**The diagram shows the distribution of energy for the molecules in a sample of gas at a given temperature, T1**

**(a) In the diagram *E*a represents the *activation energy* for a reaction. Define this term.**

**(b) On the diagram above draw another curve to show the energy distribution for the same gas at a higher temperature. Label the curve T2.**

**(c) With reference to your diagram, state and explain what happens to the rate of a reaction when the temperature is increased.**

a.: Activation energy: Minimum energy required for reaction to occur.

B: Draw a curve moved to the right, with a lower peak.

C: The rate is increased, more particles with more energy

**(a) Define the term *average bond enthalpy*, illustrating your answer with an equation for methane, CH4.**

**(b) The equation for the reaction between methane and chlorine is**

**CH4(g) + Cl2(g) → CH3Cl(g) + HCl(g)**

**Use the values from Table 10 of the Data Booklet to calculate the enthalpy change for this reaction.**

**(c) Explain why no reaction takes place between methane and chlorine at room   
temperature unless the reactants are sparked, exposed to UV light or heated.**

**(d) Draw an enthalpy level diagram for this reaction.**

A: Energy for the conversion of a gaseous molecule into gaseous atoms. Average value obtained from a number of similar compounds. CH4(g) -> C(g)+4H(g)

B: Bond Break: 1890, 654. Bond make: 2005, 769. Answer: -115 kJ mol-1

C: Molecules have insufficient energy to react at room temp, lots of unsuccessful collisions, extra energy to overcome activation energy.

D: Draw a exothermic diagram, label it.

**In aqueous solution, potassium hydroxide and hydrochloric acid react as follows.**

**KOH(aq) + HCl(aq) → KCl(aq)+ H2O(l)**

**The data below is from an experiment to determine the enthalpy change of this reaction.**

**50.0 cm3 of a 0.500 mol dm–3 solution of KOH was mixed rapidly in a glass beaker with 50.0 cm3 of a 0.500 mol dm–3 solution of HCl.**

**Initial temperature of each solution = 19.6°C  
Final temperature of the mixture = 23.1°C**

**(a) State, with a reason, whether the reaction is exothermic or endothermic.**

**(b) Explain why the solutions were mixed rapidly.**

**(c) Calculate the enthalpy change of this reaction in kJ mol–1. Assume that the specific heat capacity of the solution is the same as that of water.**

**(d) Identify the major source of error in the experimental procedure described above. Explain how it could be minimized.**

**(e) The experiment was repeated but with an HCl concentration of 0.510 mol dm–3 instead of 0.500 mol dm–3. State and explain what the temperature change would be.**

a. It is exothermic because temperature rises

b. To prevent heat loss, and so all the heat will be given out quickly

c. 100×4.18×3.5 (q = *mc*∆*T* =) 1463J. Amount of KOH used: .5\*.5= .025 mol ∆*H* = (1.463÷0.025) = –58.5 (kJ mol–1)

d. Heat lost to surroundings. Insulate the vessel, use a lid

e. 3.5 (The same) because amount of base reacted would be the same

**The data below is from an experiment used to measure the enthalpy change for the combustion of 1 mole of sucrose (common table sugar), C12H22O11(s). The time-temperature data was taken from a data-logging software programme.**

(a) Calculate Δ*T*, for the water, surrounding the chamber in the calorimeter.

(b) Determine the amount, in moles, of sucrose.

(c) (i) Calculate the enthalpy change for the combustion of 1 mole of sucrose.

(ii) Using Table 12 of the Data Booklet, calculate the percentage experimental error based on the data used in this experiment.

(d) A hypothesis is suggested that TNT, 2-methyl-1,3,5-trinitrobenzene, is a powerful explosive because it has:

a. Δ*T* = 23.70 – 23.03 = 0.67 (°C/K)

b.  = 1.281×10–3;

c. (i) Δ*H*c = (*C* Δ*T*)/*n* =  = –5.3×103 kJ mol–1

(i) Percentage experimental error == 5.4%

d.: Enthalpy change of combustion of sucrose is bigger than TNT and not important, Rate of reaction for TNT is greater than sucrose, so it is valid. Amount of gas gernated in sucrose is grater than TNT, so not important,

**(a) Define the term *average bond enthalpy*.**

**(b) Use the information from Table 10 of the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene, according to the following equation.**

**C4H8(g) + 6O2(g) → 4CO2(g) + 4H2O(g)**

1. The amount of energy to break 1 mole of covalent bonds in the gaseous state. Average calculated from a range of compounds.
2. Bonds broken: (612) + (2×348) + (8×412) + (6×496)/7580 (kJ mol–1). Bonds made: (8×743) + (8×463) / 9648 (kJ mol–1); *H* = –2068 (kJ mol–1);

**Given the following data:**

**C(s) + F2(g) → CF4(g); ∆*H*1= –680 kJ mol–1  
F2(g) → 2F(g); ∆*H*2= +158 kJ mol–1  
C(s) → C(g); ∆*H*3= +715 kJ mol–1**

**calculate the average bond enthalpy (in kJ mol–1) for the C––F bond.**

C(s) + 2F2(g) → CF4(g) ∆*H*1 = –680 kJ;  
4F(g) → 2F2(g) ∆*H*2 = 2(–158) kJ;  
C(g) → C(s) ∆*H*3 = –715 kJ;

C(g) + 4F(g) → CF4(g) ∆*H* = –1711 kJ,

so average bond enthalpy = 

= –428 kJ mol–1

Two reactions occurring in the manufacture of sulfuric acid are shown below:

reaction I S(s) +O2(g)  SO2(g) *H*Ө = –297 kJ

reaction II SO2(g) + O2(g) SO3(g) *H*Ө = –92 kJ

(i) State the name of the term *H*Ө. State, with a reason, whether reaction I would be accompanied by a decrease or increase in temperature.

(ii) At room temperature sulfur trioxide, SO3, is a solid. Deduce, with a reason, whether the *H*Ө value would be more negative or less negative if SO3(s) instead of SO3(g) were formed in reaction II.

(iii) Deduce the *H*Ө value of this reaction:

S(s) + O2(g)  SO3(g)

(i) standard enthalpy (change) of reaction;

(temperature) increase;

reaction is exothermic/sign of *H* is negative

(ii)more (negative);

heat given out when gas changes to solid/solid has less enthalpy than  
gas

(iii) –389 kJ

**(i) Define the term *average bond enthalpy*.**

**(ii) Explain why Br2 is not suitable as an example to illustrate the term *average bond enthalpy*.**

**(iii) Using values from Table 10 of the Data Booklet, calculate the enthalpy change for the following reaction: CH4(g) + Br2(g)  CH3Br(g) + HBr(g)**

**(iv) Sketch an enthalpy level diagram for the reaction in part (iii).**

**(v) Without carrying out a calculation, suggest, with a reason, how the enthalpy change for the following reaction compares with that of the reaction in part (iii):**

**CH3Br(g) + Br2(g)  CH2Br2(g) + HBr(g)**

(i) the energy needed to break one bond;  
(in a molecule in the) gaseous state;  
value averaged using those from similar compounds; 3

(ii) it is an element/no other species with just a Br-Br bond/*OWTTE;* 1

(iii) (sum bonds broken =) 412 + 193 = 605;

(sum bonds formed =) 276 + 366 = 642;

(*H* =) –37 kJ

(iv)

|  |  |  |
| --- | --- | --- |
| Enthalpy | CH4 + Br2  CH3Br + HBr |  |

(v) Aboyut the same, same number and type of bonds being formed and broken.

But–1–ene gas, burns in oxygen to produce carbon dioxide and water vapour according to the following equation. C4H8 + 6O2  4CO2 + 4H2O

(a) Use the data below to calculate the value of *H*Ө for the combustion of but-1-ene.

(b) State and explain whether the reaction above is endothermic or exothermic.

(a) (Amount of energy required to break bonds of reactants)

8×412 + 2×348 + 612 + 6×496/7580 (kJ mol1)

(Amount of energy released during bond formation)

4×2×743 + 4×2×463/9648 (kJ mol1)

*H* = 2068 (kJ or kJ mol1)

(b) exothermic **and** *HӨ* is negative/energy is released

**Calculate the enthalpy change, *H*4 for the reaction**

**C + 2H2 + O2  CH3OH *H*4**

**using Hess’s Law and the following information.**

**CH3OH + O2  CO2 + 2H2O *H*1 = 676 kJ mol1**

**C + O2  CO2 *H*2 = 394 kJ mol1**

**H2 + O2  H2O *H*3 = 242 kJ mol1**

1×H1/676;

1×H2/–394;

2×H3/– 484;

*H*4 = 202 (kJ mol1 )

**Methylamine can be manufactured by the following reaction.**

**CH3OH(g) + NH3(g)  CH3NH2(g) + H2O(g)**

**(a) Define the term *average bond enthalpy*.**

**(b) Use information from Table 10 of the Data Booklet to calculate the enthalpy change for this reaction.**

(a) energy needed to break (1 mol of) a bond in a gaseous molecule;

averaged over similar compounds; 2

(b) bonds broken identified as CO and NH;

bonds formed identified as CN and OH;

*H* = 748  768 (kJ);

=  20 kJ/kJ mol1

**(a) Define the term *average bond enthalpy*.**

**(b) Use the information from Table 10 in the Data Booklet to calculate the enthalpy change for the complete combustion of but-1-ene according to the following equation**

**C4H8(g)  4CO2(g) + 4H2O(g)**

**(c) Predict, giving a reason, how the enthalpy change for the complete combustion of but-2-ene would compare with that of but-1-ene based on average bond enthalpies.**

**(d) The enthalpy level diagram for a certain reaction is shown below.**

(a) amount of energy needed to break one mole of (covalent) bonds;

in the gaseous state;

average calculated from a range of compounds; 2

Award **[1]** each for any two points above.

(b) bonds broken: 161 + 2×348 + 8×412 + 6×496/7580 kJ mol1;

bonds made: 8×743 + 8×463/9648 kJ mol1;

(bonds broken  bonds made =) *H* = 2068(kJ mol1)

(c) same/equal, because the same bonds are being broken and formed; 1

(d) products more stable than reactants;

bonds are stronger in products than reactants/*H*P *< H*R/enthalpy/stored  
energy of products less than reactants

**The reaction between ethene and hydrogen gas is exothermic.**

**(i) Write an equation for this reaction.**

**(ii) Deduce the relative stabilities and energies of the reactants and products.**

**(iii) Explain, by referring to the bonds in the molecules, why the reaction is exothermic.**

(i) C2H4(g) + H2(g)  C2H6(g)

State symbols not required for mark

(ii) products more stable than reactants/reactants less stable than products;

products lower in energy/reactants higher in energy

(iii) (overall) bonds in reactants weaker/(overall) bonds in product stronger  
/all bonds in product are  bonds/weaker  bond broken and a  
(stronger)  bond formed

less energy needed to break weaker bonds/more energy produced  
to make stronger bonds (thus reaction is exothermic)

**(i) Define the term *average bond enthalpy*.**

**(ii) The equation for the reaction of ethyne and hydrogen is:**

**C2H2(g) + 2H2(g)  C2H6(g)**

**Use information from Table 10 of the Data Booklet to calculate the change in enthalpy for the reaction.**

**(iii) State and explain the trend in the bond enthalpies of the C–Cl, C–Br and C–I bonds.**

(i) energy required to break (a mole of) bonds in the gaseous state  
/energy given out when (a mole of) bonds are made in the  
gaseous state  
average value from a number of similar compounds

(ii) (*HӨ*reaction = (∑*BE*break  *BE*make))

= [(837) + 2(436)]  [(348 + 4(412)]

=  287(kJ/kJ mol1)

Award **[1 max]** for 287 or + 287.

(iii) (BE): CCl > CBr > CI/CX bond becomes weaker

halogen size/radius increases/bonding electrons further away from  
the nucleus/bonds become longer