**Name:**

**Periodic Table & Energy**

**End of Module Test**

|  |  |
| --- | --- |
| **Question** | **Marks** |
| **1** | **/ 3** |
| **2** | **/ 4** |
| **3** | **/ 6** |
| **4** | **/ 3** |
| **5** | **/10** |
| **6** | **/ 8** |
| **7** | **/ 7** |
| **8** | **/11** |
| **Total** | **/52** |
| **Grade** |  |

1. Explain the following: Sodium atoms are larger than magnesium atoms.

 **[Total: 3]**

1. The flowchart below shows some reactions involving barium. Write the **formulae** of substances **A**–**D** in the boxes.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **Ba(s)** | heatair | ……………………..(s)**solid A** |  |  |
|  |  |  H2O(l) |  |  |
|  |  | **Ba(OH)2 (aq)** | HC*l*(aq) | ……………………..(aq)**Solution** **B** |
|  |  |  CO2 (g) |  |  AgNO3 (aq) |
|  |  | ……………………..(s)**solid C** |  | ……………………..(s)**solid D** |

# [Total: 4]

1. A student carries out two experiments.
	1. The student bubbles some chlorine gas through a solution of sodium iodide. The solution turns a brown colour.

Explain this observation and write an equation for the reaction that takes place.

 **[2]**

* 1. The student bubbles chlorine gas through aqueous silver nitrate. A white precipitate forms. Explain this observation including equations for any reactions that take place.

 **[4]**

**[Total: 6]**

1. The Group 2 elements become more reactive down the group.

The Group 7 elements become **less** reactive down the group.

Explain this difference between Group 2 and Group 7.

 **[Total: 3]**

1. Glucose, C6H12O6, can be used directly as a source of energy in living species or fermented to produce ethanol, C2H5OH. The ethanol produced can then be used as a fuel.
	1. Energy is released from the oxidation of glucose in living species.

What name is given to this process?

 **[1]**

* 1. When ethanol is used as a fuel, a combustion reaction takes place. The equation for this process is shown below.

C2H5OH (l) + 3 O2 (g) → 2 CO2 (g) + 3 H2O Δ*H* = –1367 kJ mol–1

The table belowshows values for standard enthalpy changes of formation, Δ*H*~~o~~f..

|  |  |
| --- | --- |
| compound | Δ*H*~~o~~f (kJ mol–1) |
| CO2(g) | –394 |
| H2O(l) | –286 |

* + 1. Define the term *standard enthalpy change of formation*.

 **[3]**

* + 1. Calculate the standard enthalpy change of formation of ethanol.

Δ*H*~~o~~f = ............................................ kJ mol−1 **[3]**

* + 1. On the axes below draw the enthalpy profile diagram for the combustion of ethanol.

Label *Ea* and Δ*H* on your diagram.

C2H5OH(I) + 3O2(g)

 enthalpy

**[3]**

 progress of reaction **[Total: 10]**

1. Reactions that release energy are essential to maintain many aspects of everyday life.
	1. Enthalpy changes can be calculated using average bond enthalpies. **Table 1.1** shows some average bond enthalpies.

|  |  |
| --- | --- |
| Bond | Average bond enthalpy (kJ mol–1) |
| C–C | +347 |
| C–H | +413 |
| O–H | +464 |
| O=O | +498 |
| C=O | +805 |
| C–O | +358 |

#  Table 1.1

* + 1. Define the term *bond enthalpy*.

 **[2]**

* + 1. Ethanol, C2H5OH, can be burnt to release energy. The equation below shows the combustion of ethanol.

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| C:\Users\lfrancis.SEXEYS.000\Desktop\Ethanol.png | + | 3 O=O | → | 2 O=C=O | + | 3 H–O–H |

Use the data in **Table 1.1** to calculate a value for the enthalpy change of combustion of gaseous ethanol, Δ*H*c.

Δ*H*c = ................... kJ mol−1 **[3]**

* 1. Glucose, C6H12O6, can also be oxidised.

The table below shows some enthalpy changes of formation, Δ*H*f.

|  |  |
| --- | --- |
| compound | Δ*Hf* (kJ mol–1) |
| C6H12O6(s) | –1273 |
| CO2(g) | –394 |
| H2O(l) | –286 |

Use the data to calculate the enthalpy change of combustion, Δ*H*c, for glucose. C6H12O6(s) + 6 O2(g) → 6 CO2(g) + 6 H2O(l)

Δ*H*c = ................... kJ mol−1 **[3]**

 **[Total: 8]**

1. This question is concerned with some of the graphs you may have seen in your studies.
	1. **Graph A** shows a Boltzmann distribution.

# Graph A

* + 1. Label the axes on **graph A**. **[2]**
		2. Explain, using **graph A**, the effect of adding a catalyst on the reaction rate.

Include labelled lines of the energies involved on **graph A**.

 **[2]**

* 1. **Graph B** shows the change in mass observed in an experiment used to investigate the rate of a reaction.

**Graph B**



cotton wool

mass

zinc + hydrochloric acid

 g balance

time

* + 1. Suggest what is happening in this experiment to produce the loss of mass with time.

 **[1]**

* + 1. Exactly the same experiment was repeated but with a catalyst added.

On **graph B** sketch the line that would be produced in the presence of the catalyst. **[2]**

**[Total: 7]**

1. Many chemical reactions are reversible and are able to form equilibrium mixtures.

Hydrogen and iodine react together to form hydrogen iodide in a reversible reaction.

H2 (g) + I2 (g) ⇌ 2 HI (g) Δ*H*= +53 kJ mol-1

* 1. Explain the following observations when changes are made to an equilibrium mixture of H2 (g), I2 (g) and HI (g).

Include reference to the equilibrium position and any other factors.

* + 1. When the temperature is increased the purple colour becomes paler.

 **[2]**

* + 1. When the pressure is increased the purple colour becomes deeper.

 **[3]**

* 1. The **rate** at which equilibrium is reached could be increased by increasing the temperature or pressure.

In each case explain why the rate increases:

* + 1. on increasing the temperature,

 **[2]**

* + 1. on increasing the pressure.

 **[1]**

* 1. H2(g), I2 (g) and HI(g) were mixed together and allowed to reach equilibrium. The concentrations of the gases were then measured at various times and the results plotted. At time *t*, a change was made to the composition of the mixture.

concentration

H2(g)

HI(g)

I (g)

2

*t* time

* + 1. What change was made to the mixture at time *t* ?

 **[1]**

* + 1. Explain the changes that happen to the equilibrium mixture after time *t*.

 **[2]**

**[Total: 11]**