**RATU NAVULA COLLEGE**

**WEEK 2 YEAR 12B : CHEMISTRY NOTES & ACTIVITIES**

**LESSON NUMBER: 39**

**STRAND: 3. REACTIONS**

**SUB-STRAND: 3.3 PHYSICAL CHEMISTRY**

**LEARNING OUTCOME:**

* **Factors that affect the rate of reaction.**
* **Collision theory.**

**FACTORS WHICH AFFECT REACTION RATE:**

**The methods for speeding up a chemical reaction are:**

1. **increasing the concentration of the reacting chemicals**
2. **raising the temperature**
3. **increasing the surface area of the reactants**
4. **increasing the pressure**
5. **using a catalyst**
6. **Concentration**

**An increase in concentration of a solution increases the number of particles per unit volume. The chance of collision therefore increases and the reaction is faster.**

1. **Temperature**

**A rise in temperature by applying heat increases the speed of the particles therefore their energy. There are more collisions per second and more colliding particles have the required activation energy for the reaction.**

1. **Surface Area**

**The larger the surface area, the number of collisions per second is increased, therefore, the reaction is faster.**

**Note: The surface area of a solid can be increased by turning the solid into a powder or into finely divided particles. These react more quickly than the same mass in large lumps.**

1. **Pressure (applies only to gaseous reactants)**

**An increase in pressure of a gaseous reactant increases the number of particles per unit volume. The increase in particles increases the chances of effective collisions and as a result the reaction is faster.**

**Note: V↑P↓ and P↑V↓**

1. **Catalysts**

**Role: Catalysts lower the activation energy and therefore increase the number of effective collisions per second and the rate of reaction.**

* **most catalysts used in the industry are solids**
* **surface catalysts must be finely divided in order to have a large surface area so that particles of one reactant can attach themselves weakly to the catalysts surface. (ADSORPTION)**

**A CATALYST CAN BE DEFINED AS A SUBSTANCE WHICH ALTERS THE RATE OF REACTION (SPEEDS IT UP OR SLOWS IT DOWN) WITHOUT ITSELF UNDERGOING ANY PERMANENT CHEMICAL CHANGE**

UNCATALYSED REACTION

ENERGY CONTENT (kJ)

CATALYSED REACTION

N2 + 3H2

400

200

- 200

0

600

HEAT OF REACTION

2NH3

ACTIVATION ENERGY FOR **CATALYSED** REACTION

ACTIVATION ENERGY FOR **UNCATALYSED** REACTION

PROGRESS OF REACTION

\***Note: In the presence of the catalyst, a greater proportion of the molecules have sufficient energy to overcome the lower activation energy barrier and therefore the reaction goes much faster. The catalyst, in fact, provides a new reaction path for the breaking and rearrangement of bonds with a lower activation energy.**

**COLLISION THEORY OF REACTION**

* **States that before two or more substances can react together their particles must collide. Therefore, reactions between solids where particles vibrate, but do not move from place to place, do not occur.**
* **Another important factor in effective collision of reacting particles is the orientation of the particles. If the molecules are not properly orientated the particles will not react.**

**ACTIVATION ENERGY**

* **is the certain minimum amount of energy required to enable chemical bonds to break and a reaction to occur.**

**Exercise 6:**

1. What is the role of a catalyst in a chemical reaction?
2. Define ‘Activation Energy’

**LESSON NUMBER: 40**

**STRAND: 3. REACTIONS**

**SUB-STRAND: 3.3 PHYSICAL CHEMISTRY**

**LEARNING OUTCOME:**

* **Distinguish between exothermic and endothermic reactions.**

**HEAT CHANGES; EXOTHERMIC AND ENDOTHERMIC REACTIONS**

* **Exothermic Reaction**
  + **Heat is evolved (there is a rise in temperature)**
  + **The container becomes warm**
  + **ΔH is negative**
  + **Heat of product < Heat of reactants**
* **Endothermic Reaction**
  + **Heat is absorbed**
  + **Container feels cooler**
  + **ΔH is positive**
  + **Heat of product > Heat of reactants**

**Examples of Exothermic Reactions**

1. **When a concentrated acid is diluted by adding to water slowly, heat is always evolved. (Rise in temperature is noted)**

**H2SO4 (l) + (aq) 🡪 H2SO4 (aq) [ΔH = -70 kJ/mol]**

**Heat of solution for one mole of sulphuric acid in one litre of solution is -70 kJ/mol**

1. **Neutralisation reactions e.g. the reaction between hydrochloric acid and sodium hydroxide solution are accompanied by the release of heat.**

**Examples of Endothermic Reactions**

**NH4NO3(s) + (aq) 🡪 NH4NO3 (aq) [ΔH = +25 kJ/mol]**

**When one mole (80g) of ammonium nitrate is dissolved in 920ml of water, 1mol L-1 solution of ammonium nitrate is obtained. (a fall in temperature is recorded)**

**🡪 Another salt that loses heat when dissolved in water is ammonium chloride.**

1. **Graph showing energy changes for an exothermic reaction**

ENTHALPY

ACTIVATION ENERGY

ΔH

ΔH is negative; heat is released

🡪 products contain less energy than the reactants

INITIAL ENTHALPY OF REACTANTS

FINAL ENTHALPY OF PRODUCTS

TIME

1. **Graph showing the energy changes for an endothermic reaction:**

ENTHALPY

ΔH is positive; heat is absorbed

🡪 products contain more energy than the reactants

ACTIVATION ENERGY

FINAL ENTHALPY OF PRODUCTS

ΔH

INITIAL ENTHALPY OF REACTANTS

TIME

**Exercise 7**

1. **Copy the following diagram in the exercise book and identify A,B,C and D**

ENTHALPY

C

D

A

B

TIME

**2.Give an example of an endothermic reaction**

1. **0.5g of magnesium ribbon is added to 25ml of a 0.2 mol/L hydrochloric acid solution.**
   1. **write a balanced equation for the reaction**
   2. **state how the initial rate of reaction would change if in another experiment 0.5 g of Mg ribbon was added to 25 ml of a 0.5 mol/L hydrochloric acid solution. Give a reason.**
   3. **If 0.5 g of magnesium powder has been used instead of the ribbon in the first experiment how would the initial rate of reaction change? Give you answer in terms of simple collision theory.**

**LESSON NUMBER: 41**

**STRAND: 3. REACTIONS**

**SUB-STRAND: 3.3 PHYSICAL CHEMISTRY**

**LEARNING OUTCOME:**

* **Bond breaking and bond formation.**

**The general rule is:**

* **Breaking chemical bonds need energy.**
* **Forming chemical bonds releases energy .**

* **The symbol H is used to represent the enthalpy (heat content ) of a chemical.**
* **The enthalpy of a chemical cannot be measured by itself: however, the change in enthalpy (∆H) during a reaction can be measured.**

**\*Note - ΔH = H products – H reactants**

**Unit for ΔH = kJ/mol (kJmol-1, kilojoules per mole)**

**ΔH = enthalpy change/heat of reaction**

**H = heat content/enthalpy**

**Examples**

**1. C(s) + 2O (g) 🡪 CO2 (g) [ΔH = -390 kJ/mol]**

**The above equation means that 390 kilojoules of energy are evolved when a mole of carbon dioxide is formed.**

**2. C(s) + 2S (l) 🡪 CS2 (l) [ΔH = +106 kJ/mol]**

**This means that 106 kilojoules of heat energy are absorbed when 12g of carbon (one mole) combine with 64g of Sulphur (one mole) to form one mole of carbon disulphide.**

* ***Heat of Combustion*: for a combustion reaction**

**C(s) + O2 (g) 🡪 CO2 (g)**

* ***Heat of Solution* : of a solute is the heat change when one mole of the solute dissolves in a large volume of water**
* ***Heat of Neutralisation*: is the heat change when an alkali which provides one mole of OH-(aq) reacts with an acid which provides one mole of H+(aq)**
* ***Heat of Formation*: change in enthalpy when one mole of a compound is produced from the free elements**

**Exercise 8**

* + 1. **Consider the following reaction between hydrochloric acid and sodium hydroxide:**

**NaOH(s) + HCl(aq) 🡪 NaCl(aq) + H2O(l) [ΔH = -60 kJ/mol]**

**a) is the reaction endothermic of exothermic? Give a reason.**

**b) briefly explain why this reaction can be described as neutralization**

**2. To investigate the rate of a chemical reaction, a student used the reaction between a lump of calcium carbonate and hydrochloric acid:**

**CaCO3(s) + 2HCl(aq) 🡪 CaCl­2(aq) + H2O(l) + CO2(g)**

**Three experiments were carried out at constant temperature using the quantities and concentrations as set out in the table below:**

|  |  |  |  |
| --- | --- | --- | --- |
| **Experiment #** | **Mass of CaCO3 (g)** | **Volume of HCl (mL)** | **Concentration of HCl (mol L-1)** |
| **1** | **5** | **20** | **0.1** |
| **2** | **5** | **20** | **1.0** |
| **3** | **5** | **20** | **2.0** |

* 1. **In which reaction would the reaction rate be the fastest? Give a reason**
  2. **Which experiment would produce the smallest volume of carbon dioxide? Give a reason**
  3. **State two other methods that could be used to increase the rate of this reaction**

**LESSON NUMBER: 42**

**STRAND: 3. REACTIONS**

**SUB-STRAND: 3.3 PHYSICAL CHEMISTRY**

**LEARNING OUTCOME:**

* **Calculating enthalpy changes of reactions in Exothermic and Endothermic reactions**

**Enthalpy Change (∆H)**

* **Is the energy change occurring during a reaction.**
* **This unit of measurements is useful for calculating the amount of energy per mole either released or produced in a reaction.**

|  |
| --- |
| **∆H = Hp –Hr**  **Note:**  **If the ∆H is negative, the reaction is exothermic, heat is released.**  **If the ∆H positive, the reaction is endothermic, heat energy is absorbed.** |

* **The amount of heat energy being absorbed or released depends on the strength of the bonds being broken or formed respectively.**

**Example:**

1. **CH4(g) + 2O2(g) CO2(g) + 2H2O(g) ∆H= -895 kJ/mol.**

* **The equation shows that 895kJ/mol of heat energy is given out when methane gas undergoes combustion.**

1. **S(s) + O2 SO2 ∆H= -296.8kJ/mol**

* **The equation shows that 296.8 of heat is released when 1 mole od sulphur dioxide is formed from its elements.**

|  |
| --- |
| **Note:**   * **The overall energy change during a reaction is known as the enthalpy change(∆H) of the reaction.** * **The unit for enthalpy change is kJ/mol** * **The unit for energy is kJ** |

**Exercise 9**

1. **Give 3 differences between endothermic and exothermic reactions.**
2. **Consider the following equation which represents the burning of carbon monoxide.**

**2CO(g) + O2(g) 2O2(g) ∆H=-566kJ/mol**

**Calculate the enthalpy change when one mole of carbon monoxide burns.**

**LESSON NUMBER: 43**

**STRAND: 3. REACTIONS**

**SUB-STRAND: 3.3 PHYSICAL CHEMISTRY**

**LEARNING OUTCOME:**

**CHEMICAL EQUILIBRIUM**

* **State of equilibrium – when there it no net reaction either forwards or backwards.**
  + **Exists because the rate of forward reaction is exactly equal to the rate of reverse reaction, so that there is no longer any net reaction either way, and the reaction is said to in a state of dynamic equilibrium.**
* **Making Reversible reactions go to completion**
  + **By changing the equilibrium positions. The factors that can influence the reversible reaction are:**
    1. **Concentration**
    2. **Pressure**
    3. **Temperature**
    4. **Catalysts**
  + **The changes can be predicted using Le Chatelier’s Principle:**

“ **IF A REACTION IS AT EQUILIBRIUM AND ONE OF THE CONDITIONS IS CHANGED, THE POSITION OF THE EQUILIBRIUM WILL SHIFT IN SUCH A WAY THAT THE EFFECT OF THE CHANGE TENDS TO BE OPPOSED”**

* **Effects of changing conditions:**

1. **Changing the concentration e.g. A + B C + D**

**If more of the reactant (A or B) is added at equilibrium, the rate of the forward reaction will be increased because the concentration of the reactants is increasing. As a result, there will be more collisions between A and B. The equilibrium will shift towards the right (towards the product side). [If reactants are added or products are removed i.e. the concentration is changed, from a system at equilibrium, more products are formed]**

1. **Changing the temperature**

**An increase in temperature favours the ENDOTHERMIC reaction. In an exothermic process, a lowering of temperature favours the forward reaction.**

**When the temperature of a system in equilibrium is raised, a new equilibrium is formed which contains a higher proportion of the materials.**

1. **Using a catalyst**

**If a catalyst is used in a reversible reaction, the rates of both forward and backward reactions are increased equally, because of the lower activation energy.**

**A CATALYST DOES NOT ALTER THE EQUILIBRIUM POSITION (ie THE YIELD), BUT IT ENSURES THAT EQUILIBRIUM IS REACHED MORE QUICKLY THAN WITHOUT IT.**

1. **Changing the pressure**

**Affects reactions involving gases because gases can be compressed. If pressure is increased the system tries to decrease the effect of pressure by increasing the rate of reaction which produces fewer molecules (or moles) or a smaller volume.**

**E.g. SO2(g) + O2(g) 2SO3(g)**

**Reaction will proceed in the direction that produces a decreased volume; so the increasing of pressure increases the rate of the forward reaction.**

**REVERSIBLE REACTIONS**

**E.g. 5H2O(l) + CuSO4(s) CuSO4.5H2O(s) ΔH = -10 kJ/mol**

**CuSO4.5H2O(s) CuSO4(s) + 5H2O(l) ΔH = +10 kJ/mol**

**Combined:**

EXOTHERMIC

**5H2O(l) + CuSO4(s) CuSO4.5H2O(s)**

ENDOTHERMIC

**\*Note: ΔH symbol given with any equation for a reversible reaction refers to the forward reaction only. The reverse reaction will have a ΔH of the same value but opposite in sign.**

**Exercise 10:**

**For the following reactions what changes will you bring about to produce a new equilibrium with a higher yield of product i.e. favour the forward reaction? Explain your answers.**

**1. A2(g) + B2(g) 2AB(g) [ΔH = +210 kJ/mol]**

**2. D2(g) + 3E2(g) 2DE3**