CHEMISTRY ONLINE
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CHEMISTRY

PHYSICAL CHEMISTRY

Level & Board	AQA (A-LEVEL)
TOPIC:	CHEMICAL EQUILIBRIA
PAPER TYPE:	SOLUTION - 4
TOTAL QUESTIONS	10
TOTAL MARKS	37

Chemical Equilibria - 4

1.

Catalysts play a significant role in making chemical processes more sustainable for several reasons:

Reduced Energy Consumption:

Catalysts facilitate chemical reactions by providing an alternative reaction pathway that requires lower activation energy. This allows reactions to occur at lower temperatures and pressures than would be necessary without a catalyst.

Lowering the operating temperature and pressure saves energy by reducing the need for high-energy input for heating or pressurizing reactors, thereby decreasing overall energy consumption in the process.

Increased Efficiency and Yield:

Catalysts can enhance the efficiency of reactions by accelerating the rate of desired chemical transformations without being consumed themselves. This allows for higher yields of the desired products in a shorter time frame.

Improved selectivity, where catalysts promote the formation of specific products while minimizing unwanted by-products, contributes to higher overall process efficiency.

Decreased Environmental Impact:

Catalysts also enable the use of more environmentally benign reaction conditions, reducing the use of harsh chemicals or processes that generate hazardous waste.

Economic Advantages:

Enzymes can generate specific products.

The use of catalysts can lead to cost savings for chemical companies by reducing raw material consumption, increasing yield, and minimizing the need for expensive purification steps due to improved selectivity.

(4)

2. (C)

(Total 1 mark)

3.

(a)

The balanced equation is:

$$CH_3COOH(l) + C_2H_5OH(l) \rightleftharpoons CH_3COOC_2H_5(l) + H_2O(l)$$

Given that 72.5 cm^3 of a 1.50 mol dm^{-3} sodium hydroxide solution was used to react with the ethanoic acid, and assuming complete reaction:

Moles of sodium hydroxide used = Volume×Concentration
$=72.5\times10^{-3}\,dm3\times1.50\,mol\,dm^{-3}$
Moles of sodium hydroxide used = 0.109 mol
Moles of ethanoic acid at equilibrium = Moles of sodium hydroxide = 0.109
Moles of ester formed = Moles of water = Moles of acid reacted
= 0.200 - 0.109 = 0.091
Moles of ethanol = 0.110 - 0.091 = 0.019
The equilibrium constant Kc:

$$\mathrm{Kc} = \frac{[\mathrm{CH3COOC2H5}]\times[\mathrm{H2O}]}{[\mathrm{CH3COOH}]\times[\mathrm{C2H5OH}]}$$

$$\mathrm{Kc} = \frac{[0.091]\times(0.091)}{(0.109)\times(0.019)}$$

=3.99 no units

(5)

(b)

Similar Bond Types Broken and Made:

- Both reactants, ethanoic acid (CH_3COOH) and ethanol (C_2H_5OH), possess O-H bonds in their structures.
- In the products, the ester ($CH_3COOC_2H_5$) contains a C-O bond, and water (H_2O) contains an O-H bond.
- The reaction involves the breaking of O-H bonds in the reactants and the formation of O-H and C-O bonds in the products.

Equal Number of Bonds Broken and Made:

- In the reaction, one O-H bond is broken in both ethanoic acid and ethanol, resulting in the release of energy.
- Simultaneously, one O-H bond forms in water, and one C-O bond forms in the ester.
- The total number of bonds broken (O-H in reactants) is equal to the total number of bonds formed (O-H in water + C-O in ester), maintaining bond balance.

(4)

4. (A)

(Total 1 mark)

5. (C)

(Total 1 mark)

6.

(a)

Pressure: Typically maintained between 200 and 400 (atm).
Temperature: 500-600 °C (773 – 873 Kelvin).

(2)

(b)

Describe and explain why these conditions are a compromise between rate and equilibrium.

The relationships between pressure, temperature, rate, equilibrium, and the compromise required in the Haber process for ammonia production could be described following:

- **Rate and its Influences:**
  Higher temperatures generally increase the rate of the reaction between nitrogen and hydrogen to form ammonia. Elevated temperatures provide greater kinetic energy to the molecules, allowing them to collide more frequently and with higher energy, thereby increasing the reaction rate.
  Increased pressure increases the rate due to the closer distance of molecules or higher concentration.
- **Equilibrium Shifts:**
  - Increased pressure shifts equilibrium to the right-hand side (RHS) due to fewer gas moles on the right (favoring ammonia formation).
  - Elevated temperature shifts equilibrium to the left-hand side (LHS) due to the exothermic nature of the forward reaction.
- **Compromise in Conditions:**
  - High temperatures could lead to a lower yield as the equilibrium is shifted back towards the reactants' side.
  - Low temperatures might reduce the reaction rate.
  - High pressures can enhance the yield but can be costlier and pose safety concerns.

(6)

7. (C)

(Total 1 mark)

8.

(a)

Kc expression:

$$Kc = \frac{[HI]^2}{[H2]\ [I2]}$$

Kc (forward)×Kc (reverse)=Kc (net)
Kc (reverse)= 1/ Kc (forward)
Given that Kc (forward)=20:
Kc (reverse)=1/20=0.05

(2)

(b)

i.

- **Forward Reaction Rate:** Increasing the pressure leads to an increase in the rate of the forward reaction.
- **Reverse Reaction Rate:** Simultaneously, the rate of the reverse reaction also increases due to the pressure change.

(2)

ii.

- **No Change in Equilibrium Position:** Despite the change in pressure, the position of the equilibrium remains unaffected. The increase in pressure doesn't alter the equilibrium state.

(1)

iii.

No Change in Equilibrium Constant (Kc): The numerical value of the equilibrium constant (Kc) remains constant at a given temperature, unaffected by the pressure change.

(1)

9.

(a)

Increasing the pressure in a reaction involving gases ($2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$) would promote the formation of more product (H_2O) as per Le Chatelier's Principle. This is because the system would shift to reduce the pressure by favoring the side with fewer gas molecules. So, increasing the pressure would likely increase the rate of the reaction toward the formation of water molecules i.e product.

(2)

(b)

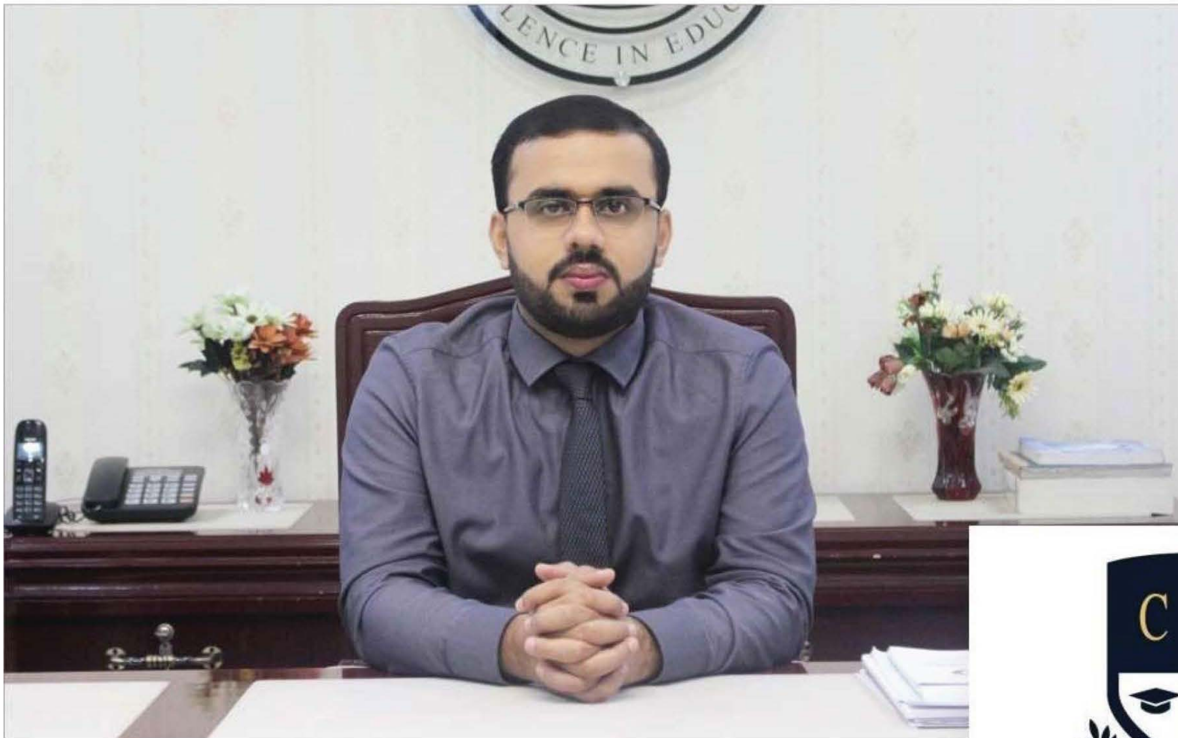
Highly exothermic reactions release a lot of heat rapidly. This sudden heat buildup can increase pressure, accelerate the reaction rate uncontrollably, and generate large volumes of gas. When these factors exceed the system's limits, it can cause an explosion due to the inability to manage the energy release and pressure buildup effectively.

(2)

10.

A homogeneous catalyst means a type of catalyst that exists in the same phase or state as the reactants in a chemical reaction. So, in a homogeneous catalytic reaction, the catalyst and the reactants are present in the same physical state, whether that be in a gaseous, liquid, or solid state.

(2)



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