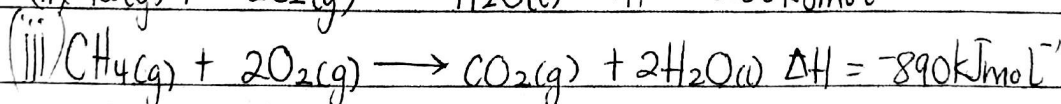
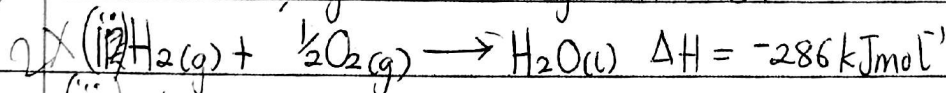
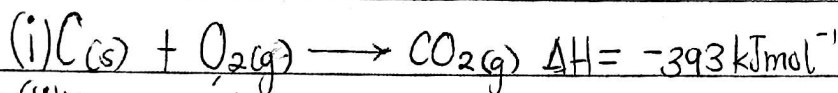
$$M_{rg} \text{ of the substance liberates } \left(\frac{V \times 4.2 \times \Delta T \times M_r}{M} \right) \text{ J}$$

Therefore the enthalpy of combustion of the substance is $-\left(\frac{V \times 4.2 \times \Delta T \times M_r}{M} \right)$

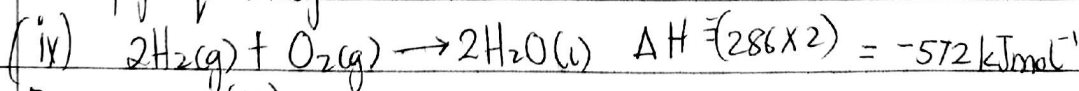
OR $\left(\frac{V \times 4.2 \times \Delta T \times M_r}{1000 \times M} \right) \text{ kJ mol}^{-1}$

Calculations involving enthalpy of formation and enthalpy of combustion.

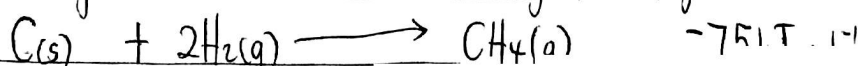
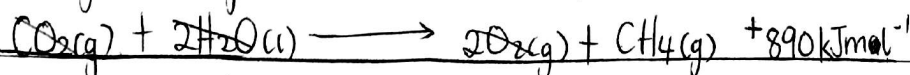
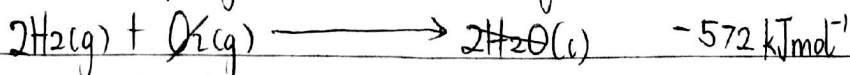
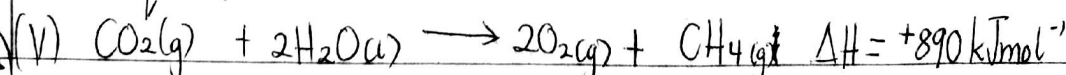
1. Calculate the enthalpy of formation of methane given that the enthalpy of combustion of Carbon is -393 kJ mol^{-1} , that of hydrogen is -286 kJ mol^{-1} and that of methane is -890 kJ mol^{-1}



Multiply eqn (ii) by 2

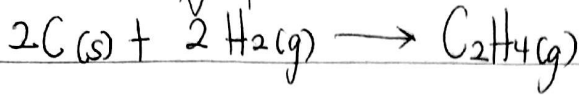


Reverse eqn (iii)



are added.

2. Calculate the enthalpy for the reaction below

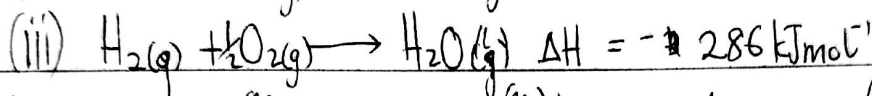
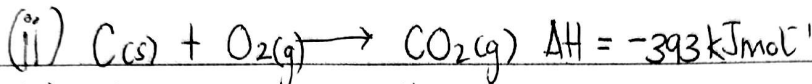
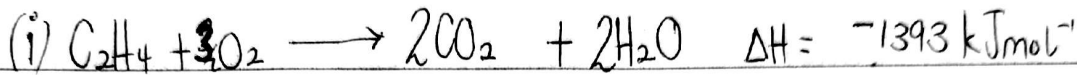


Given that:

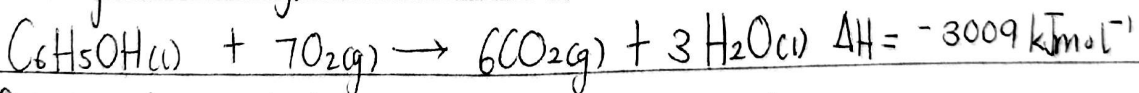
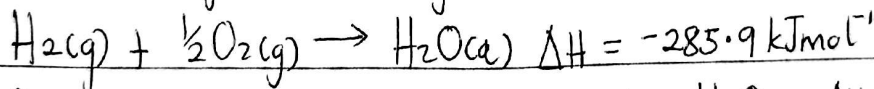
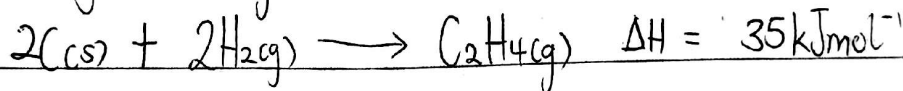
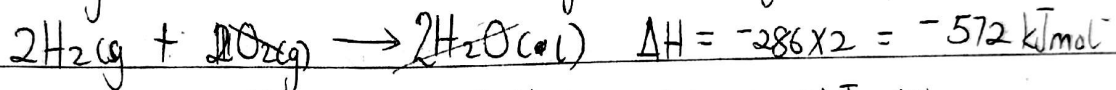
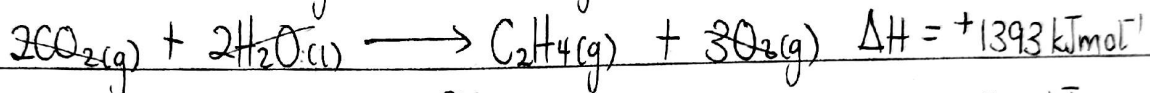
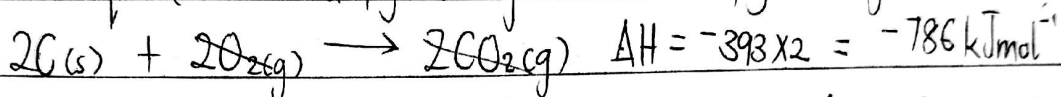
Enthalpy of combustion of carbon is -393 kJmol^{-1}

" " " " Hydrogen is -286 kJmol^{-1}

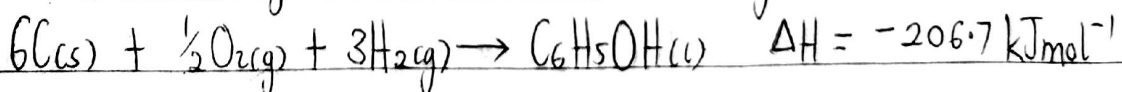
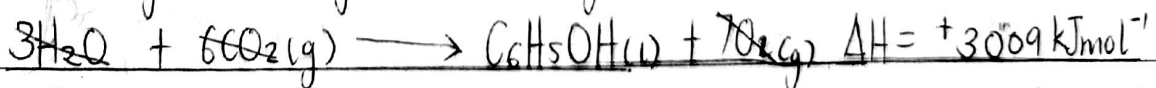
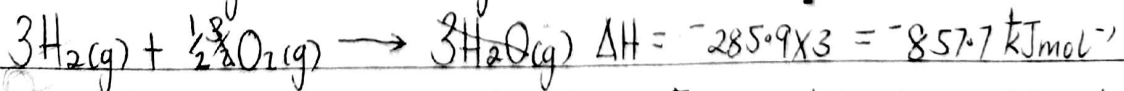
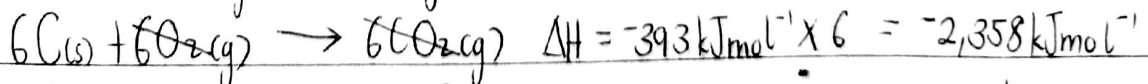
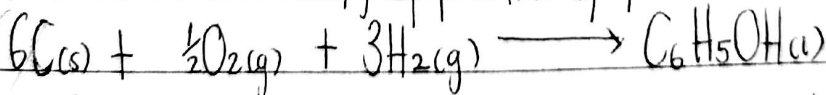
" " " " Ethene is -1393 kJmol^{-1}



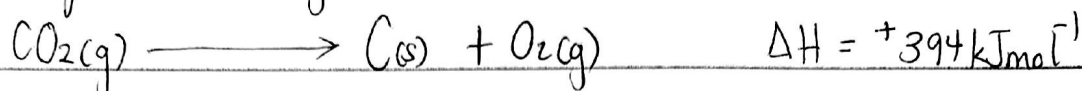
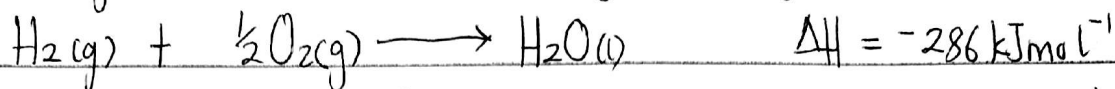
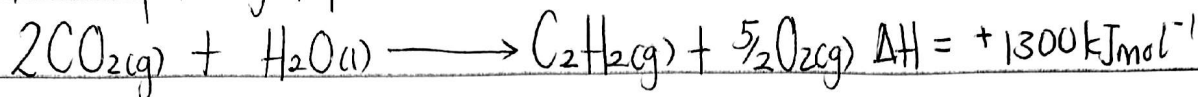
Reverse eqn (i) and multiply (ii) by 2 and multiply (iii) by 2



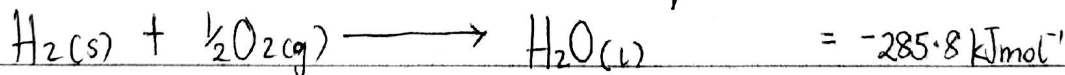
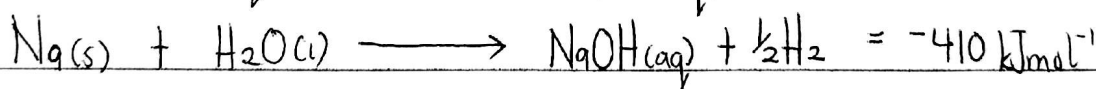
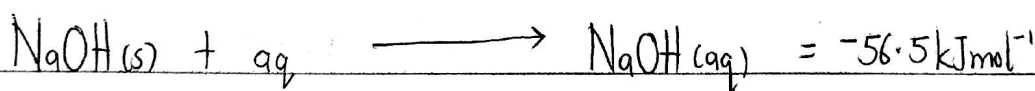
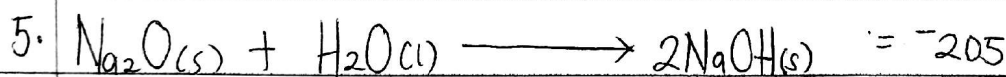
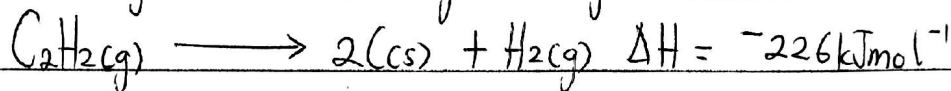
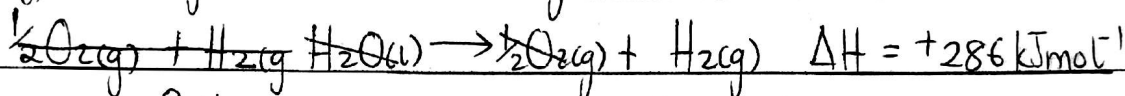
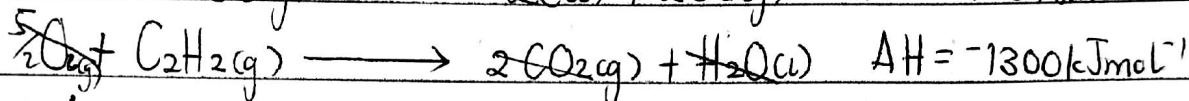
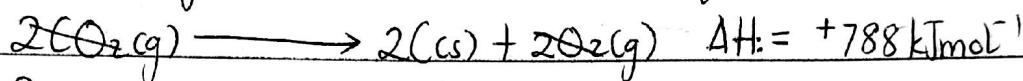
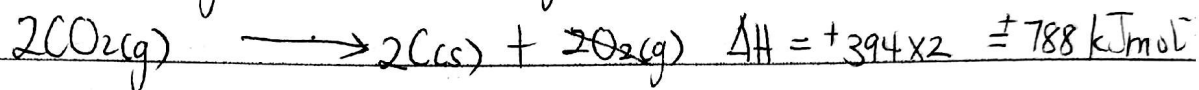
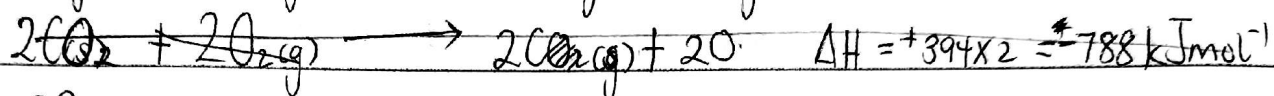
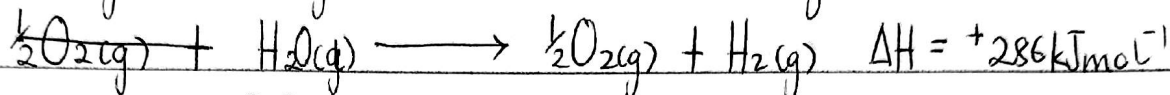
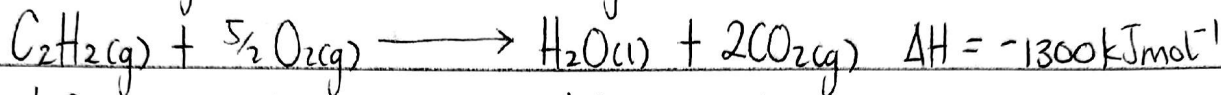
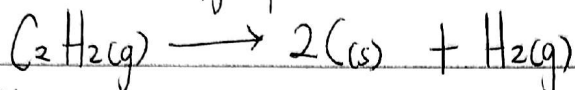
Calculate the standard enthalpy of formation of phenol.



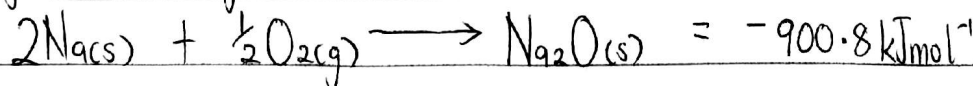
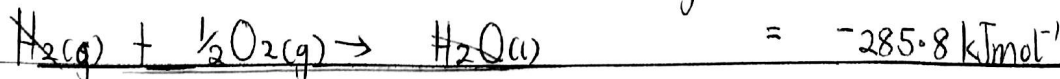
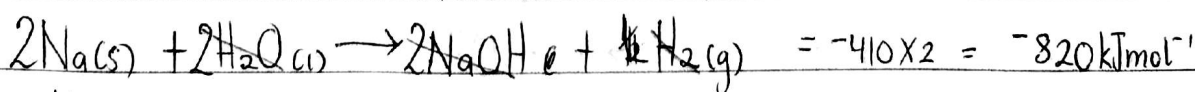
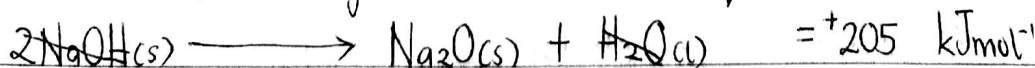
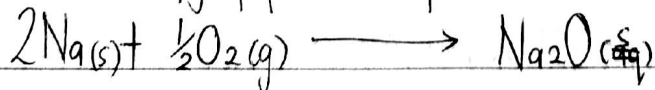
4. Given the following information.



Calculate the enthalpy of



(a) Calculate the enthalpy of formation of Na_2O .



BORN HABER CYCLE

When using the born haber cycle;

- (i) The states of all reactants and products must be shown
- (ii) The energy changes for each equation must be indicated.
- (iii) The mini equations must be balanced

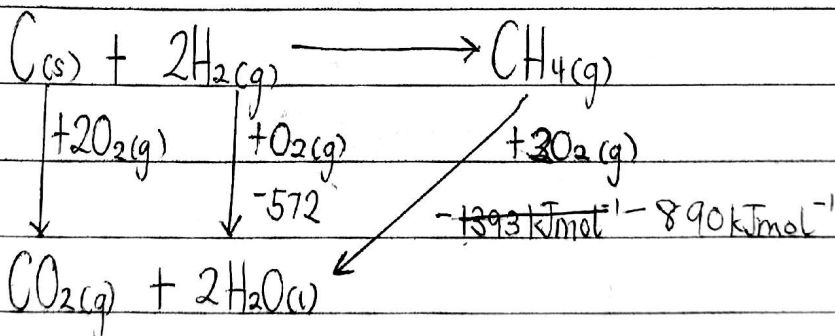
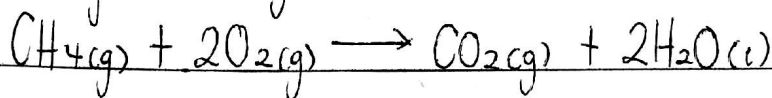
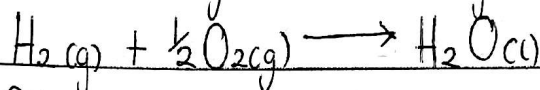
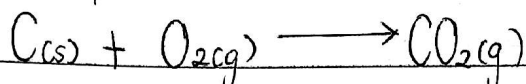
Examples

(1) Calculate the ΔH_f of methane.

$$\Delta H_f \text{ of Carbon} = -393 \text{ kJmol}^{-1}$$

$$\Delta H_c \text{ of Hydrogen} = -286 \text{ kJmol}^{-1}$$

$$\Delta H_c \text{ of CH}_4 = -890 \text{ kJmol}^{-1}$$

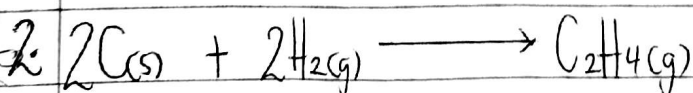


Applying Hess's law

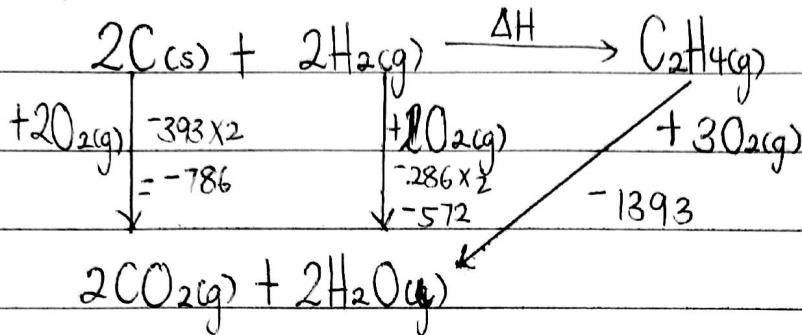
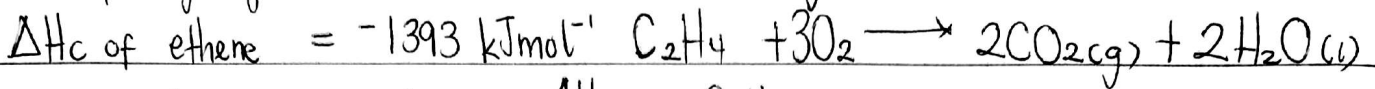
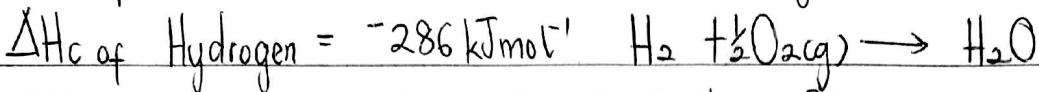
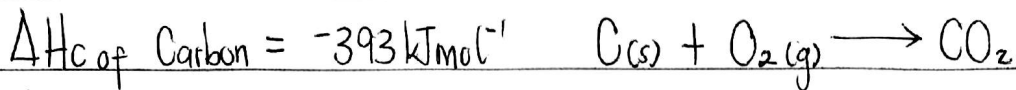
$$\Delta H + -890 = -393 + -572$$

$$\Delta H = -393 - 572 + 890$$

$$\Delta H = -75 \text{ kJmol}^{-1}$$



Given that:



$$\Delta H + -1393 = -786 + -572$$

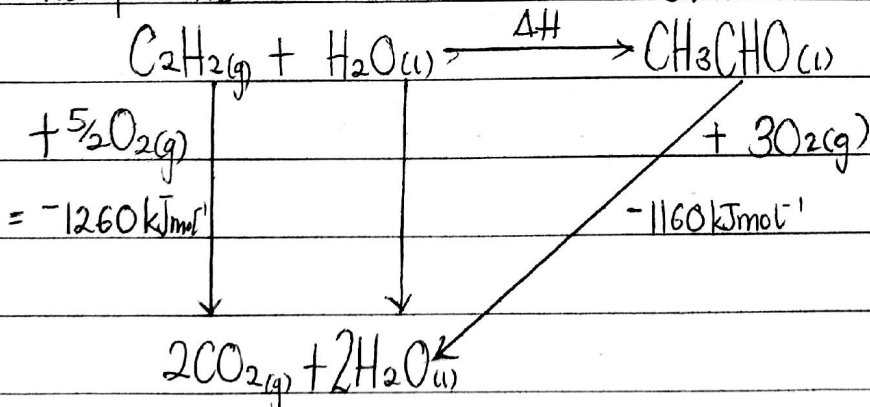
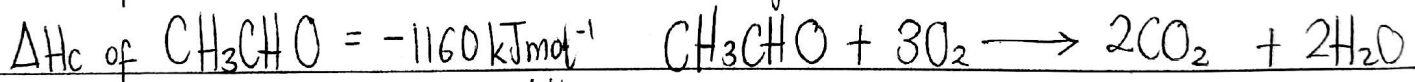
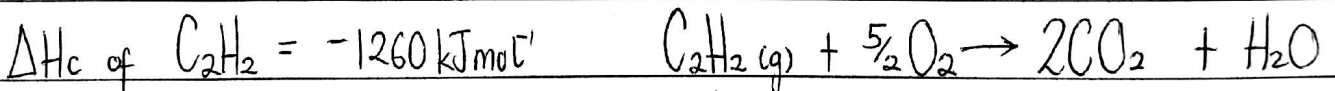
$$\Delta H = -786 + -572 + -1393$$

$$\Delta H = -1358 + 1393$$

$$\Delta H = +35 \text{ kJmol}^{-1}$$

Calculate the heat of the reaction: $\text{C}_2\text{H}_2 + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{CHO}$

Given that

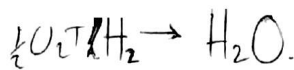


$$\Delta H + -1160 = -1260$$

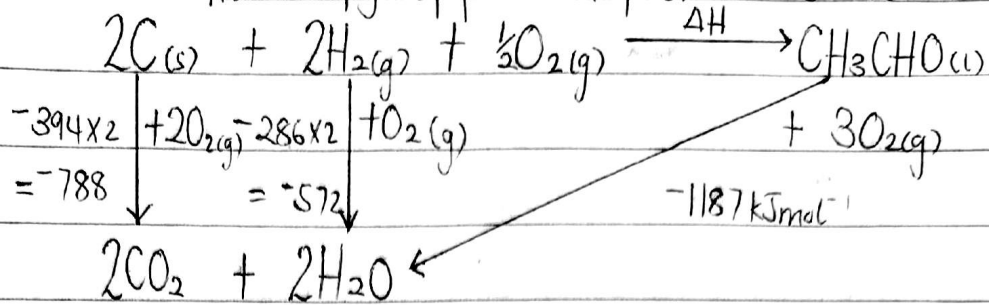
$$\Delta H = -1260 + +1160$$

$$\Delta H = -100 \text{ kJmol}^{-1}$$

4. The enthalpy of combustion of carbon, Hydrogen and ethanol are -394 , -286 & -1187



respectively. Calculate the enthalpy of formation of ethanal (CH_3CHO)



$$\Delta H = -788 + -572 - -1187$$

$$= -173 kJmol^{-1}$$

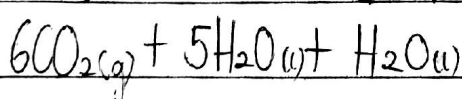
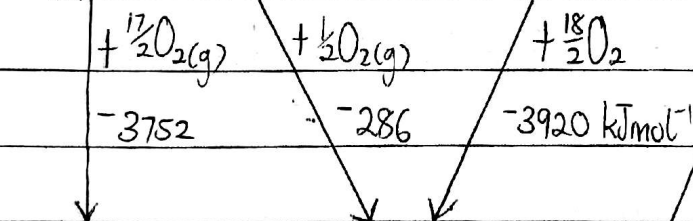
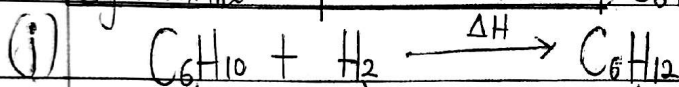
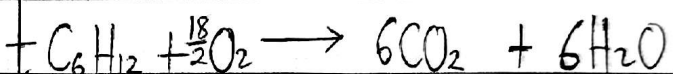
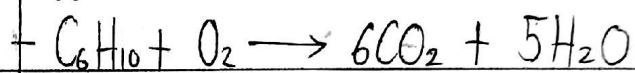
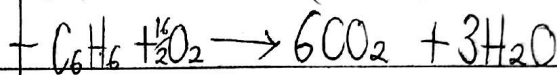
5. Given the following data:

Calculate the enthalpy of hydrogenation of:

Substance	$\Delta H_c^\circ (kJmol^{-1})$
Hydrogen	-286
Benzene	-3268
Cyclohexene	-3752
Cyclohexane	-3920

(i) Cyclohexene (C_6H_{10})

(ii) Benzene (C_6H_6)



$$\Delta H + -3920 = -3752 + -286$$

$$\Delta H = -3752 + -286 + 3920$$

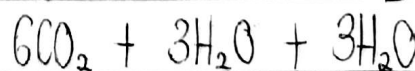
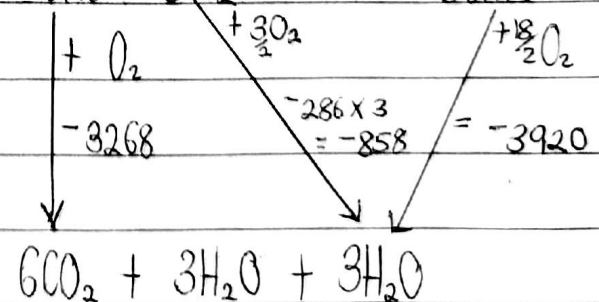
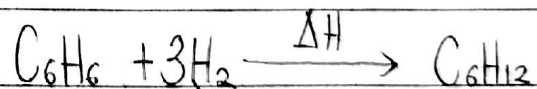
$$\Delta H = -118 kJmol^{-1}$$

$$\Delta H + -3920 = -3268 + -858$$

$$\Delta H = -3268 - 858 + 3920$$

$$\Delta H = -4126 + 3920$$

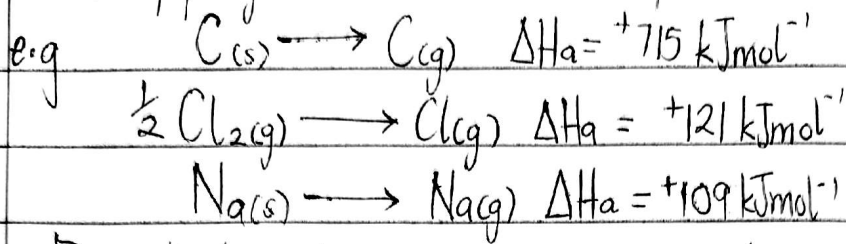
$$\Delta H = -206 kJmol^{-1}$$



17th June, 2025

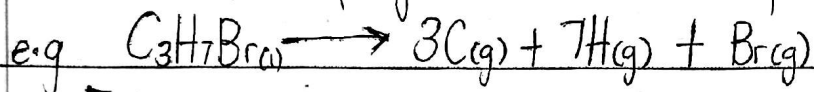
Enthalpy of atomisation.

This is the heat absorbed when an element in its normal physical state is converted into one mole of free gaseous atoms at one atmosphere and 298 K / 25°C

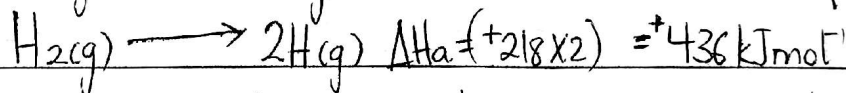


For solids, this enthalpy change may be referred to as enthalpy of sublimation. Standard enthalpy of sublimation is the heat absorbed when one mole of an element in its solid state is converted to one mole of free gaseous atoms at one atmosphere and 298 K / 25°C

For compounds, this is the heat absorbed when ^{one mole of} a compound in its normal physical state is converted into free gaseous atoms at one atmosphere and 298 K



Enthalpies of atomisation are quoted for one mole of gaseous atoms formed but not for one mole of the element e.g. if the enthalpy of atomisation of hydrogen is $+218 \text{ kJ mol}^{-1}$, this stands for: $\frac{1}{2} H_2(g) \longrightarrow H(g) \quad \Delta H_a = +218 \text{ kJ mol}^{-1}$ but not for:



Enthalpy of atomisation is always positive because energy is absorbed to convert a substance to gaseous state.

Bond energy

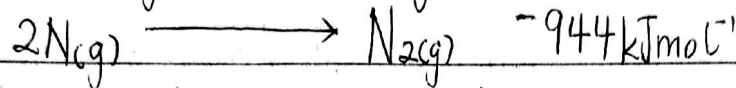
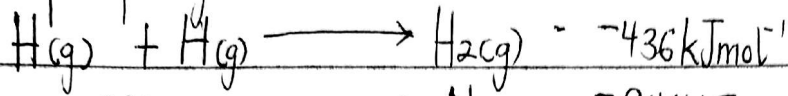
When a bond between two atoms is formed, there is release of energy and the same amount of energy must be absorbed when that bond is to be broken. Therefore bond energy can be defined in terms of bond dissociation or bond formation energy.

Bond dissociation - This is the energy absorbed to break one mole of covalent bonds into free gaseous atoms at standard conditions (1 atm & 298 K). It is an endothermic reaction e.g. $H_2(g) \longrightarrow 2H(g) \quad \Delta H = +436 \text{ kJ mol}^{-1}$



For diatomic elements, bond dissociation energy $\text{BDE} = 2 \times \Delta H_{\text{atomisation}}$.

Bond formation energy - This is the energy liberated when ^{one mole of} covalent bonds is formed from free gaseous atoms at standard conditions.



Bond energy can be used to measure the strength of the bond since it's the force with which atoms are joined together. The magnitude of the bond energy depends on the following factors: (factors that affect a covalent bond)

Bond length

Calculating heat of reaction from bond energy.

Bond energies can be used to estimate the enthalpy of the reaction. The enthalpy change in the reaction is the sum of energies required to break the bond between the reactants and that liberated when new bonds are formed. This can be summarised as:

$$\Delta H_{\text{reaction}} = \sum \text{bond energy of reactants} - \sum \text{bond energy of products}$$

The formula above applies when using bond dissociation energies only.

$$= \text{Bonds broken} - \text{bonds formed}$$

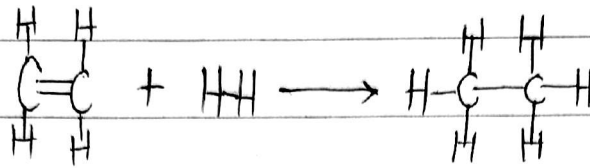
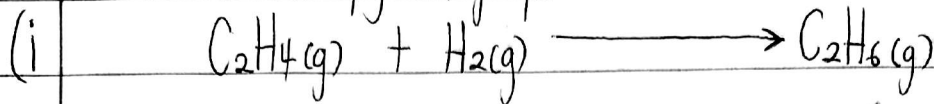
Example:

Below are some of the bond energies

Bond type	Bond energy (kJmol^{-1})
H-H	436
C-H	412
C=C	615
C-C	348

ene C_nH_{2n} double
yne C_nH_{2n-2} triple
hol C_nH_{2n+1}

Calculate the enthalpy change for the reaction;



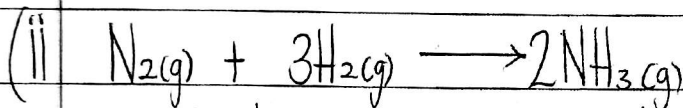
$$\Delta H_{\text{reaction}} = \sum (\text{bond energy of reactants}) - \sum (\text{bond energy of products})$$

$$= ((4 \times 412) + 615 + 436) - ((6 \times 412) + 348)$$

$$= (BE(C \equiv C) + 4BE(C-H) + BE(H-H)) - (BE(C-C) + 6BE(C-H))$$

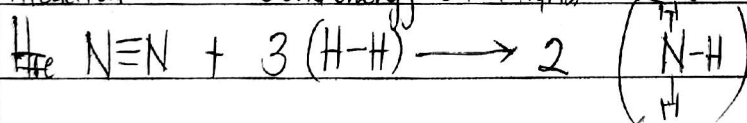
$$= 2699 - 2820$$

$$= -121 \text{ kJmol}^{-1}$$



Bond	Bond energy (kJmol ⁻¹)
N≡N	945
H-H	436
N-H	391

$$\Delta H_{\text{reaction}} = \sum (\text{bond energy of reactants}) - \sum (\text{bond energy of products})$$



$$= ((3 \times 436) + 945) - ((391 \times 3) \times 2)$$

$$= 2253 - 2346$$

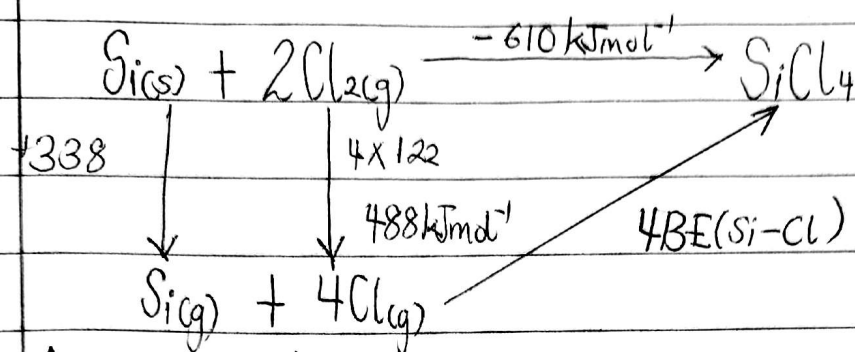
$$= -93 \text{ kJmol}^{-1}$$

18/06/25

The standard enthalpy change of formation of silicon (iv) chloride is -610 kJmol^{-1} .

Standard enthalpy changes of atomisation of silicon and chlorine are $+338$ and $+122 \text{ kJmol}^{-1}$ respectively.

- a. Use these values to construct a born haber cycle for formation of Silicon(IV)chloride.
- b. Calculate the average bond energy of Si-Cl bond.



Applying Hess's law

$$-610 = 338 + 488 + 4\text{BE}(\text{Si-Cl})$$

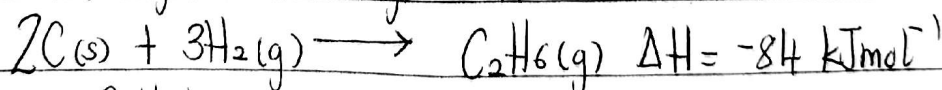
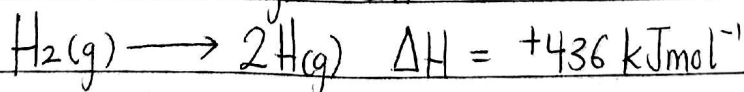
$$-610 = 826 + 4\text{BE}(\text{Si-Cl})$$

$$\frac{-1436}{4} = \frac{4\text{BE}(\text{Si-Cl})}{4}$$

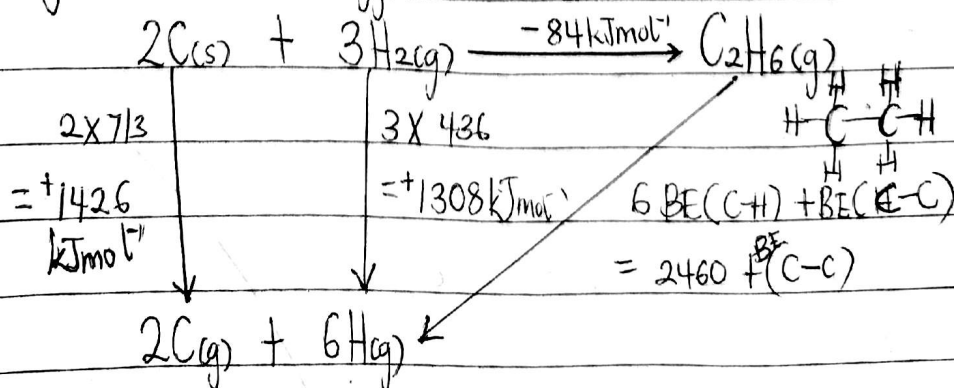
$$\text{BE}(\text{Si-Cl}) = -359 \text{ kJmol}^{-1}$$

The average bond energy of Si-Cl bond is -359 kJmol^{-1}

2. Calculate the C-C bond energy in ethane given that ;



Average C-H bond energy is 410 kJmol^{-1}



Applying Hess's law

$$-84 = 1426 + 1308 - (2460 + BE(C-C))$$

$$-84 = 1426 + 1308 - 2460 - BE(C-C)$$

$$-84 = 274 - BE(C-C)$$

$$\frac{-358}{-1} = \frac{-BE(C-C)}{-1}$$

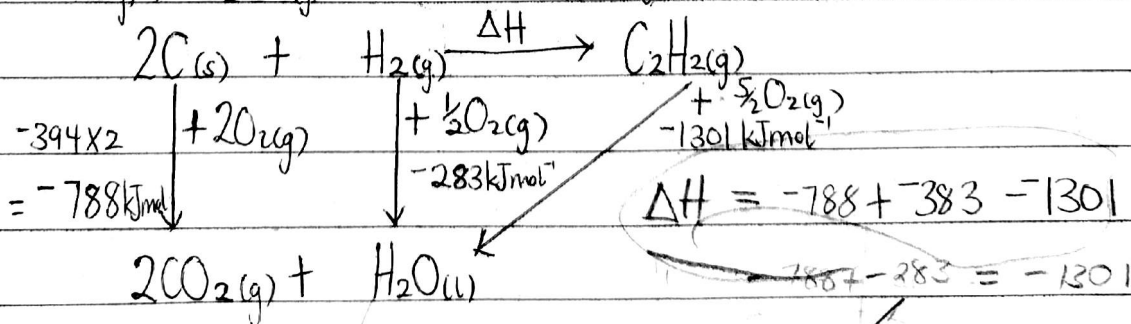
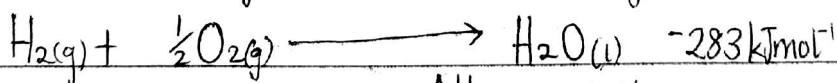
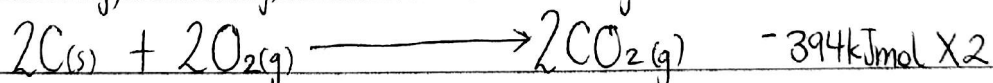
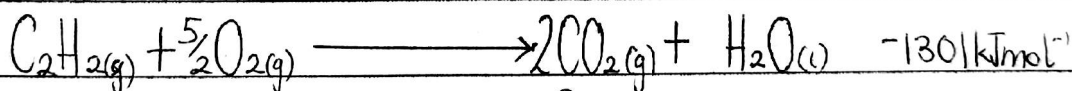
$$BE(C-C) = +358 \text{ kJmol}^{-1}$$

19th June, 2025.

The heat of combustion of ethyne is -1301 kJmol^{-1} . If the enthalpy of formation of CO_2 and H_2O are -394 and -283 kJmol^{-1} respectively, determine the enthalpy of formation of ethyne.

(b) Given that the heat of atomisation of hydrogen and carbon are $+218$ and $+713 \text{ kJmol}^{-1}$ and the C-H bond energy is -410 kJmol^{-1} , determine the $\text{C}\equiv\text{C}$ bond energy.

Solution.



$$\Delta H = -788 + -283 - 1301$$

$$-788 - 283 = -1071$$

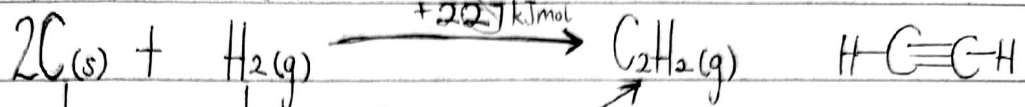
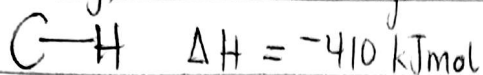
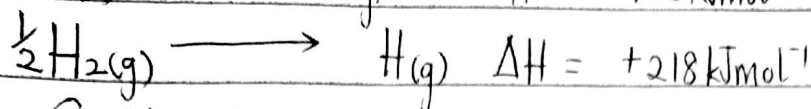
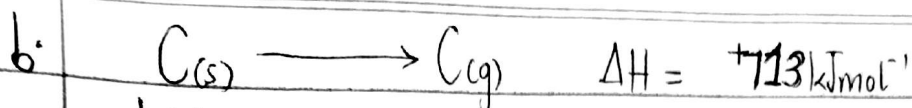
Applying Hess's law

$$\Delta H + -1301 = -788 + -283 = -1071$$

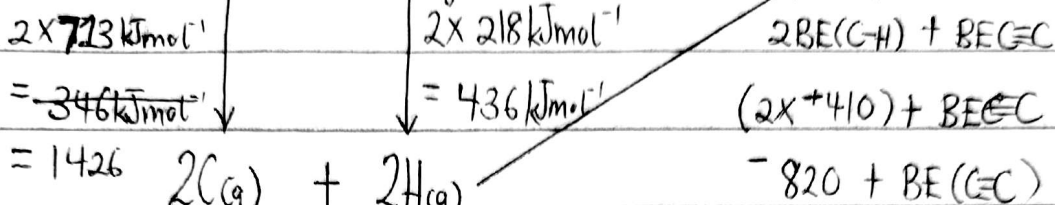
$$\Delta H = -1071 + 1301$$

$$\Delta H - 1301 = -1071$$

$$\Delta H = -1071 + 1301$$



Polish
Nonwegian



$227 = \frac{1426}{2} + 436 + -820 + BE(C \equiv C)$

$227 = 1602 + BE(C \equiv C)$

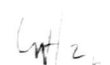
$227 - 1602 = BE(C \equiv C)$

$BE(C \equiv C) = -1375 \text{ kJmol}^{-1} - 815 \text{ kJmol}^{-1}$

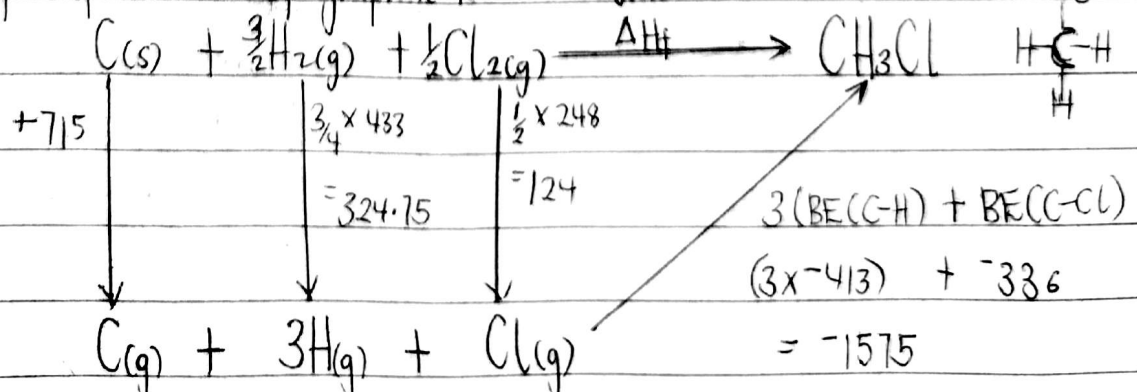
The bond energy of $C \equiv C$ is $-1375 \text{ kJmol}^{-1} - 815 \text{ kJmol}^{-1}$

Calculate the enthalpy of formation of Chloromethane given that

Bond Bond energy (kJmol^{-1})

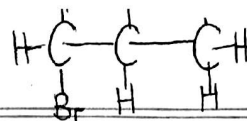


Enthalpy of vapourisation of graphite is $+715 \text{ kJmol}^{-1}$



$\Delta H_f = 715 + 324.75 + 124 + -1575$

$$C_nH_{2n+2} \text{ Alkanes } \Delta H_r^{+ve -ve} = BE(\text{Reactants}) - BE(\text{Products})$$



PAGE NO:

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C_nH_{2n} Alkene

$$\Delta H_f = -411.25 \text{ kJmol}^{-1}$$

C C C

21st June, 2025

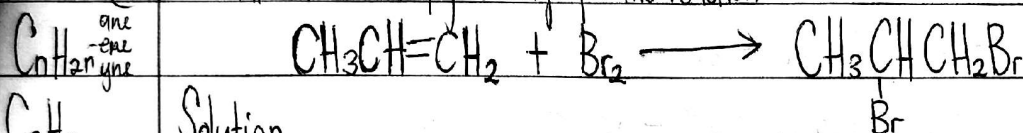
C_nH_{2n}	Bond	H-H	C-Br	C-C	C=C	Br-Br	C-H
	BE (kJmol ⁻¹)	435	415	284	256	193	413

a. Given that the enthalpy of atomisation of carbon is $+713 \text{ kJmol}^{-1}$, Calculate the enthalpy of formation of:

(i) 1,2-dibromopropane

(ii) Propene

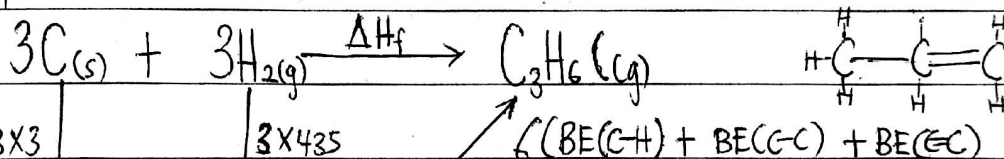
(b) Calculate the enthalpy change for the reaction.



C_3H_6

Solution.

CH_3CH (ii)



$$+713 \times 3$$

$$3 \times 435$$

$$6(BE(C-H)) + BE(C-C) + BE(C=C)$$

$$= 2139$$

$$= 1305$$

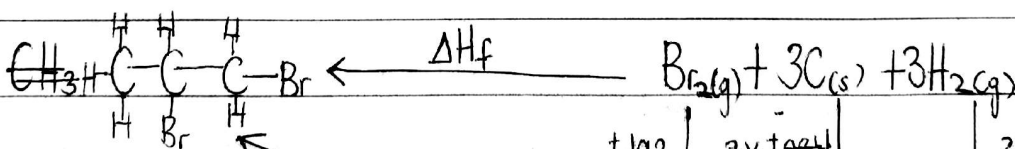
$$(6 \times 413) + 284 + 256$$



$$= -3018 \text{ kJmol}^{-1}$$

$$\Delta H_f = 2139 + 1305 - 3018$$

$$\Delta H_f = 426 \text{ kJmol}^{-1}$$



$$\begin{array}{c} +193 \\ \downarrow \\ 3 \times 284 \\ \downarrow \\ 713 \times 3 \\ = 2139 \end{array} \quad \begin{array}{c} \downarrow \\ 3 \times 435 \\ = 1305 \end{array}$$



$$2BE(C-Br) + 2BE(C-C) + 6(BE(C-H))$$

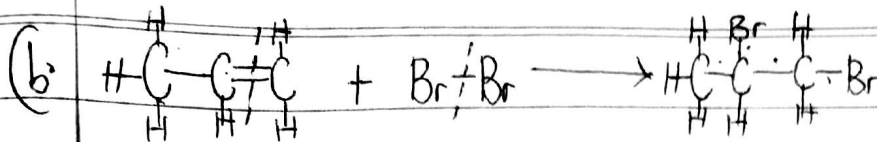
$$(2 \times 413) + (284 \times 2) + (6 \times 413)$$

$$= -3876 \text{ kJmol}^{-1}$$

$$\Delta H_f = +193 + 2139 + 1305 - 3876$$

$$= -239 \text{ kJmol}^{-1}$$

ΔH



$$\Delta H_r = \sum (\text{Energy of bonds broken}) - \sum (\text{Energy of bonds formed})$$

$$= (\text{C}=\text{C} + \text{Br}-\text{Br}) - (2(\text{C}-\text{Br}) + (\text{C}-\text{C}))$$

$$= (256 + 193) - ((2 \times 415) + 284)$$

$$= 449 - 830 = -381$$

$$= -665 \text{ kJmol}^{-1} \quad \text{OR}$$

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

$$= (6(\text{BE}(\text{C}-\text{H}) + \text{BE}(\text{C}-\text{C}) + \text{BE}(\text{C}=\text{C}) + \text{BE}(\text{Br}-\text{Br})) - (6(\text{BE}(\text{C}-\text{H}) + 2(\text{C}-\text{Br}) + 2(\text{C}-\text{C})))$$

$$= ((6 \times 413) + 284 + 256 + 193) - ((6 \times 413) + (2 \times 415) + (2 \times 284))$$

$$= 2478 + 3018 - 3876 = 3211 - 3876$$

$$= -665 \text{ kJmol}^{-1}$$

BORN HABER CYCLES AND LATTICE ENERGY

Lattice energy.

Ionic lattice arise as a result of attraction between the positive and negative ions. The measure of the strength of the lattice is given by the lattice energy and the higher the value of lattice energy, the stronger the ionic bond formed.

Lattice energy is the heat absorbed when one mole of a crystal lattice or ionic compound is broken into its constituent gaseous ions. It is also referred to as the heat given out when one mole of ionic compound / crystal lattice is formed from its constituent gaseous ions.

Lattice energy is the enthalpy change that occurs when one mole of crystal lattice or ionic compound / solid is formed from or broken into its constituent gaseous ions at standard conditions (1 atmosphere & 25°C)

2/4
2/5
2/9

Factors affecting lattice energy.

- **Ionic charge** - Lattice energy is directly proportional to the ionic charge because the higher the ionic charge, the stronger the electrostatic forces of attraction between the oppositely charged ions. This implies that more energy is required to break down the crystal lattice **eg** consider sodium chloride & magnesium chloride whose lattice energies are -771 kJmol^{-1} and -2524 kJmol^{-1} respectively.

Magnesium chloride has a greater lattice energy than NaCl because Magnesium ions have a greater charge than sodium ions hence they attract the Cl^- ions more strongly than the Sodium ions therefore the magnitude of electrostatic forces of attraction between Mg^{2+} and Cl^- are stronger than those between the Na^+ and Cl^- ions.

- **Ionic radius** - Lattice energy is inversely proportional to the ionic radius i.e. larger cations have less attraction for the anions because of reduced ^{effective nuclear} attraction. The smaller the ionic radius the closer the oppositely charged ions and the stronger the ionic bond.

Consider the lattice energies of the following compounds:

ionic radius	Compound	Lattice energy (kJmol^{-1})
○	NaF	917
○	NaCl	771
○	NaBr	763
○	NaI	684

irrisability ↑
unthe table

Lattice energy decreases from sodium fluoride to sodium Iodide. This is because down the group of halogens they form ions with increasing ionic radius. Attraction between sodium ions and the anions decreases with increase in the ionic radius of the halides therefore the strength of the electrostatic forces of attraction between sodium ion and the anions decreases from NaF to NaI.

Calculation of lattice energy.

Calculation of lattice energy using born-haber cycle.

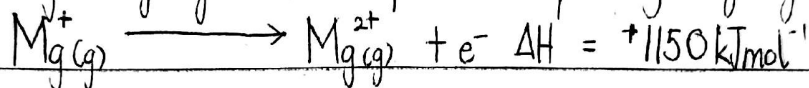
The following energy changes are involved in calculation of lattice energy ;

→ Atomisation / bond dissociation energy

→ Ionisation energy of metals - this is the minimum amount of energy ^(absorbed) required to remove 1 mole of electrons from one mole of free gaseous atoms / ions to form one mole of positively charged gaseous ions. eg $\text{Na(g)} \longrightarrow \text{Na}^+(\text{g}) + \text{e}^- \Delta H_{\text{IE}} = +414 \text{ kJ mol}^{-1}$

- The first ionisation energy is the minimum amount of energy required to remove 1 mole of e^- from one mole of gaseous atoms to form one mole of unipositively charged ions.

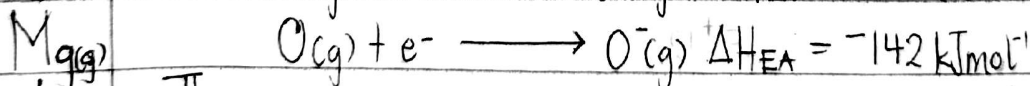
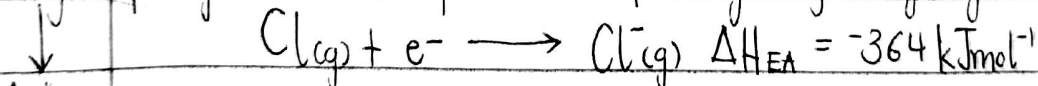
- Second ionisation energy is the minimum energy required to remove an electron from a unipositively charged gaseous ion to form a double positively charged gaseous ion. ie



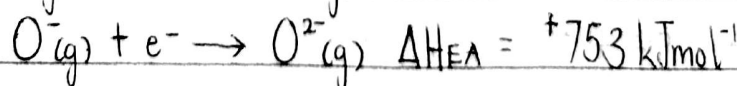
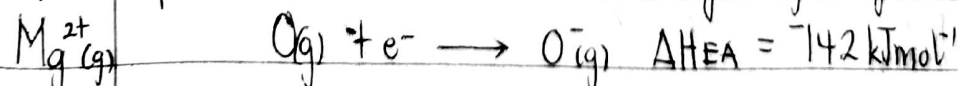
→ Electron affinity - energy change that occurs when one mole of electron is added to one mole of free gaseous atoms / ions to form one mole of $^-$ vely charged gaseous ions. It is also successive ie it has first electron affinity and second electron affinity etc.

First electron affinity - energy given out when one mole of electrons is added to one mole of

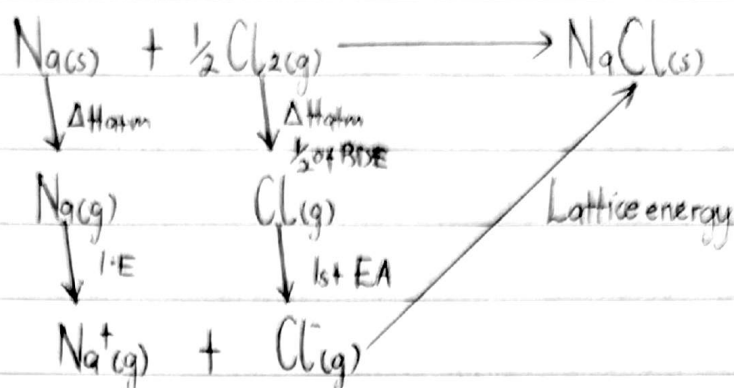
Mg(g) free gaseous atoms to form one mole of uninegatively charged gaseous ions.



The first electron affinity is exothermic ^{where} for as the subsequent electron affinity ie second, third, fourth are endothermic because energy is required to overcome electron repulsion in order to add an electron the negatively charged ions.



→ Lattice energy



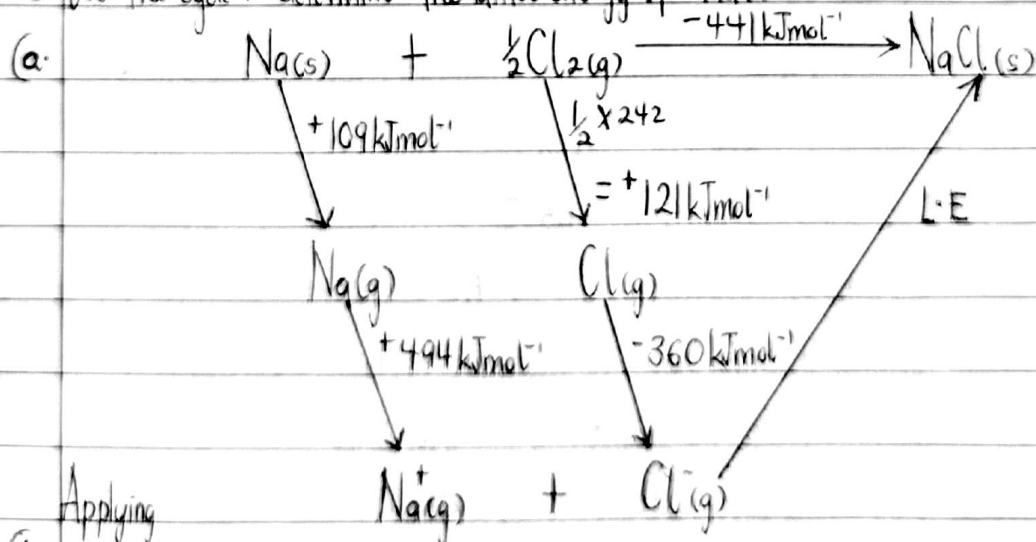
Example

1. You are provided with the following thermochemical data:

Process	$\Delta H^{\circ} (\text{kJ mol}^{-1})$
Bond dissociation energy of Chlorine	+242
Atomisation of sodium	+109
First ionisation energy of sodium	+494
First electron affinity of chlorine	-360
Enthalpy of formation of NaCl	-441

a. Draw a born haber cycle for the formation of NaCl from its elements.

b. Use the cycle to determine the lattice energy of NaCl



Applying

(b) $-441 = +109 + 494 + -360 + 121 + \text{L.E.}$

$$-441 = 364 + \text{L.E.}$$

$$\text{L.E.} = -441 - 364$$

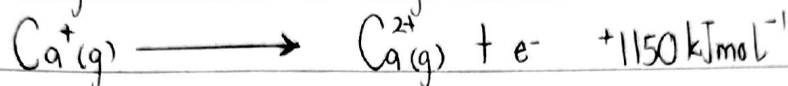
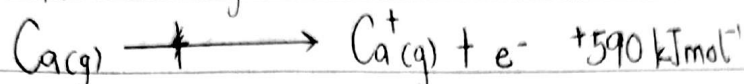
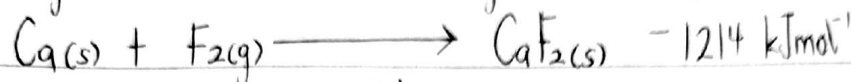
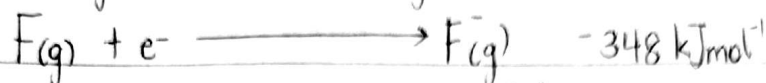
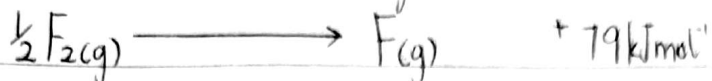
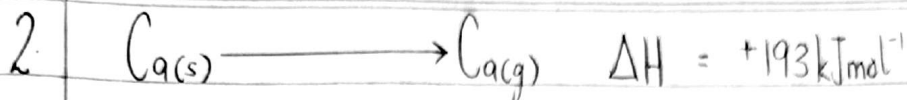
$$\text{L.E.} = -805 \text{ kJ mol}^{-1}$$

$$\Delta H = \frac{1}{2} BE$$

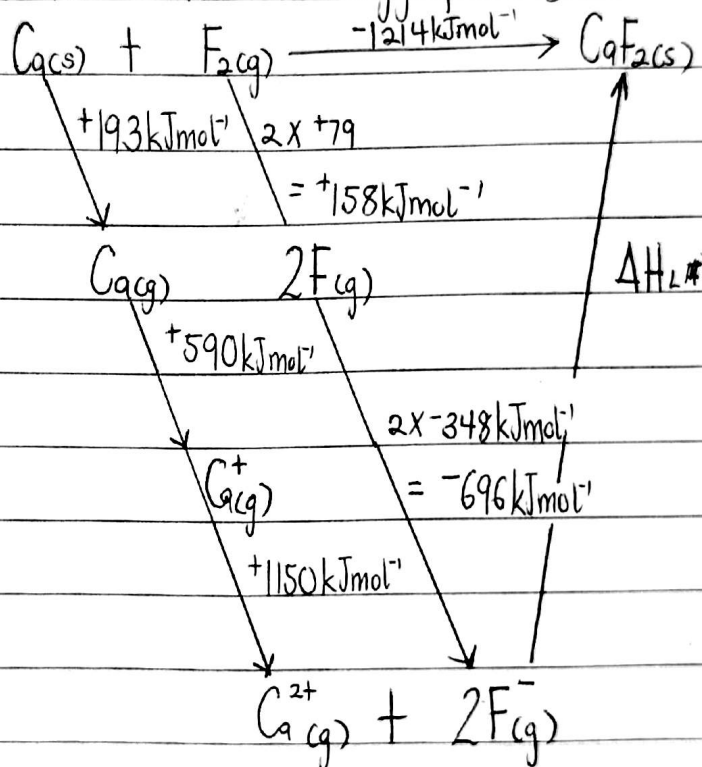
$$BE = 2 \Delta H$$

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Draw a cycle and determine the lattice energy of $\text{CaF}_2(\text{s})$



$$-1214 = 193 + 158 + 590 + 1150 + -696 + \Delta H_L$$

$$-1214 = 1395 + \Delta H_L$$

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

3. Process $\Delta H^\circ (\text{kJ mol}^{-1})$

Bond dissociation energy of oxygen

+498

Atomisation energy of barium

+178

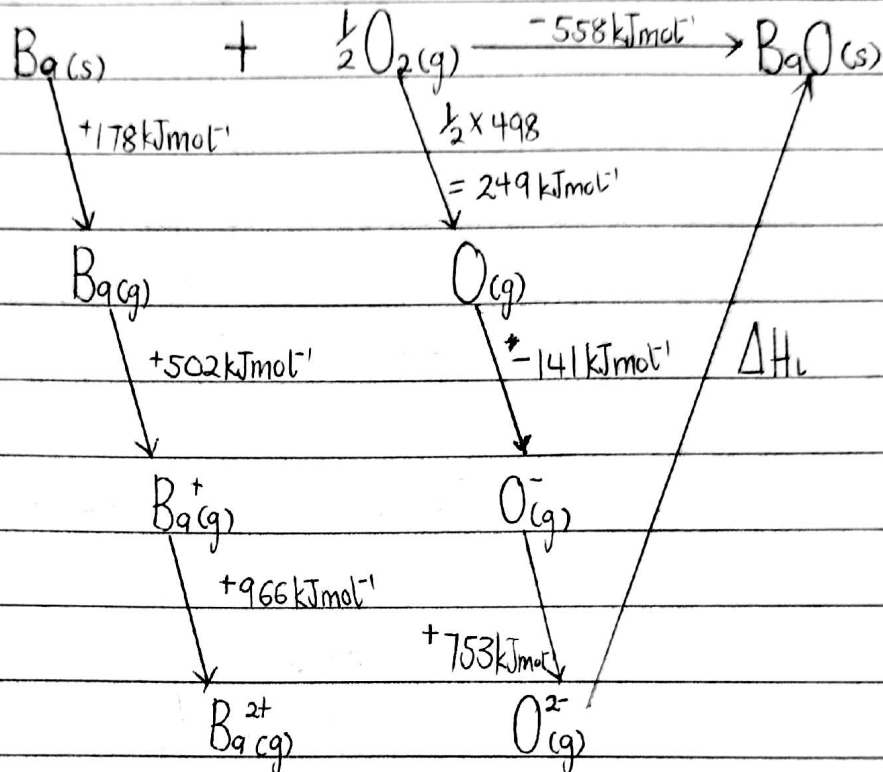
First ionisation energy of barium

+502

Second ionisation energy of barium

+966

Enthalpy of formation of bariumoxide	-558
First electron affinity of oxygen	-141
Second electron affinity of oxygen	+753



$$-558 = +178 + 249 + 502 + 966 + 753 + \Delta H_L - 141$$

$$-558 = 2507 + \Delta H_L$$

$$\Delta H_L = -3065 \text{ kJmol}^{-1}$$

4. Aluminium reacts with Oxygen to form Al_2O_3 according to the following equation:



You are given the following data;

First ionisation energy of Al = +578 kJmol⁻¹

1st Electron affinity of O_2 = -141.4 kJmol⁻¹

Heat of atomisation of Al = +324.3 kJmol⁻¹

2nd " " " " = +844 kJmol⁻¹

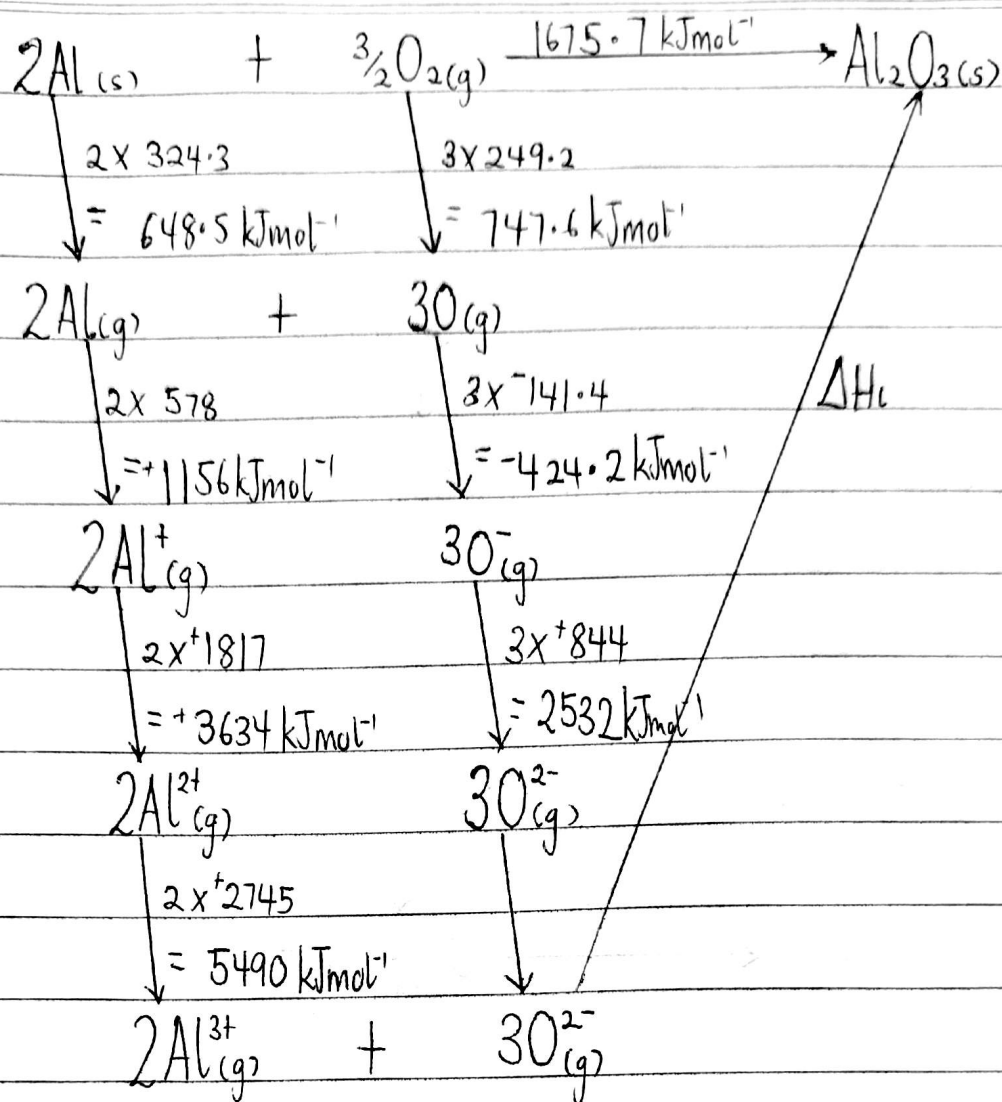
2nd ionisation energy of Al = +1817 kJmol⁻¹

(i) Draw an energy level diagram for the formation of Al_2O_3

3rd ionisation energy of Al = +2745 kJmol⁻¹

(ii) Calculate the lattice energy of Al_2O_3

Atomisation energy of O_2 = +249.2 kJmol⁻¹



Applying Hess's law.

$$1675.7 \text{ kJmol}^{-1} = 5490 + 3634 + 2532 + -424.2 + 747.6 + 1156 + 648.5 + \Delta$$

$$1675.7 = 13783.9 + \Delta H_L$$

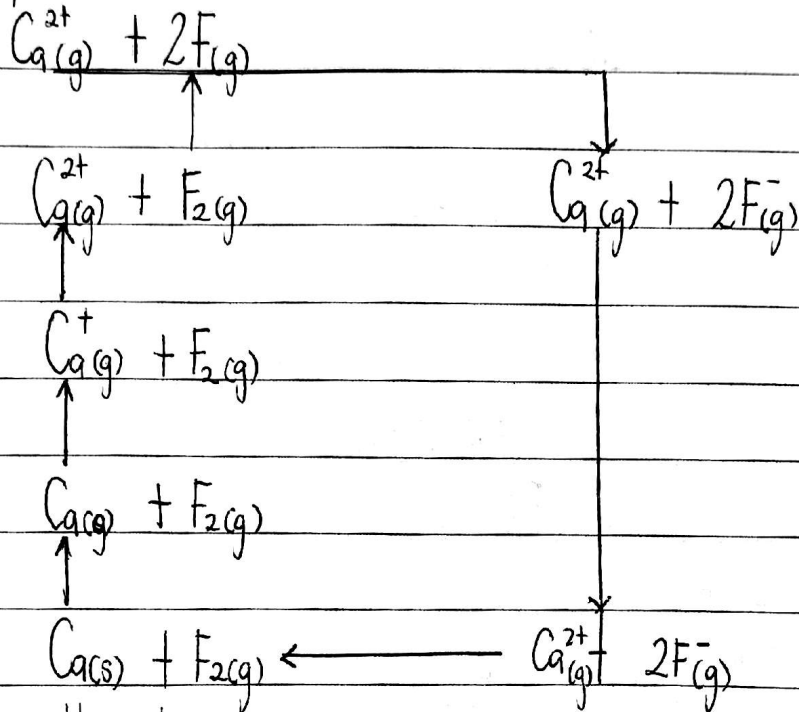
$$1675.7 - 13783.9 = \Delta H_L$$

$$\Delta H_L = -12108.2 \text{ kJmol}^{-1}$$

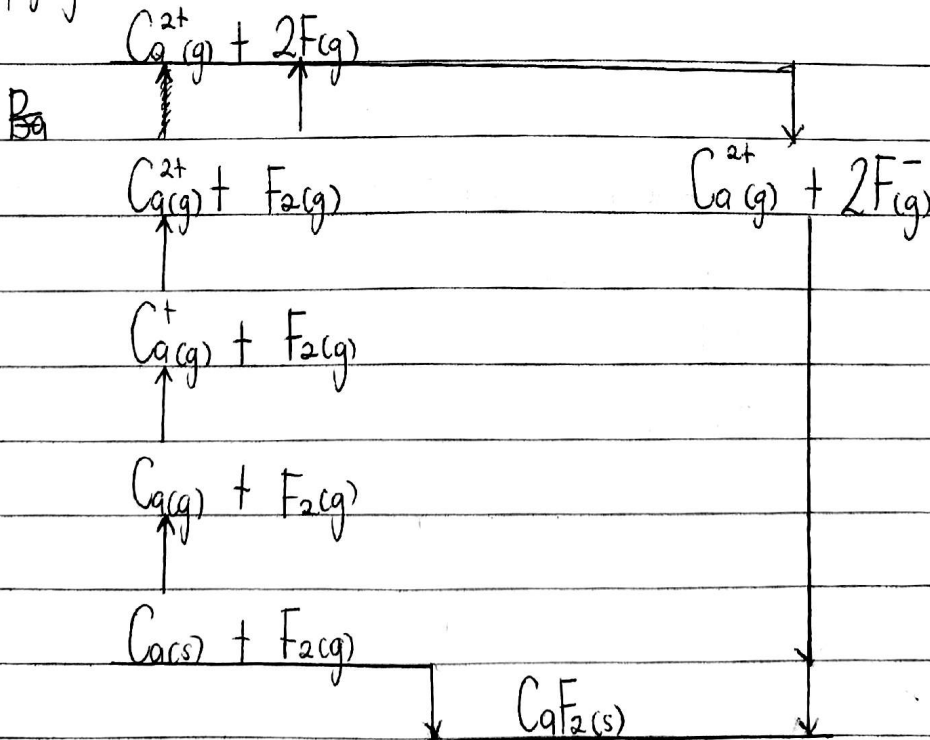
Tuesday 24th June, 2025

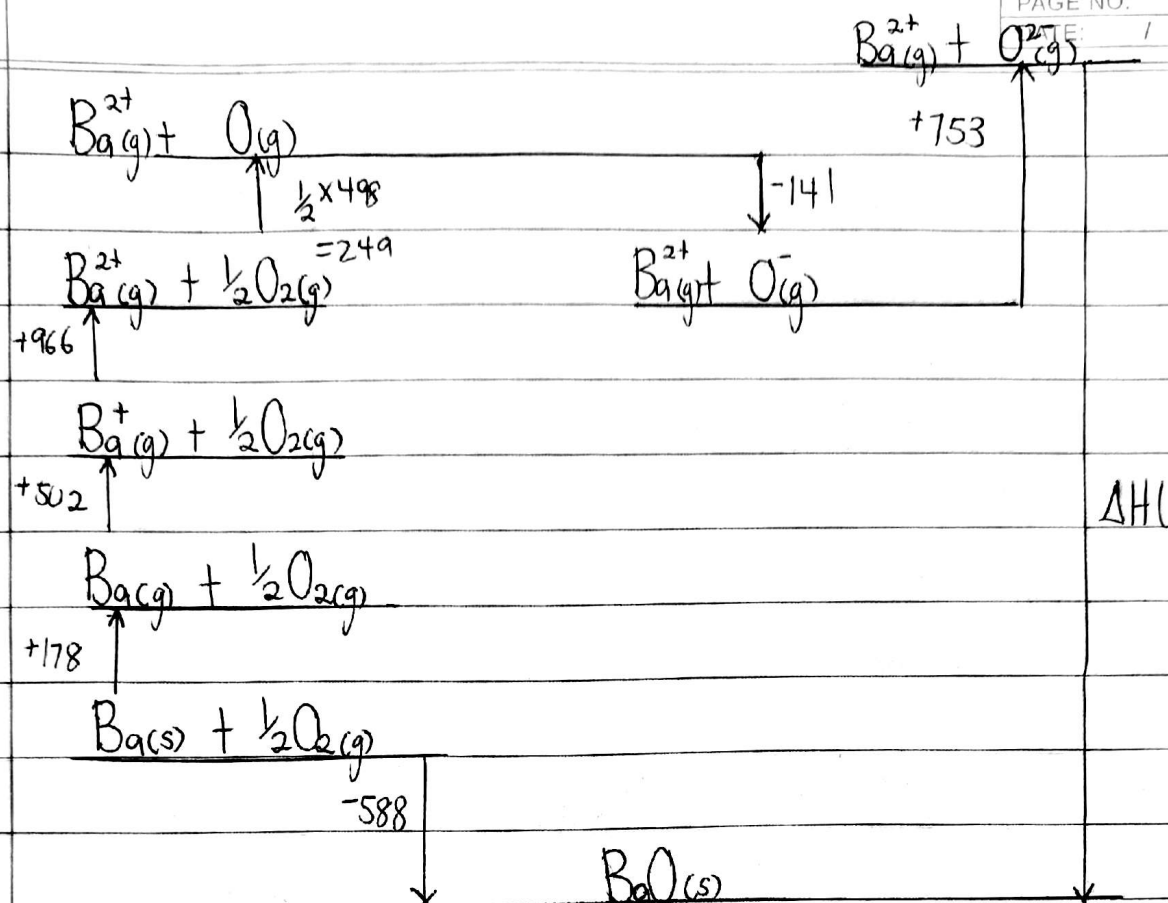
Using energy level diagram to find lattice energy

Example



Applying Hess's law





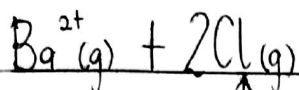
Applying Hess's law.

$$-588 = \Delta H_l + 753 + 741 + +966 + 502 + 178 + 249$$

$$-588 = \Delta H_l + 2507$$

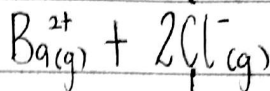
$$\Delta H_l = -3095 \text{ kJmol}^{-1}$$

Assignment: Process	$\Delta H (\text{kJmol}^{-1})$
Atomisation of Chlorine	+122
" " Barium	+178
1 st ionisation energy of Barium	+502
2 nd " " " "	+966
Lattice energy of barium chloride	-2069
Enthalpy of formation of BaCl_2	-860
Construct an energy level diagram and use it to calculate electron affinity of Chlorine atom.	



$$244 \text{ kJ mol}^{-1} \quad +122 \times 2$$

$$(\text{E.A. of Cl}) \times 2$$

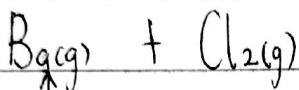


$$+966 \text{ kJ mol}^{-1}$$

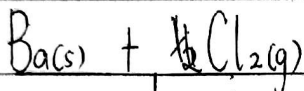


$$+502 \text{ kJ mol}^{-1}$$

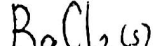
$$-2069 \text{ kJ mol}^{-1}$$



$$+178 \text{ kJ mol}^{-1}$$



$$-860 \text{ kJ mol}^{-1}$$



Applying Hess's law.

$$-860 = 2\text{EA} + -2069 + 244 + 966 + 502 + 178$$

$$-860 = 2\text{EA} - 179$$

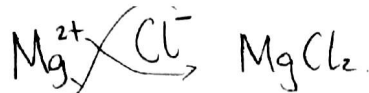
$$\frac{-681}{2} = \frac{2\text{EA}}{2}$$

$$\text{EA} = -340.5 \text{ kJ mol}^{-1}$$

26/06/2025 Solubility of ionised salts

Ionic salts dissolve in water to form a solution and during this process two energies are involved and these include: lattice energy & hydration/solvation energy.

(i) Hydration energy: This is the heat given out when one mole of gaseous ions is completely surrounded by water molecules to form an infinitely dilute solution. Hydration energy is always exothermic because during hydration ions of opposite charges are surrounded by water molecules via dipole interactions and because of these interactions heat is given out to the surrounding.



Factors affecting hydration energy.

Ionic charge. Hydration energy increases with increase in the charge of the ion. This is because highly charged ions have a high charge density and this strongly attracts many water molecules towards themselves resulting in release of a lot of energy.

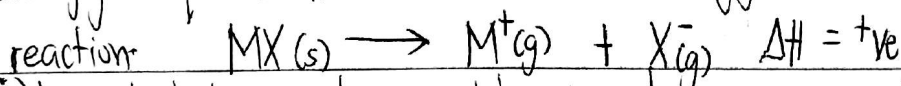
Ionic radius. Ions with small ionic radius are highly hydrated than ions with big ionic radius.

(ii) **Enthalpy of solution.** This is the enthalpy change that occurs when one mole of ionic compound is dissolved in a given volume of water to form an infinitely dilute solution at standard conditions.

Relationship between lattice energy, hydration energy and enthalpy of solution.

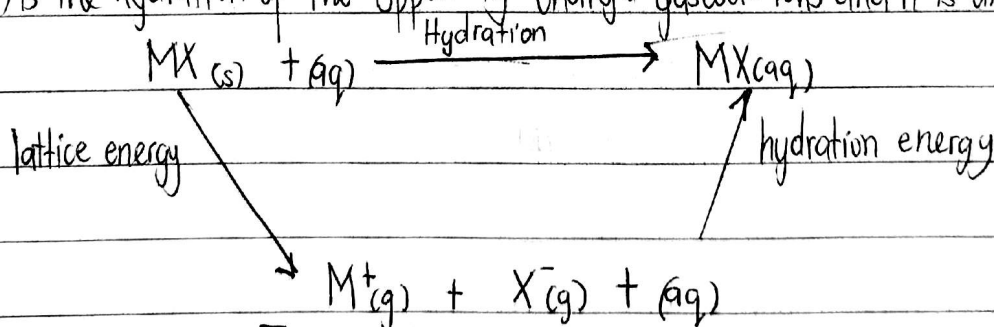
Two steps are involved when an ionic compound is dissolving in water.

i) Separating the ionic compound / lattice into two oppositely charged ions. The energy required is the lattice / dissociation energy and it is an ~~exothermic~~ endothermic reaction.



ii) Is the hydration of the oppositely charged gaseous ions and it is an exothermic reaction.

Selole



- Note i** - For an ionic compound to dissolve easily in water, the lattice dissociation energy should be overcome by the hydration energy such that an overall negative enthalpy of solution is obtained.
- When the lattice dissociation energy is greater than the hydration energy, the enthalpy of solution is endothermic and for salts that dissolve endothermically, their solubility increases with

increase in temperature.

- When the hydration energy is greater than the lattice energy, the enthalpy of solution is exothermic. For salts that dissolve exothermically, their solubility decrease with increase in temperature.
- Compounds which dissolve exothermically are more soluble than those which dissolve endothermically.

The lattice and hydration energies of salts AX and BX are given in the table below.

Salt	Lattice energy (kJmol^{-1})	Hydration energy (kJmol^{-1})
AX	+880	-860
BX	+790	-800

- Calculate the enthalpy of solution for each salt.
- Which salt is more soluble in water at a given time and give a reason for your answer.
- How do you expect the solubility of these two salts to change with temperature.

For AX: $\Delta H_{\text{solution}} = \text{Lattice energy} + \text{hydration energy}$
 $= +880 + -860$
 $= +20 \text{ kJmol}^{-1}$

For BX: $\Delta H_{\text{solution}} = \text{Lattice energy} + \text{hydration energy}$
 $= +790 + -800$
 $= -10 \text{ kJmol}^{-1}$

BX is more soluble than AX because BX dissolves exothermically while AX dissolves endothermically.

Given the following thermochemical data

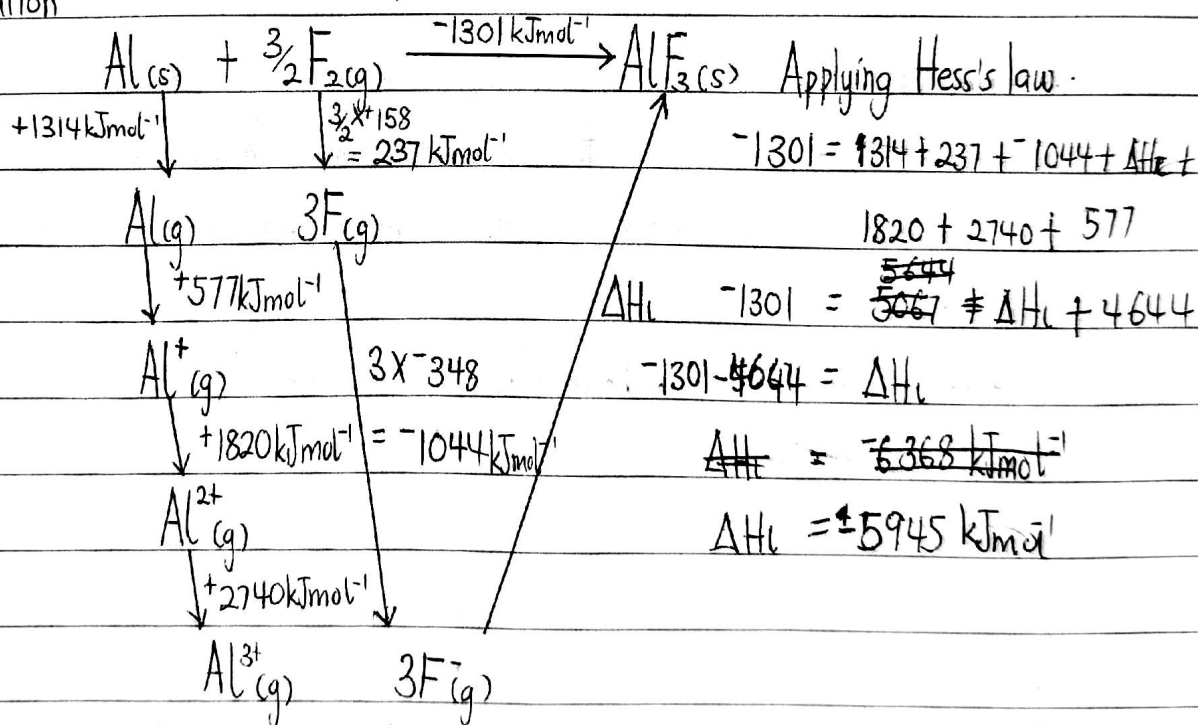
Process	ΔH° (kJmol^{-1})
Enthalpy of formation of aluminium fluoride	-1301

Enthalpy of atomisation of aluminium	+1314
BDE of Fluorine	+158
1 st atomisation energy of aluminium	+577
2nd " " " "	+1820
3rd " " " "	+2740
First electron affinity of fluorine	-348

Draw a born haber cycle for the formation of Aluminium fluoride and use it to determine its lattice energy.

Given that the hydration energies of aluminium ions and fluoride ions are -4690 & -364 kJmol^{-1} respectively. Calculate the enthalpy of solution of aluminium fluoride and comment on its solubility.

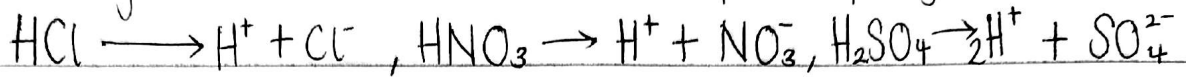
Solution



$\Delta H_{\text{solution}} = \text{Lattice energy} + \text{hydration energy}$
 $= 5945 + (-4690 - 364 \times 3)$
 $= 5945 - 5052 = +893 \text{ kJmol}^{-1}$
 $\Delta H_{\text{solution}} = +163 \text{ kJmol}^{-1}$

Strong acids Standard enthalpy of neutralisation.

-HCl This is the heat change that occurs when one mole of hydrogen ions from the acid and one mole of hydroxide ions from a base react to form one mole of water. The enthalpies of neutralisation of strong acids by strong bases are usually higher and constant. This is because there are strong acids and bases such as HCl and therefore completely ionise i.e.



And the reverse is true for weak acids because they don't undergo complete ionisation. i.e. the enthalpies of ^{neutralisation} weak acids and bases are usually low and not constant because they don't undergo complete ionisation e.g. $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$. This is also because some of the energy is absorbed and used to ionise the weak acid or alkali thus evolving less heat.

weak acid
 $\text{CH}_3\text{COOH} \rightleftharpoons$

$\text{CH}_3\text{COO}^- + \text{H}^+$

An experiment to determine enthalpy of neutralisation.

A known volume of acid V_{cm^3} and • known molarity M is put in a well insulated calorimeter. Its initial temperature t_1 is read and recorded. The same volume of base/alkali which is equimolar is put in another insulated calorimeter and its initial temperature t_2 is read and recorded.

The alkali is slowly added to the acid as we continuously stir gently. The highest temperature t_3 is read and recorded.

Assumptions

The density of the solution is equal to that of water. Also the specific heat capacity of the solution is equal to that of water. ($4.2 \text{ J g}^{-1} \text{ } ^\circ\text{C}^{-1}$).

No heat was lost to the calorimeter

Results

$$\text{Mean initial temperature} = \left(\frac{t_1 + t_2}{2} \right) ^\circ\text{C}$$

$$\text{Final temperature} = t_3$$

$$\Delta\theta = t_3 - \left(\frac{t_1 + t_2}{2}\right)$$

$$\text{Volume of solution} = V + V \\ = 2V$$

$$\text{Mass of solution} = V\rho \\ = 2V \times 1$$

$$= 2Vg$$

$$\text{Heat change} = mC\Delta\theta$$

$$= 2V \times 4.2 \left(t_3 - \left(\frac{t_1 + t_2}{2}\right) \right)$$

Initial moles of acid reacted

1000 cm³ of acid contains M moles

V cm³ of acid contain $\frac{MV}{1000}$ moles of acid.

$$\frac{MV}{1000} \times 2V \times 4.2 \left(t_3 - \left(\frac{t_1 + t_2}{2}\right) \right)$$

1000 moles of acid produced

1 mole of acid produces

1M HCl

1M KOH

When 100 cm³ of acid was added to 100 cm³ of alkali, temperature rose by 12°C. Find the enthalpy of neutralisation of HCl

$$\text{Total volume} = (100 + 100) \text{ cm}^3 \\ = 200 \text{ cm}^3$$

$$\text{Mass of solution} = V\rho \\ = 200 \text{ cm}^3 \times 1 \text{ g/cm}^3 \\ = 200 \text{ g}$$

$$\Delta\theta = 12^\circ\text{C}$$

$$\text{Heat change} = mC\Delta\theta$$

$$= 200 \text{ g} \times 4.2 \times 12$$

$$= 10,080 \text{ J}$$

Moles of HCl that reacted

1000 cm³ contain 1 mole of HCl

100 cm³ contain $\left(\frac{1}{1000} \times 100\right)$ moles of HCl

100 cm³ of solution contain 0.1 moles of HCl

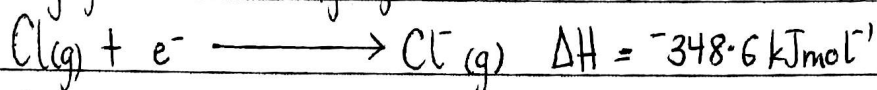
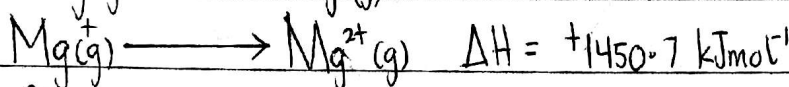
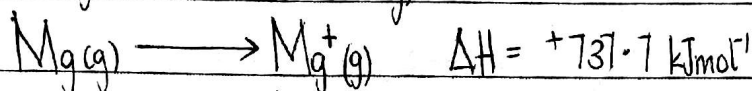
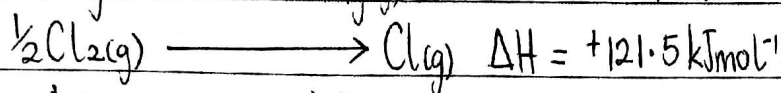
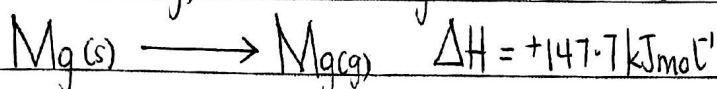
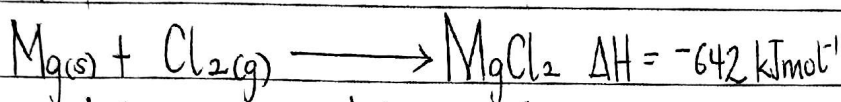
0.1 moles of HCl produced 10,080 J

$$1 \text{ mole of HCl produces } \left(\frac{10,080 \times 1}{0.1} \right) \text{ J}$$

$$= 100,800 \text{ J}$$

3rd July, 2025.

Given:



(a) Use the data above to construct an energy level diagram for the formation of solid MgCl₂.

(b) Use the diagram to calculate the lattice energy of MgCl₂.

(c) Given that the hydration energies of magnesium ions and chloride ions are -1891 kJmol⁻¹ and -381 kJmol⁻¹ respectively, calculate the enthalpy of solution of MgCl₂ and comment on its solubility in H₂O.

Mg²⁺

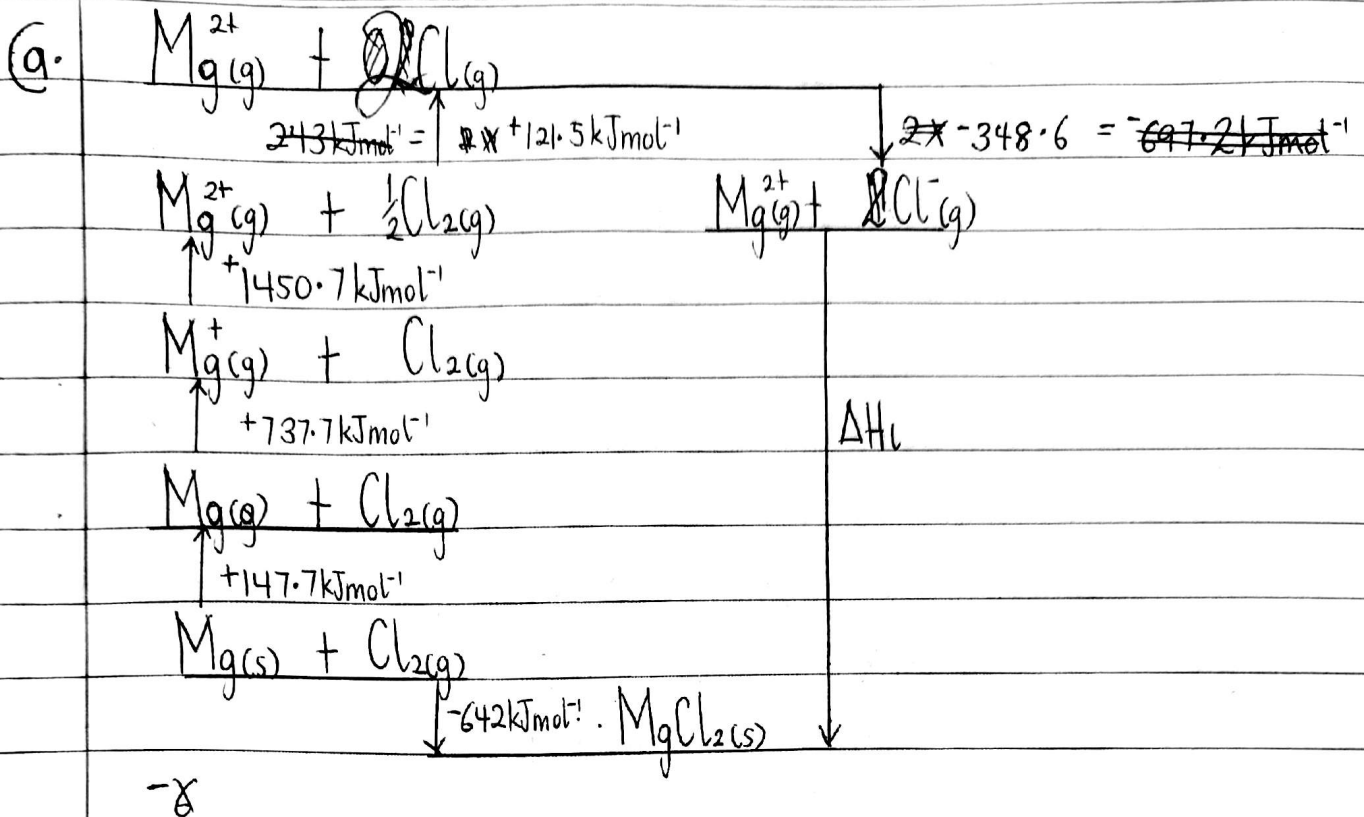
Cl⁻ Applying Hess's law

$$-642 = \Delta H_f + \overset{-348.6}{\cancel{697.2}} + \overset{121.5}{\cancel{243}} + 1450.7 + 737.7 + 147.7$$

$$-642 = \Delta H_f + \cancel{1881.9} \quad 2109$$

$$\cancel{-642 - 1881.9} = \Delta H_f \quad \cancel{-642} - 2109 = \Delta H_f$$

$$\Delta H_f = \cancel{-2523.9 \text{ kJmol}^{-1}} \quad \Delta H_f = -2751 \text{ kJmol}^{-1}$$



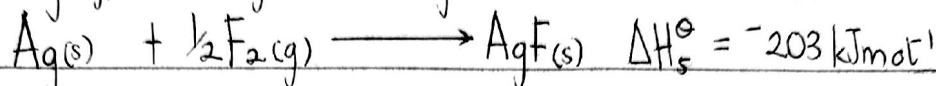
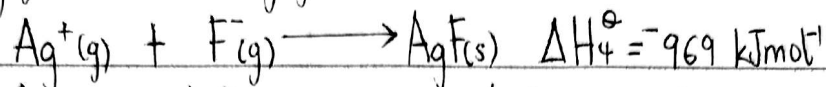
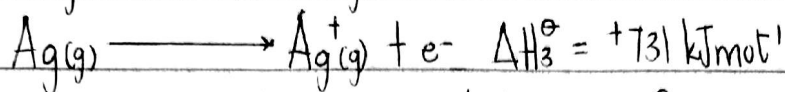
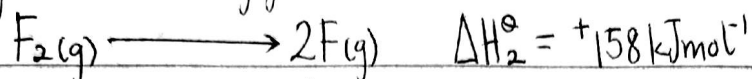
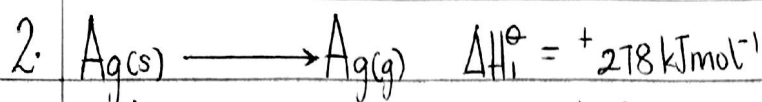
(c) $\Delta H_s = \Delta H_L + (\text{Hydration energy})$

$$\begin{aligned}
 &= +2523.9 + (-1891 + 2 \times -381) \\
 &= +2523.9 + -2653 \\
 &= -129.1 \text{ kJ mol}^{-1}
 \end{aligned}$$

(c) $\Delta H_s = \Delta H_L + \sum \text{Hydration energy}$

$$\begin{aligned}
 &= 2751 + (-891 + -381) \\
 &= 2751 + -1272 \\
 &= +1479 \text{ kJ mol}^{-1}
 \end{aligned}$$

$\therefore$ The enthalpy of solution of MgCl_2 is $+1479 \text{ kJ mol}^{-1}$. MgCl_2 dissolves endothermically therefore its solubility increases with increase in temperature.



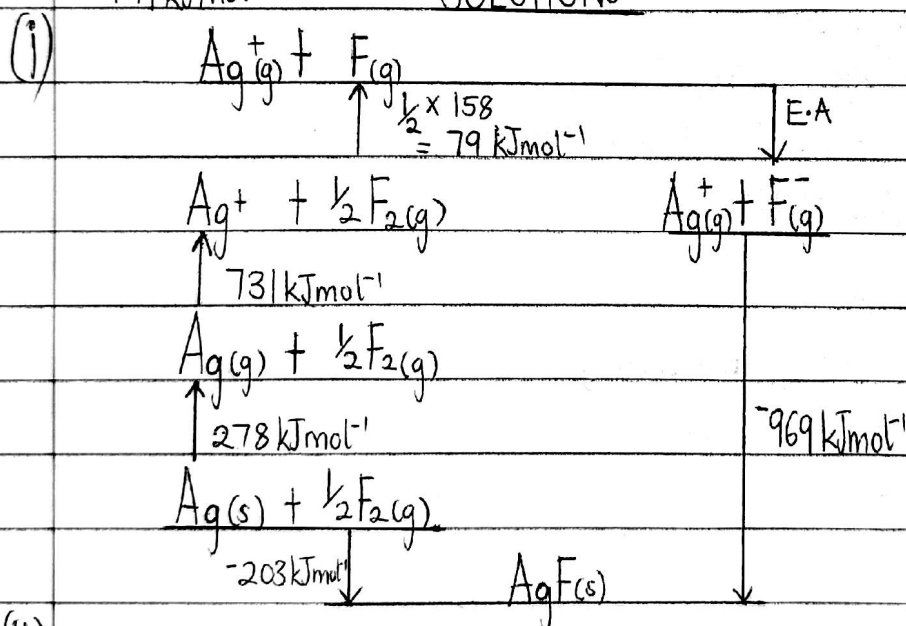
(i) Draw an energy level diagram to illustrate the energy changes for formation of silver fluoride.

(ii) Calculate the electron affinity of fluorine.

(iii) Draw a Born-Haber cycle to illustrate the relationship between the energy changes that determine the magnitude of heat solution.

(iv) Calculate the heat solution of silver fluoride if the hydration energy of silver fluoride is -991 kJmol^{-1}

SOLUTIONS



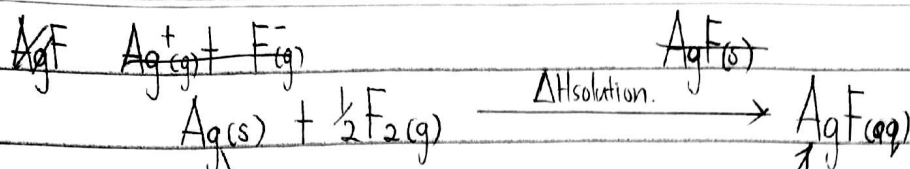
$$(ii) -203 = E.A + 79 + 731 + 278 + 969$$

$$-203 = E.A + 1088 - 969$$

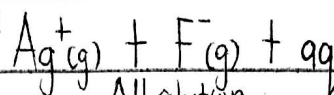
$$-203 = E.A + 119$$

$$-203 - 119 = E.A$$

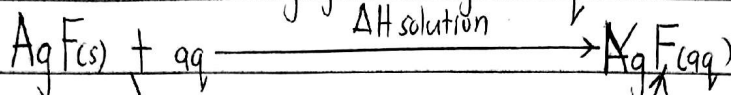
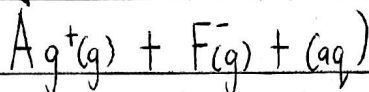
$$E.A = -322 \text{ kJmol}^{-1}$$

 ΔH_c

OMITTED

Hydration energy = -991 kJ mol^{-1} 

(iii)

 -969 kJ mol^{-1} lattice energyHydration energy = -991 kJ mol^{-1} 

(iv)

$$\Delta H_{\text{solution}} = \text{lattice energy} + \text{hydration energy}$$

$$= -969 + -991$$

$$\Delta H_{\text{solution}} = -1960 \text{ kJ mol}^{-1}$$

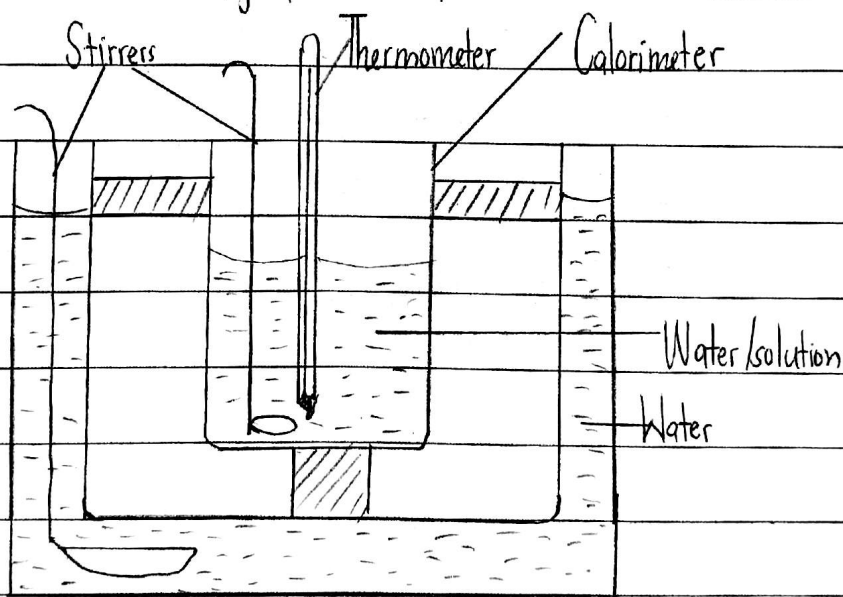
NOTE:

4/07/2025 Anhydrous salts dissolve exothermically whereas hydrated salts dissolve endothermically. This is because for anhydrous salts, their hydration energy is high since salt ~~has~~ has not given out any energy due to interaction with H_2O molecules. Hence the hydration energy will be higher than the lattice energy making the enthalpy of solution negative.

For hydrated salts, the ions are partly hydrated hence less heat given off when dissolved in water as hydration energy therefore lattice energy will outweigh the hydration energy making the enthalpy of solution positive i.e. the enthalpy of solution of hydrated Copper (II) sulphate ($CuSO_4 \cdot 5H_2O$) is endothermic whereas that of anhydrous $CuSO_4$ is exothermic.

Experiment to determine enthalpy of solution of ionic salts.

Setup:



A known volume of water $V \text{ cm}^3$ is put in a well insulated calorimeter fitted with a thermometer and a stirrer. The initial temperature of water is noted from the thermometer ($T_1^\circ\text{C}$).

A known mass ($W \text{ g}$) of finely powdered ionic salt is added to the water and the mixture is stirred thoroughly until the highest or lowest steady temperature ($T_2^\circ\text{C}$) is attained.

Treatment of results

Assumptions

- i) Heat gained by the calorimeter is neglected.
- ii) Since the solution is very dilute, its density is the same as that of water i.e. 1 g/cm^3

iii) The specific heat capacity of the solution is the same as that of water since the solution is very dilute i.e. $4200 \text{ J kg}^{-1} \text{ K}^{-1}$ or $4.2 \text{ J g}^{-1} \text{ K}^{-1}$

$$\begin{aligned} \text{Mass of solution} &= V \times 1 + W \\ &= (V + W) \text{ g} \end{aligned}$$

$$\text{Temperature change } (\Delta T) = T_2 - T_1 \text{ } ^\circ\text{C}$$

$$\begin{aligned} \text{Heat change} &= \text{Mass of solution} \times \text{Specific heat capacity} \times \text{Temperature change} \\ &= ((V + W) \times 4.2 \times \Delta T) \end{aligned}$$

If the molar mass of the salt is M_r .

$$\text{Wt of salt liberate/absorb } ((V + W) \times 4.2 \times \Delta T) \text{ J}$$

$$\text{Mrg of the salt will liberate/absorb } \left(\frac{((V + W) \times 4.2 \times \Delta T) \times M_r}{W} \right) \text{ J}$$

Therefore the enthalpy of solution of the salt is:

$$\pm \left(\frac{((V + W) \times 4.2 \times \Delta T) \times M_r}{W} \right) \text{ J mol}^{-1} / \pm \left(\frac{((V + W) \times 4.2 \times \Delta T) \times M_r}{1000 \times W} \right) \text{ kJ mol}^{-1}$$

8th July, 2025

Solution.

(i) Anhydrous CuSO_4

Heat change = Mass of solution $\times$ Specific heat capacity $\times$ Temperature change

$$= (50.14) \times 4.2 \times 8$$

$$= 1814.4 \text{ J}$$

$$\text{RMM of } \text{CuSO}_4 = 64 + 32 + 16 \times 4$$

$$= 160$$

4g of CuSO_4 liberated 1814.4 J

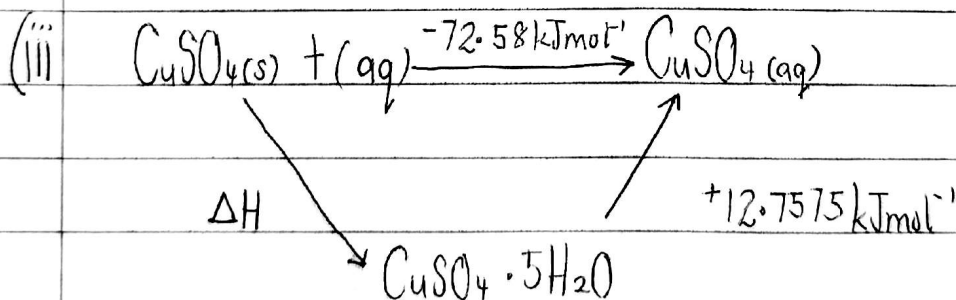
$$160 \text{ g of } \text{CuSO}_4 \text{ liberate } \left(\frac{1814.4 \times 160}{4} \right) \text{ J}$$

$$= 72576 \text{ J} \approx 72.58 \text{ kJmol}^{-1}$$

Therefore the enthalpy of solution of anhydrous CuSO_4 is $-72.58 \text{ kJmol}^{-1}$

(ii) $+12.7575 \text{ kJmol}^{-1}$

$$\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 250$$



Applying Hess's law

$$-72.58 = \Delta H + 12.7575$$

$$-72.58 - 12.7575 = \Delta H$$

$$\Delta H = -85.34 \text{ kJmol}^{-1}$$

50 cm³ of 0.5 M Sodium hydroxide and 50 cm³ of 0.25 M H₂SO₄ both at 50° 20°c were mixed and stirred in a calorimeter of negligible heat capacity. The temperature of the mixture rose to 23.2°c. Calculate the enthalpy of neutralisation of sulphuric acid.

$$\begin{aligned}\text{Heat change} &= \text{Mass of solution} \times \text{Specific heat capacity} \times \text{Temperature change} \\ &= (100 \times 1) \times 4.2 \times (23.2 - 20) \\ &= 100 \times 4.2 \times 3.2 \\ &= 1344 \text{ J} \approx 1.344 \text{ kJmol}^{-1}\end{aligned}$$

Moles of H₂SO₄.

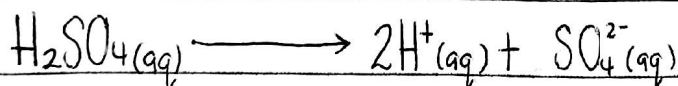
1000 cm³ of solution contain 0.5 moles of H₂SO₄

$$\begin{aligned}50 \text{ cm}^3 \text{ of solution contain } &\left(\frac{0.5 \times 50}{1000} \right) \text{ moles of H}_2\text{SO}_4 \\ &= 0.0125 \text{ moles of H}_2\text{SO}_4\end{aligned}$$

0.0125 moles of H₂SO₄ evolve 1.344 kJmol⁻¹

$$1 \text{ mole of H}_2\text{SO}_4 \text{ evolves } \left(\frac{1.344 \times 1}{0.0125} \right) \text{ kJ}$$

$$= 53.76 \text{ kJmol}^{-1} \quad 107.52 \text{ kJmol}^{-1}$$



1 mole of H₂SO₄ liberates 2 moles of H⁺

$$\begin{aligned}0.0125 \text{ moles of " liberate } &(2 \times 0.0125) \text{ moles of H}^+ \\ &= 0.025 \text{ moles of H}^+\end{aligned}$$

0.025 moles of H⁺ liberate 1344 J

$$1 \text{ mole of H}^+ \text{ " } \frac{1344 \times 1}{0.025}$$

$$= 53760 \text{ Jmol}^{-1} \approx 53.76 \text{ kJmol}^{-1}$$

Therefore the enthalpy of neutralisation is -53.76 kJmol⁻¹

3. 25cm^3 of 1M HCl at 25°C was placed in a plastic cup and 25cm^3 of 1M NaOH at was added to the acid. The mixture was stirred and the temperature rose to 28.2°C .

Calculate the enthalpy of neutralisation.

$$\text{Heat change} = \text{Mass of solution} \times \text{Specific heat capacity} \times \text{Temperature change}$$

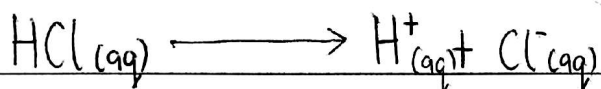
$$\begin{aligned}\text{Mass of solution} &= (25 + 25) \times 1 \\ &= 50\text{g}\end{aligned}\quad \begin{aligned}\text{Temperature change} &= 28.2 - \frac{(25 + 25)}{2} \\ &= 28.2 - 25.5\end{aligned}$$

$$\begin{aligned}\text{Heat change} &= 50 \times 4.2 \times 2.7 \\ &= 567\text{J}\end{aligned}\quad \begin{aligned}&= 2.7^\circ\text{C}\end{aligned}$$

Moles of HCl

1000cm^3 of solution contains 1 mole of HCl

$$\begin{aligned}25\text{cm}^3 \text{ of solution contains } &\left(\frac{1}{1000} \times 25\right) \text{ moles of HCl} \\ &= 0.025 \text{ moles of HCl}\end{aligned}$$



0.025 moles of HCl / H^+ liberate 567kJ

$$\begin{aligned}1 \text{ mole of HCl liberates } &\left(\frac{567 \times 0.1}{0.025}\right) \text{J} \\ &= 22680\text{Jmol}^{-1}\end{aligned}$$

The enthalpy of neutralisation is 22.68kJmol^{-1}

10/7/25 Enthalpy of precipitation.

1. Define the term enthalpy of precipitation.

2. Describe an experiment to determine enthalpy of precipitation of MgCO_3 & NaCO_3 2M

3. Calculate the enthalpy of precipitation of Magnesium nitra carbonate when 25cm^3 of MgNO_3 is added to 25cm^3 of 2M HCl and the temperature of the solution decreased by 6°C .
(density of soln = 1g/cm^3 S.H.C = $4.2\text{Jkg}^{-1}\text{K}^{-1}$)

Standard heat of displacement

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When excess zinc was added to 100cm^3 of 0.2M Cu_2SO_4 solution, the temperature of the mixture rose by 12.9°C . State what was observed.

- (i) Write an equation for the reaction.
- (ii) Calculate the enthalpy change for the reaction.

Standard enthalpy of precipitation

This is the heat change that occurs when 1 mole of a precipitate is formed from its ions in aqueous solution.

Experiment to determine heat of precipitation of magnesium carbonate from magnesium nitrate and sodium carbonate.

A known volume, $V\text{cm}^3$ of magnesium nitrate of known concentration $C\text{mol dm}^{-3}$ is measured and placed in a plastic beaker.

The initial temperature of magnesium nitrate solution, $T_1^\circ\text{C}$ is noted from the thermometer.

A equal volume, $V\text{cm}^3$ of sodium carbonate of the same concentration is also measured and poured into a separate beaker and its initial temperature $T_2^\circ\text{C}$ is noted from the thermometer.

The two solutions are added together quickly and carefully while stirring and the maximum temperature $T_3^\circ\text{C}$ is attained and noted.

Assumptions:

The solution has the same S.H.C as water = $4.2\text{ Jg}^{-1}\text{C}^{-1}$

Heat loss to the surroundings is negligible

Heat capacity of the beaker is negligible

Density of solution is 1gcm^{-3}

Treatment of results.

Initial temperature of magnesium nitrate solution = $T_1^\circ\text{C}$

Initial temperature of sodium carbonate = $T_2^\circ\text{C}$

Average initial temperature of solution = $\left(\frac{T_1 + T_2}{2}\right)^\circ\text{C}$

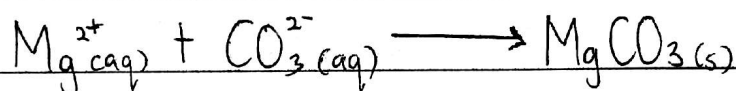
Temperature rise = $\left[T_3 - \left(\frac{T_1 + T_2}{2}\right)\right]^\circ\text{C}$

Total volume of solution = $(V + V) = 2V\text{cm}^3$

Total mass of solution = $(\text{Vol} \times \text{density}) = (2V \times 1) = 2V\text{g}$

Heat change = mass of solution $\times$ S.H.C $\times$ temperature rise
 $= (2V \times 4.2 \times \Delta T) \text{ J}$

Moles of magnesium ions that reacted = $\left(\frac{V \times C}{1000}\right)$



mole ratio $\text{Mg}^{2+} : \text{MgCO}_3 = 1:1$

$\left(\frac{V \times C}{1000}\right)$ moles of Mg^{2+} precipitate $\left(\frac{V \times C}{1000}\right)$ moles of MgCO_3

$\left(\frac{V \times C}{1000}\right)$ moles of MgCO_3 are precipitated with evolution of $(2V \times 4.2 \times \Delta T) \text{ J}$

1 mole of MgCO_3 is precipitated with $\frac{(1000 \times 2V \times 4.2 \times \Delta T)}{V \times C} \text{ J}$

Enthalpy of precipitation of $\text{MgCO}_3 = \left(\frac{1000 \times 2V \times 4.2 \times \Delta T}{V \times C}\right) \text{ Jmol}^{-1}$

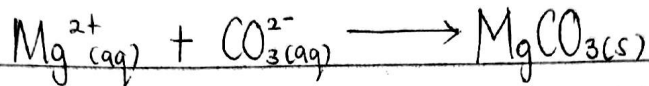
$$= \left(\frac{2V \times 4.2 \times \Delta T}{V \times C}\right) \text{ kJmol}^{-1}$$

Calculate the enthalpy of precipitation of MgCO_3 when 25cm^3 of 2M $\text{Mg}(\text{NO}_3)_2$ is added to 25cm^3 of 2M $(\text{Na})_2\text{CO}_3$ in the temp of the solution decreased by 6°C (S.H.C of solution = 4.2 , density of solution = 1g/cm^3)

Solution $\Delta T = 6^\circ\text{C}$ $S.H.C = 4.2$

Total mass of solution = $(25 + 25) = 50$ Total mass = $50 \times 1 = 50\text{g}$.

$$\begin{aligned}\text{Heat change} &= mC\Delta T \\ &= 50 \times 4.2 \times 6 \\ &= 1260\text{J}\end{aligned}$$



1000cm^3 of solution contain 2 moles of Mg^{2+}

25cm^3 of solution contain $\left(\frac{2}{1000} \times 25\right)$ moles of Mg^{2+}
 $= 0.05$ moles of Mg^{2+}

1 mole of Mg^{2+} ions precipitates 1 mole of MgCO_3 .

0.05 moles of Mg^{2+} ions precipitate $\left(\frac{1}{1} \times 0.05\right)$ moles of MgCO_3
 $= 0.05$ moles of MgCO_3

0.05 moles of MgCO_3 were precipitated with evolution of 1260J

1 mole of MgCO_3 is " " " " $\left(\frac{1260 \times 1}{0.05}\right) = 25,200\text{J}$

Enthalpy of precipitation of $\text{MgCO}_3 = -25,200\text{Jmol}^{-1}$
 $= -25.2\text{kJmol}^{-1}$

Enthalpy of displacement (standard)

This is the heat evolved when 1 mole of ions of a less electropositive element is displaced from its solution by a more electropositive element at standard conditions.

It's denoted by (-)ve sign

Experiment to determine enthalpy of displacement between zinc powder and copper (II) sulphate solutions-

A known volume ($V\text{cm}^3$) of copper (II) sulphate solution of a known

concentration (cmol dm^{-3}) is put in a plastic beaker.

The initial temperature $T_1^\circ\text{C}$ of the solution is measured and recorded from the thermometer.

An excess of zinc powder is added quickly and carefully to solution while stirring. The final steady temperature, $T_2^\circ\text{C}$ is read and recorded from the thermometer.

Assumptions

Density of the solution is equal to that of water $= 1.0 \text{ g cm}^{-3}$

S.H.C of solution is equal to that of water $= 4.2 \text{ J g}^{-1} \text{ K}^{-1}$

No heat is lost to the surrounding.

Heat capacity of the beaker is negligible.

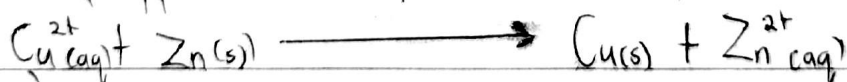
Treatment of results.

$$\text{Temperature rise} = \Delta T = (T_2 - T_1)^\circ\text{C}$$

$$\text{Mass of solution} = (\text{Volume} \times \text{density}) = (V \times 1) = V \text{ g}$$

$$\begin{aligned} \text{Heat change} &= \text{mass of solution} \times \text{S.H.C} \times \text{temperature rise} \\ &= (V \times 4.2 \times (T_2 - T_1)) \text{ J} \end{aligned}$$

$$\text{Moles of copper (II) ions displaced} = \left(\frac{V \times C}{1000} \right)$$



$\left(\frac{V \times C}{1000} \right)$ moles of Cu^{2+} are displaced with evolution of $(V \times 4.2 \times \Delta T) \text{ J}$

$$\begin{aligned} \text{1 mole of } \text{Cu}^{2+} \text{ is displaced with evolution of } & - \left(\frac{V \times 4.2 \times \Delta T}{\frac{V \times C}{1000}} \right) \text{ J mol}^{-1} \\ & = - \left(\frac{V \times 4.2 \times \Delta T}{V \times C} \right) \text{ kJ mol}^{-1} \end{aligned}$$

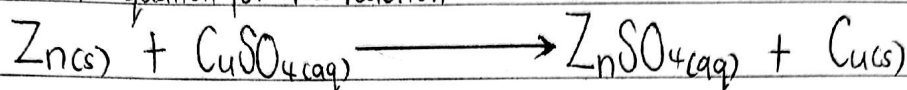
When excess Zinc was added to 100 cm^3 of 0.25 M CuSO_4 solution, the

temperature of the mixture rose by 12.9°C .

(i) State what was observed

A brown solid is seen at the bottom of the container.

(ii) Write an equation for the reaction



(iii) Calculate the ΔH for the reaction.

$$S.H.C = 4.2 \quad \Delta T = 12.9^\circ\text{C}$$

$$\begin{aligned} \text{Total mass} &= \text{Vol} \times \text{density} \\ &= (100 \times 1) = 100\text{g} \end{aligned}$$

1000 cm³ of soln contain 0.25 moles

$$100\text{cm}^3 \text{ " " " } \left(\frac{0.25}{1000} \times 100 \right)$$

$$\text{Heat change} = m c \Delta T$$

$$= 0.025 \text{ moles}$$

$$= 100 \times 4.2 \times 12.9$$

$$= 5418 \text{ J}$$

0.025 moles of Cu^{2+} are displaced with evolution of 5418 J

$$1 \text{ mole " " " " " " " } \left(\frac{5418 \times 1}{0.025} \right)$$

$$= 216,720 \text{ J mol}^{-1}$$

$$= 216.72 \text{ kJ mol}^{-1}$$

25th September, 2025

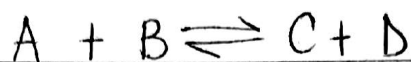
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CHEMICAL EQUILIBRIA

Chemical equilibrium is concerned with reversible reactions. A reversible reaction is a reaction which does not go to completion i.e. not all reactants are converted to products and the reaction proceeds in both the forward and backward directions.

Consider a reversible reaction:



In the forward reaction A & B react to give C & D as products. In the backward reaction, C & D combine to give A & B as products. The forward and backward reactions proceed until a point is reached where their rates become equal and at this point equilibrium is attained and this equilibrium is dynamic in nature.

$$\text{Rate of reaction} \propto [\text{Reactants}]$$

$$\text{Rate of reaction} = K [\text{Reactants}]$$

Where K is the rate constant. If K_1 and K_2 are the rate constants for the forward and backward reactions respectively.

$$\text{Rate of forward reaction} = K_1 [A][B]$$

$$\text{Rate of backward reaction} = K_2 [C][D]$$

At equilibrium rate of forward reaction = rate of backward reaction.

$$\text{i.e. } K_1 [A][B] = K_2 [C][D]$$

$$\frac{K_1}{K_2} = \frac{[A][B]}{[C][D]} \quad \text{But } \frac{K_1}{K_2} = K_c \text{ (equilibrium constant)}$$

$$\frac{K_1}{K_2} = \frac{[C][D]}{[A][B]} \quad K_c = \frac{[C][D]}{[A][B]}$$

Equilibrium constant is defined as the ratio of ^{Product of} molar concentration of products to the product of molar concentration of reactants all raised to their appropriate powers from a stoichiometric (well balanced) equation of a reversible reaction at

Stud