



Grade 11



Chemistry



**Student Short Note with
Practice Questions
Based on Currently Used
Ethiopian curriculum**

By: Megersa Legesse

(0921079334) &

Gulilat Tayu

(0946377283)

April, 2020

2020, Oromia Educational Bureau

UNIT FOUR

1. Chemical Kinetics

1.1. INTRODUCTION

This unit is organized into four sections.

1. The first section is concerned with rate of reaction, with emphasis on how to determine rate of reaction and factors that influence rate of reaction.
2. The second section deals with theories of reaction rates, including collision theory and transition-state theory.
3. The third section deals with rate equation or rate law, which includes order of reaction, rate constant, concentration-time equation and half-life of a reaction.
4. The fourth section deals with reaction mechanism, which includes the molecularity of elementary reactions and the rate-determining step.

Chemical kinetics:- Is the area of chemistry concerned with the speeds, or rates, at which reactions occur. The word “**kinetic**” suggests movement or change. Chemical kinetics refers to the **rate of reaction**, which is the change over time in the concentration of a reactant or a product.

Chemical kinetics also the study of rates of chemical processes. This includes investigations of how different experimental conditions can influence the speed of a chemical reaction. The investigations yield information about the reaction's mechanism and transition states. The study of chemical kinetics also uses the construction of mathematical models that can describe the characteristics of a chemical reaction.

Rate of reaction (speed) is not fixed for a particular type of reaction, since it depends on several factors. The rate of reaction gets altered when subjected to small changes in temperature, concentration, pressure etc.

In 1864, Peter Waage and Cato Guldberg pioneered the development of chemical kinetics by formulating the law of mass action, which states that the speed of a chemical reaction is proportional to the quantity of the reacting substances.

Some chemical reactions proceed quickly, whereas others require days, months or even years to give products. These differences depend primarily on the differences in the chemical nature of the reacting substances. The rate of reaction defined for both reversible and irreversible reactions as well as slow and fast reactions

Fast reactions

- Combustions of fuel
- Reactions of active metal with acids
- Precipitation reactions
- Firing of gun
- Neutralization reactions

Slow reactions

- Rusting of Iron
- Oxidations of copper
- souring of milk
- Fermentations
- Geological reactions

Grade-11 Chemistry Short Note With Practice Questions

- Active metal with water

How do you think that the rate of reaction is measured?

Rate of reaction is the change in concentration of a reactant or a product over a given period of time. During the course of a reaction, the concentrations of the reactants decrease while those of the products increase. Thus, the rate of a reaction is calculated by dividing the change in the concentrations of reactants or products by the time taken for the change to occur.

For a simple reaction, $A \rightarrow B$

The rate of reaction is expressed in terms of a reactant concentration as

$$r = -\Delta [A] / \Delta t = [A]_f - [A]_i / \Delta t$$

Where r is rate of reaction, A is a reactant, B is a product, $[A]_i$ and $[A]_f$ are initial and final concentration of A respectively, $\Delta [A]$ is change in concentration of A and Δt is the change in time.

$[A]_f < [A]_i$. Thus, $\Delta [A] = [A]_f - [A]_i$, is negative. This indicates that the reactant is being consumed. However, the rate of a reaction is a positive quantity.

In terms of product concentration, the rate of a reaction is given by:

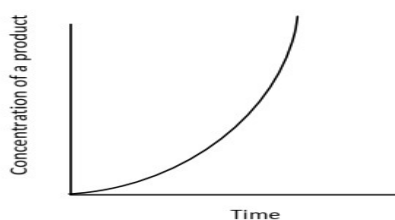
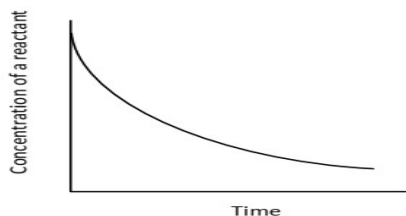
$$r = \Delta [B] / \Delta t = [B]_f - [B]_i / \Delta t$$

$[B]_f > [B]_i$. Thus, $\Delta B = [B]_f - [B]_i$ is positive because $[B]$ increases and the rate of a reaction is also positive.

The unit of rate is $\text{mol L}^{-1} \text{s}^{-1}$ (Ms^{-1}) and the unit of concentration is mole per litre (mol/L)

1.2 Reaction Rate and Stoichiometry

Rate expressions in terms of the disappearance of the reactants and the appearance of the products:



A graph of changes in concentration of a reactant and a product over time where time is written as the x-axis and concentration as the y-axis.

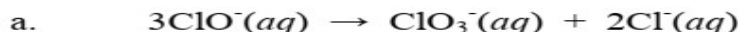
In general, for the reaction, $aA + bB \rightarrow cC + dD$

$$r = -\frac{1}{a} \frac{\Delta [A]}{\Delta t} = -\frac{1}{b} \frac{\Delta [B]}{\Delta t} = \frac{1}{c} \frac{\Delta [C]}{\Delta t} = \frac{1}{d} \frac{\Delta [D]}{\Delta t}$$

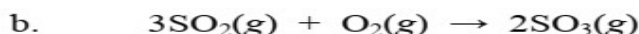
Grade-11 Chemistry Short Note With Practice Questions

Example

1. Write the rate expressions for the following reactions, in terms of the disappearance of the reactants and the appearance of the products:



$$\text{rate of reaction} = -\frac{\Delta[\text{ClO}^-]}{3\Delta t} = \frac{\Delta[\text{ClO}_3^-]}{\Delta t} = \frac{\Delta[\text{Cl}^-]}{2\Delta t}$$



$$\text{rate of reaction} = -\frac{\Delta[\text{SO}_2]}{3\Delta t} = -\frac{\Delta[\text{O}_2]}{\Delta t} = \frac{\Delta[\text{SO}_3]}{2\Delta t}$$



$$\text{rate of reaction} = -\frac{\Delta[\text{C}_2\text{H}_4]}{\Delta t} = -\frac{\Delta[\text{Br}_2]}{\Delta t} = \frac{\Delta[\text{C}_2\text{H}_4\text{Br}_2]}{\Delta t}$$

2. Consider the hypothetical reaction.



What is the rate of formation of $\text{XY}_2(g)$

A. $-5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$

C. $2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$

B. $5.0 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$

D. $-2.5 \times 10^{-2} \text{ mol L}^{-1} \text{ sec}^{-1}$

- C. Answer:- C The formation of the product is half of the rate of the loss of Y. It is correct to change the sign to positive, because we are representing the formation of product, and the negative sign on the reactant to represent loss.

2. For the chemical reaction $\text{H}_2 + \text{Br}_2 \rightarrow 2\text{HBr}$, what can be said of the relative rates of consumption of the reactants when compared to the formation of product?

- A. HBr will be produced at half the rate that bromine is consumed.
- B. HBr will be produced at the same rate that bromine is consumed.
- C. Br_2 and H_2 will be consumed at different rates.
- D. HBr will be produced at twice the rate that bromine is consumed.
- E. It is not possible to tell anything about relative rates without the rate law.

Answer: - D The stoichiometry of the overall reaction will give us answers to the questions about comparative rates of formation of products compared to loss of reactants.

3. For the reaction $\text{A} + 3\text{B} \rightarrow 2\text{C}$, how does the rate of disappearance of B compare to the rate of production of C?

- A. The rate of disappearance of B is 1/2 the rate of appearance of C
- B. The rate of disappearance of B is 3/2 the rate of appearance of C
- C. The rate of disappearance of B is 2/3 the rate of appearance of C
- D. The rate of disappearance of B is 1/3 the rate of appearance of C

Answer: - B Write in terms of B

1.3 Determination of Rate of Reaction

- The rate of reaction can be determined both quantitatively and qualitatively.

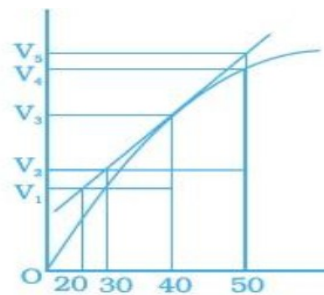
Grade-11 Chemistry Short Note With Practice Questions

- ✓ **Qualitatively:** - An idea about the rate of reaction can be obtained by observing either the speed of appearance of the reactants or the speed of appearance (formation) of the products.
- ✓ **Quantitatively:-** Quantity obtained from formula

EXAMPLE

1. A graph of volume of hydrogen released vs time for the reaction between zinc and dil.HCl is given in Fig. On the basis of this mark the correct option is?

- (i) Average rate upto 40s is $\frac{V_3 - V_2}{40}$
- (ii) Average rate upto 40 seconds is $\frac{V_3 - V_2}{40 - 30}$
- (iii) Average rate upto 40 seconds is $\frac{V_3}{40}$
- (iv) Average rate upto 40 seconds is $\frac{V_3 - V_1}{40 - 20}$



Answer: - iii because average rate up to 40 means $V_3 - V_0 / 40 - 0 = V_3 / 40$

2. Consider the decomposition of N_2O_5 to give NO_2 and O_2 :



Time (s)	Concentration (M)		
	N_2O_5	NO_2	O_2
0	0.0200	0	0
100	0.0169	0.0063	0.0016
200	0.0142	0.0115	0.0029
300	0.0120	0.0160	0.0040
400	0.0101	0.0197	0.0049
500	0.0086	0.0229	0.0057
600	0.0072	0.0256	0.0064
700	0.0061	0.0278	0.0070

From the graph looking at $t = 300$ to 400 s

$$\text{Rate } O_2 = \frac{0.0009M}{100s} = 9 \times 10^{-6} \text{ Ms}^{-1}$$

Why do they differ?

$$\text{Rate } NO_2 = \frac{0.0037M}{100s} = 3.7 \times 10^{-5} \text{ Ms}^{-1}$$

Recall:

$$\text{Rate } N_2O_5 = \frac{0.0019M}{100s} = 1.9 \times 10^{-5} \text{ Ms}^{-1}$$



To compare the rates one must account for the **stoichiometry**.

$$\text{Rate } O_2 = \frac{1}{1} \times 9 \times 10^{-6} \text{ Ms}^{-1} = 9 \times 10^{-6} \text{ Ms}^{-1}$$

$$\text{Rate } NO_2 = \frac{1}{4} \times 3.7 \times 10^{-5} \text{ Ms}^{-1} = 9.2 \times 10^{-6} \text{ Ms}^{-1}$$

$$\text{Rate } N_2O_5 = \frac{1}{2} \times 1.9 \times 10^{-5} \text{ Ms}^{-1} = 9.5 \times 10^{-6} \text{ Ms}^{-1}$$

Now they agree!

what about with the time interval ($t = 100$ to 500 s)?

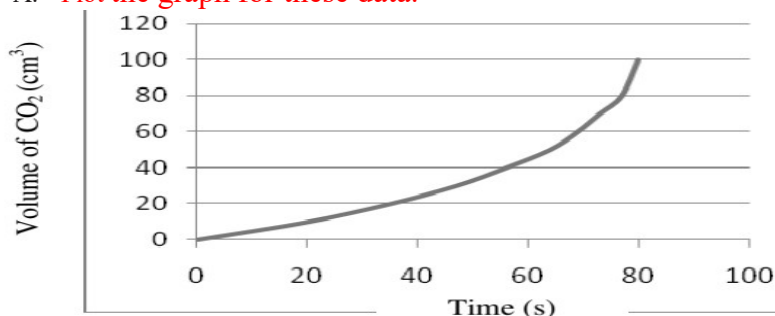
Grade-11 Chemistry Short Note With Practice Questions

3. The results obtained during the reaction between excess of powdered calcium carbonate and hydrochloric acid, are given in the following table:



Volume of $\text{CO}_2 / \text{cm}^3$	0	20	35	47	56	64	69	73	77	80
Time/s	0	10	20	30	40	50	60	70	80	100

- A. Plot the graph for these data.



- B. What volume of gas has formed after 50 seconds?

$$r = \frac{\Delta[\text{CO}_2]}{\Delta t} = \frac{(64-0)\text{cm}^3}{(50-0)\text{s}} = \frac{64}{50}\text{cm}^3\text{s}^{-1} = 1.28\text{cm}^3\text{s}^{-1}$$

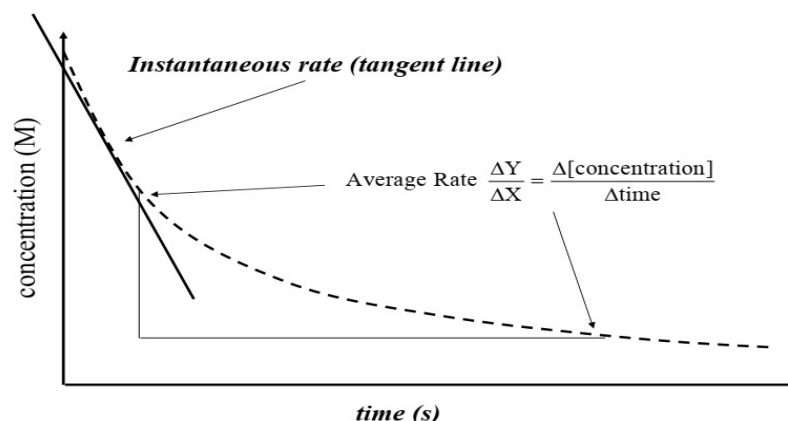
- C. What volume of gas has formed after 15 seconds?

$$r = \frac{\Delta[\text{CO}_2]}{\Delta t} = 1.28\text{cm}^3\text{s}^{-1} = \frac{V_2 - V_1}{(15-0)\text{s}} \Rightarrow V_2 - 0 = 19.2\text{cm}^3 \Rightarrow V_2 = 19.2\text{cm}^3$$

- D. How long did it take for 40 cm³ of carbon dioxide to form?

$$r = \frac{\Delta[\text{CO}_2]}{\Delta t} = 1.28\text{cm}^3\text{s}^{-1} = \frac{40\text{cm}^3}{\Delta t} \Rightarrow \Delta t = \frac{40\text{cm}^3}{1.28\text{cm}^3\text{s}^{-1}} \Rightarrow \Delta t = 31.25\text{s}$$

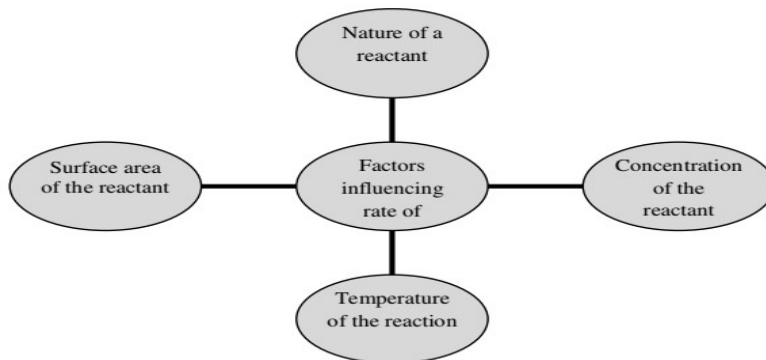
Instantaneous rate:- is defined as the rate of a reaction for a given instant of time and it is obtained from the slope of the tangent to the concentrations curve at that time



1.2 Factors Influencing the Reaction Rates

In general, different chemical reactions have different rates. Even a chemical reaction involving the same reactants may have different rates under different conditions.

Grade-11 Chemistry Short Note With Practice Questions



- 1. Temperature of reactants:-** When temperature increases, rate of reaction increases
Reason:- When temperature increases overall kinetic energy of all particles of reactants increases, so more molecules will have energy greater than activation energy. So, the overall frequency of collisions increases, therefore frequency of effective collisions increases, and thus the reaction rate increases
- 2. Concentration of reactant:-** The concentration of one or more reactants increases, rate of reaction increases.
Reason:- Increasing the concentration one or more reactants produce more contacts between the reacting particles, which results in increasing the rate of reaction.

For gaseous reactants, an increase in pressure increases reactions rate
Reason:- Increasing pressure means the volume decreases implies contact (concentration) increase as result collision increases and reaction will be fast. However, pressure change has no effect on the rate of reaction if the reactants are **solids** or **liquids**.
- 3. Surface Area of Reactants:- as surface area increase rate of reaction increases**
Reason:- When the size of the substance subdivisions decreases, surface area increases. It makes the effective collisions increase and therefore the rate of reaction also increases.
- 4. Nature of reactants:-** The rate of a reaction depends on the chemical nature of the substances participating in the reaction. A more active a substance the reactive
- 5. Effect of Catalyst on Rate of Reaction** A catalyst is a substance that changes reaction rate by providing a different reaction mechanism one with lower activation energy

Characteristics of catalyst

- Speeds up the reaction forward and backward equally
- May even form temporary chemical bonds with the reactants
- Recovered unchanged in original form at the end of the reaction
- Does not alter the position of equilibrium
- It provides an alternate mechanism by reducing activation energy
- Does not consumed itself in the reactions

Grade-11 Chemistry Short Note With Practice Questions

EXAMPLE

1. The rate of a chemical reaction
 - A. Is always dependent of the concentration of all reactants.
 - B. Is always increased with increasing temperatures.
 - C. Is greater with higher activation energies.
 - D. Is independent of surface area.

Answer: (B) The collision model indicates that molecular collision is necessary for reaction. Because an increase in temperature raises molecular velocities and the percentage of effective collisions, the reaction rate increases

2. Which of the following statements is correct?
 - A. The rate of a reaction decreases with passage of time as the concentration of reactants decreases.
 - B. The rate of a reaction is same at any time during the reaction
 - C. The rate of a reaction is independent of temperature change.
 - D. The rate of a reaction decreases with increase in concentration of reactant(s) Ans A
3. How does a catalyst increase the rate of a chemical reaction?
 - A. By changing the magnitude of the enthalpy change.
 - B. By providing a pathway for the reaction that has lower activation energy.
 - C. By decreasing the potential energy of the products
 - D. By increasing the potential energy of the reactants
 - E. By giving energy to the reactants so more of them have energies greater than the activation energy Answer B

4. Why do large crystals of sugar burn more slowly than finely ground sugar?

Answer:- The rate of combustion is proportional to the surface area of sugar exposed to oxygen. Smaller crystals have more surface area and burn faster.

5. How do homogeneous catalysts and heterogeneous catalyst differ?

A homogeneous catalyst is a catalyst that exists in the same phase (liquid or gas) as the reactants. A

heterogeneous catalyst, on the other hand, is a catalyst that exists in a different phase from the reactants.

1.3 THEORIES OF REACTION RATES

Two important theories help us to explain why different reactions occur at different rates. These are

- The collision theory and
- The transition-state theory.

These two theories are distinct but in complete agreement, each emphasizing different aspects of reaction processes.

The collision theory

- The fundamental concept of the collision theory of rate of reaction is that, in order for a reaction to occur between reacting species (*atoms, ions or molecules*), they must first collide (come in contact). However, not all collisions between reacting species result in a reaction. This is because collisions between reactants can be either effective or ineffective.
- *Effective collisions*: - are collisions that result in a reaction to form the desired products.
- *Ineffective collisions*: - are collisions that do not result in a reaction to form the desired products.

Postulates of collision theory are summarized below.

- For a reaction to occur between reactants:
 1. The particles of the reacting substances must collide.
 2. The particles of the reacting species must have proper orientation.
 3. The particles of the reactants must collide with energy greater than or equal to the energy of activation

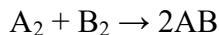
Note:- Linear collision is more effective than non linear collision

Transition State Theory (Activated Complex and Energy Barrier)

According to transition-state theory, the collision between two reacting species results in the formation of *an activated complex or transition state*. An activated complex or transition state is

- High energy state species that is temporarily formed by the collision of reactant molecules before they form products.
- The species can either fall to back(reactant) or form product
- Short-lived , unstable and barrier
- In the activated complex, the original bonds are lengthened and weakened, and the new bonds are only partially formed

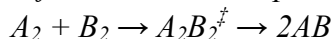
Let us consider the reaction:



The reactants, A_2 and B_2 , approach each other and collide; they begin climbing up the left side of the potential-energy hill, or barrier.

If they have less energy than the energy of activation, E_a , they fail to climb the barrier and, instead, roll back down its left side. In this case, no reaction occurs.

The formation of an activated complex in the form of an equation can be shown as,



Where $A_2B_2^\ddagger$ is the activated complex or transition state

Grade-11 Chemistry Short Note With Practice Questions

Example

1. Which applies to this reaction as it proceeds at constant temperature?
 - I. The rate of the reaction decreases.
 - II. The effectiveness of collisions between reactant molecules remains the same.
 - III. The frequency of collisions between reactant molecules remains the same.

A. I only B. II only C. I and II only D. II and III only

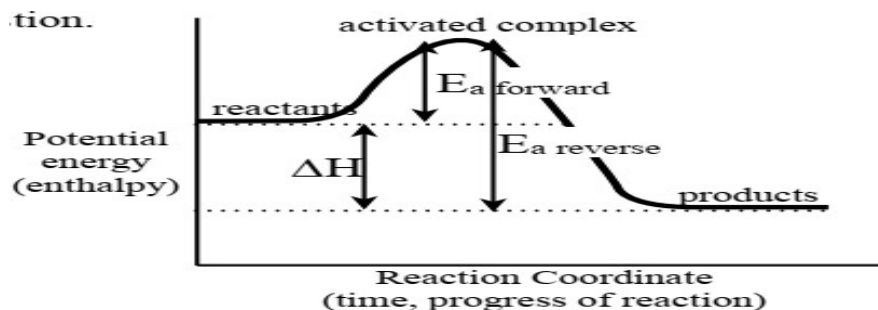
Answer C: - As the concentrations of the reactants decrease over the course of a reaction, the frequency of collisions will decrease, but the effectiveness of the collisions will remain the same, as that is a function of the molecules themselves, the temperature, and the presence of any catalyst.

2. Collision theory predicts all of the following EXCEPT that
 - A. A reaction will only occur if the collision geometry is correct
 - B. A reaction will not occur if the collision occurs with energy that is less than the activation energy.
 - C. A reaction will not occur if the reactants do not collide
 - D. Frequency of collisions will increase with increasing temperature.
 - E. More successful collisions will occur for a reaction with larger activation energy.

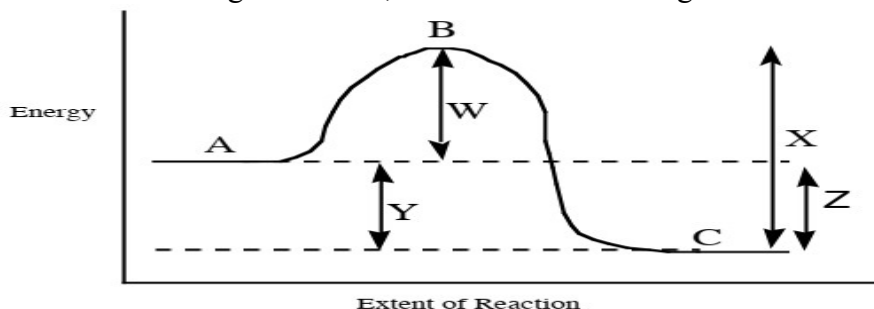
Answer E: - Higher activation energy is a higher hump, and is harder to get over, making for fewer successful collisions.

3. The values for the change in enthalpy, ΔH , and the activation energy, E_a , for a given reaction are known. The value of E_a for the reverse reaction equals
 - A. E_a for the forward reaction.
 - B. The sum of $-(\Delta H)$ and E_a
 - C. The difference between ΔH and E_a
 - D. $-(E_a)$ for the forward reaction.
 - E. The sum of E_a and ΔH

Answer B, see the fig

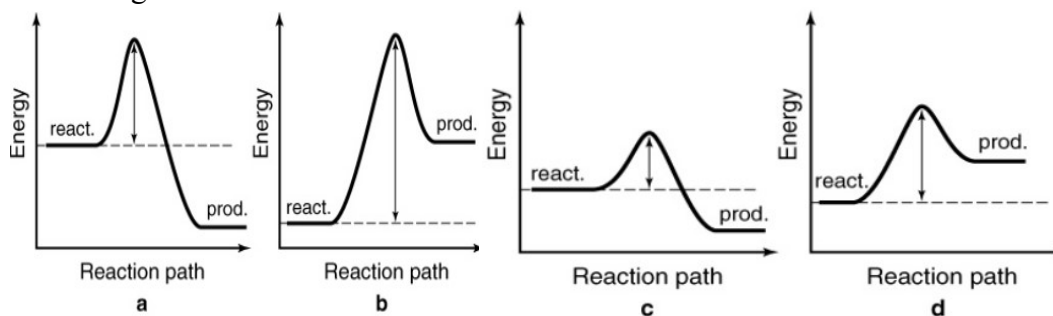


4. For the reaction diagram shown, which of the following statements is true?



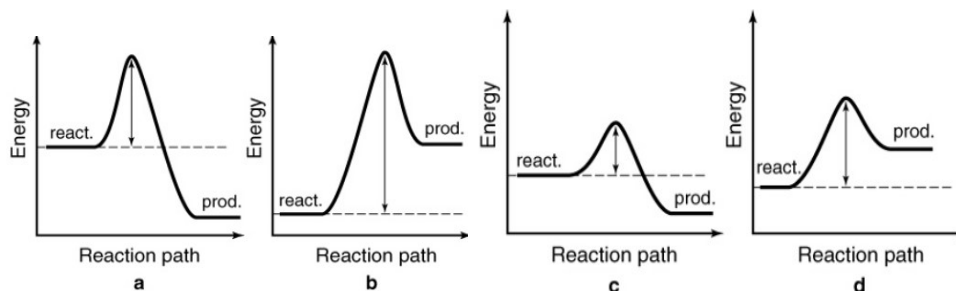
Grade-11 Chemistry Short Note With Practice Questions

- A. Line W represents the ΔH for the forward reaction; point B represents the transition state
- B. Line W represents the activation energy for the forward reaction; point B represents the transition state
- C. Line Y represents the activation energy for the forward reaction; point C represents the transition state
- D. Line X represents the ΔH for the forward reaction; point B represents the transition state. **Ans. B**
5. The energy profiles for four different reactions are shown. Which reaction requires the most energetic collisions to reach the transition state?



Answer B

6. Collision theory assumes that the rate of a reaction depends on _____
- A. The energy of collisions.
- B. The orientation of colliding molecules.
- C. The energy of collisions and the orientation of colliding molecules.
- D. The change in energy between the products and the reactants. **Ans. C**
7. The energy needed for a reaction to proceed from reactants to products is called:
- A. Collision energy C. Kinetic energy
- B. Potential energy D. Activation energy. **Ans. D**
8. The following energy profiles for four different reactions are shown. Which reaction is the most endothermic?



Ans. B

1.4 RATE EQUATION OR RATE LAW

1.4.1 Order of Reaction and Rate Constant

- **The rate law or rate equation:** - for a chemical reaction is an equation which links the reaction rate with the concentrations or pressures of reactants and certain constant parameters (*normally rate coefficients and partial reaction order*). Each concentration is expressed with an order (exponent).

Grade-11 Chemistry Short Note With Practice Questions

For the general reaction: $aA + bB \rightarrow cC + d$

Rate = $k[A]^x[B]^y$ x and y are the reactant order determined from experiment.

x and y are **NOT** the stoichiometric coefficients.

Consider a reaction with the rate equation

$r = k[A]^2[B]$, order with respect to A is 2 and 1 with respect to B

$2 + 1 = 3$ over all order

If the $[A]^2$ is doubled and remain constant the $[B]$ the rate will be increase by factor of 4

If the $[A]^2$ is tripled and remain constant the $[B]$ the rate will be increase by factor of 9

Change in initial rate with change in $[A]_0$ for rate = $k[A]^m$

For a reaction of	If $[A]_0$ is doubled	If $[A]_0$ is tripled
Zero-order ($m = 0$)	rate is unchanged	rate is unchanged
First-order ($m = 1$)	rate is doubled	rate is tripled
Second-order ($m = 2$)	rate is quadrupled ($\times 2^2$)	rate is 9 times greater ($\times 3^2$)

1.5 Determining Reaction Order: The Method of Initial Rates

Example

1. The reaction, $2I^-(aq) + S_2O_8^{2-}(aq) \rightarrow I_2(aq) + 2SO_4^{2-}(aq)$ was studied at 25°C. The following results were obtained where

From the following data, A. determine the order

B. Rate law and

C. Rate constant.

$[I^-]_0$ (mol/L)	$[S_2O_8^{2-}]_0$ (mol/L)	Initial Rate (mol/L • s)
0.080	0.040	12.5×10^{-6}
0.040	0.040	6.25×10^{-6}
0.080	0.020	6.25×10^{-6}
0.032	0.040	5.00×10^{-6}
0.060	0.030	7.00×10^{-6}

- The rate law for the reaction is given by: - Rate (M/min) = $k[I]^x[S_2O_8^{2-}]^y$, but we should have to calculate for the value of X and Y.

✓ To have the value of X keeps constant the value of y and make varies the value of X.

✓ To have the value of y keeps constant the value of X and vary the value of y

Taking the ratio of the rates of runs 2 and 1 one finds:

A. Determine the rate law.

Grade-11 Chemistry Short Note With Practice Questions

$$\text{Rate} = k[\text{I}^-]^x [\text{S}_2\text{O}_8^{2-}]^y$$

Compare experiments 1 and 2.

$$\frac{\text{rate}_2}{\text{rate}_1} = \frac{k[\text{I}^-]_2^x [\text{S}_2\text{O}_8^{2-}]_2^y}{k[\text{I}^-]_1^x [\text{S}_2\text{O}_8^{2-}]_1^y}$$

$$\frac{6.25 \times 10^{-6}}{12.5 \times 10^{-6}} = \frac{k(0.040)^x (0.040)^y}{k(0.080)^x (0.040)^y}, 0.5 = 0.5^x, x = 1$$

$$\frac{\text{rate}_3}{\text{rate}_1} = \frac{k[\text{I}^-]_3^x [\text{S}_2\text{O}_8^{2-}]_3^y}{k[\text{I}^-]_1^x [\text{S}_2\text{O}_8^{2-}]_1^y}$$

$$\frac{6.25 \times 10^{-6}}{12.5 \times 10^{-6}} = \frac{k(0.080)^x (0.020)^y}{k(0.080)^x (0.040)^y}, 0.5 = 0.5^y, y = 1$$

$$\text{Rate} = k[\text{I}^-][\text{S}_2\text{O}_8^{2-}]$$

SO order with respect to $[\text{I}^-]$ is 1 ($x = 1$) and $[\text{S}_2\text{O}_8^{2-}]$ is 1 ($y = 1$)

- B. Calculate a value for the rate constant for each experiment and an average value for the rate constant. Rate constant can be calculated from all or you can choose one the them

$$\text{Experiment 1} = \frac{12.5 \times 10^{-6} \text{ mol}}{\text{L} \cdot \text{s}} = k \left(\frac{0.080 \text{ mol}}{\text{L}} \right) \left(\frac{0.040 \text{ mol}}{\text{L}} \right), k = 3.9 \times 10^{-3} \frac{\text{L}}{\text{mol} \cdot \text{s}}$$

$$\text{Experiment 2} = \frac{6.25 \times 10^{-6} \text{ mol}}{\text{L} \cdot \text{s}} = k \left(\frac{0.040 \text{ mol}}{\text{L}} \right) \left(\frac{0.040 \text{ mol}}{\text{L}} \right), k = 3.9 \times 10^{-3} \frac{\text{L}}{\text{mol} \cdot \text{s}}$$

$$\text{Experiment 3} = \frac{6.25 \times 10^{-6} \text{ mol}}{\text{L} \cdot \text{s}} = k \left(\frac{0.080 \text{ mol}}{\text{L}} \right) \left(\frac{0.020 \text{ mol}}{\text{L}} \right), k = 3.9 \times 10^{-3} \frac{\text{L}}{\text{mol} \cdot \text{s}}$$

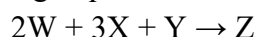
$$\text{Experiment 4} = \frac{5.00 \times 10^{-6} \text{ mol}}{\text{L} \cdot \text{s}} = k \left(\frac{0.032 \text{ mol}}{\text{L}} \right) \left(\frac{0.040 \text{ mol}}{\text{L}} \right), k = 3.9 \times 10^{-3} \frac{\text{L}}{\text{mol} \cdot \text{s}}$$

$$\text{Experiment 5} = \frac{7.00 \times 10^{-6} \text{ mol}}{\text{L} \cdot \text{s}} = k \left(\frac{0.060 \text{ mol}}{\text{L}} \right) \left(\frac{0.030 \text{ mol}}{\text{L}} \right), k = 3.9 \times 10^{-3} \frac{\text{L}}{\text{mol} \cdot \text{s}}$$

$$K_{\text{mean}} = 3.9 \times 10^{-3} \text{ L/mol} \cdot \text{s}$$

C. Rate law = $k[\text{I}^-][\text{S}_2\text{O}_8^{2-}]$

2. The following 2 questions refer to the following equation and table below:



Trial	$[\text{W}] \text{ mol L}^{-1}$	$[\text{X}] \text{ mol L}^{-1}$	$[\text{Y}] \text{ mol L}^{-1}$	Rate of Production of $\text{Z} \text{ mol L}^{-1}$
1	0.001	0.001	0.001	2.1×10^{-4}
2	0.002	0.001	0.001	2.1×10^{-4}
3	0.001	0.003	0.001	6.3×10^{-4}
4	0.002	0.001	0.0005	5.25×10^{-5}

The rate law is:

A. Rate = $k[\text{W}]^2 [\text{X}]^3 [\text{Y}]$

C. Rate = $k[\text{X}]^3 [\text{Y}]^2$

B. Rate = $k[\text{W}]^2 [\text{X}] [\text{Y}]$

D. Rate = $k[\text{X}] [\text{Y}]^2$ Answer D or $[\text{W}]^0 = 1$

3. The reaction $\text{X} + \text{Y} \rightarrow \text{Z}$ has the rate law, Rate = $k[\text{X}]^2 [\text{Y}]$. If the concentration of X is doubled and the concentration of Y is tripled, the rate would be increased by a factor of

A. 2 B. 4 C. 5 D. 6 E. 12

ANS E, Rate = $k[\text{X}]^2 [\text{Y}]$ thus $[2]^2 [3] = 12$

Grade-11 Chemistry Short Note With Practice Questions

4. For a reaction $2A + B \rightarrow 2C$, with the rate equation: $\text{Rate} = k[A]^2[B]$
- The order with respect to A is 1 and the order overall is 1.
 - The order with respect to A is 2 and the order overall is 2.
 - The order with respect to A is 2 and the order overall is 3.
 - The order with respect to B is 2 and the order overall is 2.
 - The order with respect to B is 2 and the order overall is 3. *ANS. C*
5. For the rate law, $\text{Rate} = k[Q]^2[R]^2$, which statement is incorrect?
- The reaction is fourth order overall, and doubling both the [Q] and [R] while keeping the temperature constant will lead to a bigger k and a 16-fold increase in the rate
 - The reaction is fourth order overall, and doubling both the [Q] and [R] while keeping the temperature constant will lead to a 16-fold increase in the rate
 - The reaction is fourth order overall, and doubling both the [Q] and [R] while keeping the temperature constant will lead to a bigger k and an 8-fold increase in the rate
 - The reaction is second order overall, and doubling both the [Q] and [R] while keeping the temperature constant will lead to a bigger k and a 16-fold increase in the rate
 - The reaction is second order overall, and doubling both the [Q] and [R] while keeping the temperature constant will lead to a 16-fold increase in the rate
- ANS. B. Changing concentrations will not affect the rate constant*
6. If tripling the concentration of a single reactant R in a multi-reactant reaction (while leaving all other conditions unchanged) leads to a nine-fold increase in rate, it can be deduced
- That the complete rate law is $\text{Rate} = k[R]$
 - That the complete rate law is $\text{Rate} = k[R]$
 - That the complete rate law is $\text{Rate} = k[R]^3$
 - that the complete rate law is $\text{Rate} = k[R]^9$
 - Only that the order with respect to [R] is 2 *ANS. E*
7. For the rate law shown below, where the rate is measured in units of $M s^{-1}$ and the concentrations of X and Y are measured in M, which of the following statements is true?
- $$\text{Rate} = k[X]^2[Y]$$
- The units of $k = M^2 s^{-1}$
 - Doubling the concentration of y while keeping all other conditions constant will not affect the rate.
 - The reaction can be described as third-order.
- I only
 - III only
 - I and II only
 - II and III only
 - I, II, and III
- ✓ *ANS. B* Add the individual orders to determine the overall order. For item I, the units would be $M^{-2}s^{-1}$, and doubling Y would double the rate.
8. Consider the reaction: $P_4 + 6 H_2 \rightarrow 4 PH_3$. A rate study of this reaction was conducted at 298 K. The data that were obtained are shown in the table.

$[P_4]$, mol/L	$[H_2]$, mol/L	Initial Rate, mol/(L · s)
0.0110	0.0075	3.20×10^{-4}
0.0110	0.0150	6.40×10^{-4}
0.0220	0.0150	6.39×10^{-4}

Grade-11 Chemistry Short Note With Practice Questions

- A. What is the order with respect to: $P_4 = 0$, $H_2 = 1$.
- B. Write the rate law for this reaction. $rate = k[H_2]^1[P_4]^0$
- C. Determine the value and units of the rate constant, k . **plug and check using the rate law & data from experiment 1 and solving for k , we get $k = 0.0427 \text{ s}$**

1.6 The Arrhenius Equation

- **Svante Arrhenius**: - Observed that a mathematical relationship connects **activation energy**, **temperature**, and **the rate constant**. This relationship is known as the Arrhenius equation and is written as:

Arrhenius equation

$$k = Ae^{\frac{-E_a}{RT}}$$

k is the rate constant

T is the temperature in K

E_a is the activation energy

R is the ideal-gas constant (8.314 J/Kmol)

A is known the **frequency or pre-exponential factor**

In addition to carrying the units of the rate constant, “**A**” relates to the frequency of collisions and the orientation of a favorable collision probability

Both A and E_a are *specific to a given reaction*.

- ❖ **Arrhenius** discovered that most reaction-rate data obeyed an equation based on three factors:
1. The number of collisions per unit time.
 2. The fraction of collisions that occur with the correct orientation.
 3. The fraction of the colliding molecules that have energy greater than or equal to E_a .

By combining the equations for the values of k at two **different temperatures** T_1 and T_2 for the same reaction, a relationship is obtained

$$\ln \frac{k_2}{k_1} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right) \quad \text{OR} \quad \text{Log} k_2/k_1 = E_a/2.303R (1/T_1 - 1/T_2)$$

Example

1. Which of the following statements are in accordance with the Arrhenius equation?
 - A. Rate of a reaction increases with increase in temperature.
 - B. Rate of a reaction increases with decrease in activation energy.
 - C. Rate constant decreases exponentially with increase in temperature.
 - D. Rate of reaction decreases with decrease in activation energy. **ANS:- A & B**
2. Consider the Arrhenius equation given below and mark the correct option.

$$K = A^{E_a/RT}$$

Grade-11 Chemistry Short Note With Practice Questions

- A. Rate constant increases exponentially with increasing activation energy and decreasing temperature.
- B. Rate constant decreases exponentially with increasing activation energy and decreasing temperature.
- C. Rate constant increases exponentially with decreasing activation energy and decreasing temperature.
- D. Rate constant increases exponentially with decreasing activation energy and increasing temperature. *Ans. D*
3. The activation energy of a first order reaction is 50.2 kJ/mol at 25°C. At what temperature will the rate constant double?

$$(1) \quad k_2 = 2k_1 \qquad (2) \quad \ln\left(\frac{k_2}{k_1}\right) = \ln\left(\frac{2k_1}{k_1}\right) = \ln(2) = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

$$(3) \quad \frac{E_a}{R} = \frac{50.2 \text{ kJ/mol} \times \frac{10^3 \text{ J}}{1 \text{ kJ}}}{8.314 \frac{\text{J}}{\text{mol} \cdot \text{K}}} = 6.04 \times 10^3 \text{ K}$$

$$(4) \quad \ln(2) = 0.693 = 6.04 \times 10^3 \text{ K} \times \left(\frac{1}{298 \text{ K}} - \frac{1}{T_2} \right) \qquad \text{A } 10^\circ\text{C change of temperature doubles the rate!!}$$

$$(5) \quad \frac{1}{T_2} = 3.24 \times 10^{-3} \text{ K}^{-1} \longrightarrow \boxed{T_2 = 308 \text{ K}}$$

4. Consider the reaction: $2B \rightarrow C + 3D$. In one experiment it was found that at 300 K the rate constant is 0.134 L/(mol.s). A second experiment showed that at 450 K, the rate constant was 0.569 L/(mol.s). Determine the activation energy for the reaction.

at 300 K: $k_{300} = Ae^{\frac{-E_a}{RT}}$

at 450 K: $k_{450} = Ae^{\frac{-E_a}{RT}}$

$$\ln \frac{k_{450}}{A} = \frac{-E_a}{RT}$$

$$\ln(k_{450}) - \ln(A) = \frac{-E_a}{RT} \quad \text{where } \ln(A) = \ln(k_{300}) - \frac{-E_a}{RT}$$

so that

$$\ln(k_{450}) - \left[\ln(k_{300}) - \frac{-E_a}{RT} \right] = \frac{-E_a}{RT}$$

$$\ln\left(\frac{k_{450}}{k_{300}}\right) = \frac{E_a}{R} \left(\frac{1}{T_{300}} - \frac{1}{T_{450}} \right)$$

plug and solve for E_a , $E_a = 10.8 \text{ kJ}$

5. Given that E_a for the hydrolysis of sucrose is $108 \times 10^3 \text{ kJ/mol}$, compare the rate constant of this reaction at 37°C (T_1) with the rate constant of the same reaction at 27°C (T_2).
- Given that E_a for the hydrolysis of sucrose is $108 \times 10^3 \text{ kJ/mol}$, compare the rate constant of this reaction at 37°C (T_1) with the rate constant of the same reaction at 27°C (T_2).

Grade-11 Chemistry Short Note With Practice Questions

$$T_1 = 37 + 273 = 310 \text{ K} \quad T_2 = 27 + 273 = 300 \text{ K}$$

$$R = 8.31 \text{ J/K.molEa} = 108 \times 10^3 = 1.08 \times 10^5 \text{ J/mol} \quad A = \text{not given}$$

$$\ln k_1 - \ln k_2 = \left(\ln a - \frac{E_a}{RT_1} \right) - \left(\ln a - \frac{E_a}{RT_2} \right)$$

$$\ln k_1 - \ln k_2 = \ln a - \frac{E_a}{RT_1} - \ln a + \frac{E_a}{RT_2} = -\frac{E_a}{RT_1} + \frac{E_a}{RT_2}$$

$$\ln \frac{k_1}{k_2} = -\frac{E_a}{RT_1} + \frac{E_a}{RT_2} \quad (\text{since } \ln k_1 - \ln k_2 = \ln k_1 / k_2)$$

$$\ln \frac{k_1}{k_2} = -\frac{1.08 \times 10^5}{8.314 \times 310} + \frac{1.08 \times 10^5}{8.314 \times 300} = 1.396$$

$$\frac{k_1}{k_2} = 4.04$$

The reaction at 37 °C is about 4 times faster than at 27 °C.

1.7 Concentration-Time Equation (Integrated Rate Law)

- ❖ Rate law tells us how the rate of reaction depends on concentrations of the reactants at a particular moment

1.8 Zero Order Reaction

- Whose rate of reaction does not depend on the reactant concentration
- The rates of these **zero-order reactions** do not vary with increasing nor decreasing reactants concentrations. This means that the rate of the reaction is equal to the rate constant, k , of that reaction.



$$[A]_0 - [A]_t = -kt_0$$

$$[A]_t = -kt + [A]_0 \quad \text{Integrated rate equation for 0}^{\text{th}} \text{ order reaction}$$

Where $[A]$ is concentration of A at t ; $[A]_0$ is the concentration of A at $t = 0$.

- This is a linear equation with the form of $y = mx + b$.
 - For a zero-order reaction, a plot of $[A]$ versus-time is a straight line, whose slope is $-k$
- Example

1. For all zero-order reactions
 - A. A plot of time vs. concentration squared is linear.
 - B. The rate is independent of time.
 - C. The rate constant is zero.
 - D. E_a is very low

Answer: (B) For zero-order reactions, $\text{Rate} = k[A]^0$. Because anything raised to the zero power is equal to one, $\text{Rate} = k$. This is another way of saying that the rates of zero-order reactions do not change; they do not speed up and they do not slow down, they either take place or they do not

Grade-11 Chemistry Short Note With Practice Questions

2. The decomposition of the reaction $2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$, is a zero order reactions that occurs on a hot platinum surface . The concentrations N_2O reduced from 0.4M to 0.1M in 20s. Then what is rate constant?
A. 0.015 B. 0.15 C. 0.25 D. -0.015 *ANS A*
3. $A \rightarrow C$ is zero order reaction. If the reaction is 30% complete after 10 min. What is the rate constant for a reaction? If the initial concentration of A was 16M
A. 0.48 B. 48 C. -0.48 D. -48 *ANS A*
- ✓ *Solution After 10min 30% of A is changed in to C. This means 70% of A remains after 10min. Thus $[A]_{20min} = \frac{70}{100} \times 16 = 11.2M$*

$$[A]_{10min} = -Kt + [A]_{20min}$$

$$11.2M = -k \times 10min + 16, \text{ Then calculate for } k = 0.48$$

1.9 First-Order Reactions

- Whose rate of reaction depends on the reactant concentration raised to the first power

$$\ln [A]_t = -kt + \ln [A]_0$$

Integrated rate equation for 1st order reaction
Taking the exponent to each side of the equation:

$$\frac{[A]}{[A]_0} = e^{-kt} \quad \text{or} \quad [A] = [A]_0 e^{-kt}$$

The linear equation $y = mx + b$, in which m is the slope of the line that is the graph of the equation. Thus, a plot of $\ln[A]$ versus t gives a straight line with a slope of k (or m).

EXAMPLE

1. A certain first-order reaction is 45.0% complete in 65 s. What is the rate constant for this process?

If $[A]_0 = 100\%$, then after 65s, 45.0% of A has reacted or $[A]_t = 55\%$ For first order reactions:

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\ln[55] - \ln[100] = -k(65 \text{ s}), \quad k = 9.2 \times 10^{-3} \text{ s}, \text{ use scientific calculator}$$

2. The rate law for the decomposition of phosphine (PH_3) is

$$\text{Rate} = -\frac{\Delta[PH_3]}{\Delta t} = k[PH_3]$$

It takes 120s for 1.00 M PH_3 to decrease to 0.250 M. How much time is required for 2.00 M PH_3 to decrease to a concentration of 0.350 M?

Grade-11 Chemistry Short Note With Practice Questions

$$\ln[\text{PH}_3]_t - \ln[\text{PH}_3]_0 = -kt$$

$$\ln[0.250]_t - \ln[1.00]_0 = -k(120 \text{ s}) \quad k = 0.0116 \text{ s}^{-1}$$

$$\ln[0.350]_t - \ln[2.00]_0 = -0.0116 \text{ s}^{-1}(t) \quad t = 150. \text{ s}$$

3. A sealed rigid container holds two first order reactant gases, A and B. If additional moles of gas B are added to the container while keeping the temperature constant, what can be said of the reaction under these conditions?
- The rate of the reaction will increase, as will the rate constant, k .
 - The rate of the reaction will remain unchanged as will the rate constant, k .
 - The rate of the reaction will increase but the rate constant, k , will remain unchanged.
 - The rate of the reaction will remain unchanged but the rate constant, k , will increase.
 - The rate of reaction will decrease as will the rate constant, k .
- ✓ **ANS C.** The rate will increase due to the concentration increase, but the rate constant will remain the same.
4. The value of rate of a pseudo first order reaction
- Depends on the concentration of reactants present in small amount.
 - Depends on the concentration of reactants present in excess.
 - Is independent of the concentration of reactants.
 - Depend only on one of the reactant
 - B & D
- ANS E
5. The first-order reaction $\text{A} \rightarrow \text{B}$, has $k = 5.67 \text{ s}^{-1}$. If $[\text{A}]_0 = 0.500 \text{ M}$, how long will it take $[\text{A}] = 0.124 \text{ M}$? A. 0.122s B. 0.244s C. 0.100s D. 0.488s
- ANS D

1.10 Second Order Process:



$$\text{Rate}(\text{Ms}^{-1}) = -\frac{d[\text{A}]}{dt} = k[\text{A}]^2$$

Integrating as before we find:

$$\frac{1}{[\text{A}]} = kt + \frac{1}{[\text{A}]_0}$$

A plot of $1/[\text{A}]$ versus t is a straight line with slope k and intercept $1/[\text{A}]_0$

For a second order reaction, a plot of $\ln[\text{A}]$ vs. t is **not linear**.

Example

1. The reaction: $2\text{NOBr}(\text{g}) \rightarrow 2\text{NO}(\text{g}) + \text{Br}_2(\text{g})$ is a second order reaction with respect to NOBr. $k = 0.810 \text{ M}^{-1}\cdot\text{s}^{-1}$ at 10°C . If $[\text{NOBr}]_0 = 7.5 \times 10^{-3} \text{ M}$, how much NOBr will be left after a reaction time of 10 minutes?
- ✓ **SOLUTION:** One can solve for the amount of NOBr after 10 minutes by substituting the given data into the integrated rate law for a second-order reaction.

$$\frac{1}{[\text{NOBr}]_t} = kt + \frac{1}{[\text{NOBr}]_0} \quad (\text{Second Order})$$

Grade-11 Chemistry Short Note With Practice Questions

$$\frac{1}{[\text{NOBr}]_t} = (0.810 \text{ M}^{-1} \cdot \text{s}^{-1}) \times (600 \text{ s}) + \frac{1}{7.5 \times 10^{-3} \text{ M}}$$

$$\frac{1}{[\text{NOBr}]_t} = 6.19 \times 10^2 \text{ M}^{-1}$$

$$[\text{NOBr}]_t = 1.6 \times 10^{-3} \text{ M}$$

2. What are the units for each of the following if the concentrations are expressed in moles per liter and the time in seconds?
- | | |
|--|---|
| A. Rate of a chemical reaction | The units for rate are always mol/L•s. |
| B. Rate constant for a zero-order rate law | Rate = k; k must have units of mol/L•s |
| C. Rate constant for a first-order rate law | Rate = k[A], k must have units of s ⁻¹ |
| D. Rate constant for a second-order rate law | Rate = k[A] ² , k must have units of L/mol•s |
3. Which answer choice correctly matches the integrated rate law shown:

$$1/[A]_t = kt + 1/[A]_0$$

- A. A zero order reaction where a plot of $1/[A]_t$ versus t will be a straight line with slope = k
 B. A first order reaction where a plot of $1/[A]_t$ versus t will be a straight line with slope = k
 C. A first order reaction where a plot of $1/[A]_t$ versus t will be a straight line with slope = -k
 D. A second order reaction where a plot of $1/[A]_t$ versus t will be a straight line with slope = k
 E. A second order reaction where a plot of $1/[A]_t$ versus t will be a straight line with slope = -k

ANS. D

1.11 The Half-life of a Reaction

➤ *Half-life is the time taken for the concentration of a reactant to drop to half its original value*

This means that half life is the time it takes for the concentrations of A to fall from $[A]_0$ to $\frac{1}{2}[A]_0$

HALF-LIFE FOR FIRST- ORDER

$$[A] = \frac{[A]_0}{2}$$

For a first order process the half life ($t_{1/2}$) is found mathematically from

$$(1) \ln[A] = -kt + \ln[A]_0$$

$$(2) \ln[A] - \ln[A]_0 = -kt$$

$$(3) \ln\left(\frac{[A]}{[A]_0}\right) = -kt$$

$$(4) \ln\left(\frac{[A]_0}{[A]}\right) = kt$$

$$(5) \ln\left(\frac{[A]_0}{[A]_0/2}\right) = kt_{1/2}$$

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

Grade-11 Chemistry Short Note With Practice Questions

Example

1. A certain reaction proceeds through first order kinetics. The half-life of the reaction is 180s. What the rate constant?

$$t_{\frac{1}{2}} = \frac{\ln 2}{k} = \frac{0.693}{k} \quad k = \frac{\ln 2}{t_{\frac{1}{2}}} = \frac{\ln 2}{180s} = 0.00385s^{-1}$$

2. A reaction is first order in A. If the rate constant of the reaction is $3.45 \times 10^{-3} s^{-1}$, what is the half-life ($t_{1/2}$) of the reaction?

A. $4.98 \times 10^{-3} s$ B. 100 s C. 201 s D. $3.45 \times 10^{-3} s$ ANS C

3. Iodine -131 has a half life of 8 days. If there are 200g of this sample how much of I-131 will remain after 32days. (the rxn is first order) A. 12.5g B. 25g C. 6.25g D. 50g

✓ **SOLUTION**, One half- life is 8 days total half- life is $32/8 = 4$

~~200~~--8day--~~100~~---8day---~~50~~--- 8day --~~25~~ ---8day--12.5 total of 32 days **Ans. A**

4. If a certain first order rxn has a half-life of 30min how much time it required for this reactions to be 35% complete?

A. 18.7min B. 36min C. 70min D. 40min ANS A

Half-life for a second-order reaction

- Unlike a first order reaction, the rate constant for a second order process depends on the initial concentration of a reactant.

$$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$$

at the half-life, $[A]_t = \frac{1}{2} \times [A]_0$

Substituting and solving, $t_{1/2} = \frac{1}{k[A]_0}$

Example

1. The reaction $2\text{NOBr (g)} \rightarrow 2\text{NO (g)} + \text{Br}_2\text{ (g)}$ is a second order reaction with respect to NOBr. $k = 0.810 \text{ M}^{-1}\text{s}^{-1}$ at 10°C . If $[\text{NOBr}]_0 = 7.5 \times 10^{-3} \text{ M}$, Determine the half-life of this reaction.

✓ To determine the half-life for this reaction, we substitute the initial concentration of NOBr and the rate constant for the reaction into the equation for the half-life of a second-order reaction.

$$t_{1/2} = \frac{1}{k[A]_0} \quad t_{1/2} = \frac{1}{0.810 \text{ M}^{-1} \cdot \text{s}^{-1} (7.5 \times 10^{-3} \text{ M})} = 164s$$

2. If the reaction $2\text{HI} \rightarrow \text{H}_2 + \text{I}_2$ is second order, which of the following will yield a linear plot?

A. $\log [\text{HI}]$ vs time

C. $[\text{HI}]$ vs time

B. $\ln[\text{HI}]$ vs time

D. $1/[\text{HI}]$ vs time ANS. D

3. Nitrosyl chloride, NOCl , decomposes slowly to NO and Cl_2 as shown in the equation below.



The rate constant, k equals 0.020 L/mol.s at a certain temperature.

Grade-11 Chemistry Short Note With Practice Questions

If the initial concentration 0.050 M, what will the concentration be after 30 minutes? Assume the reaction is second order

- A. 0.018M B. 0.18M C. 18M D. 1.8M *ANS A*

1.12 REACTION MECHANISM

- ❖ **A mechanism:** - is a collection of elementary processes (also called elementary steps or elementary reactions) that explains how the overall reaction proceeds.
- ❖ **Elementary steps:** - The individual reactions that make up a reaction mechanism.
- ❖ **Intermediates:** - A substance that is found in elementary steps but not in overall rxns
- ❖ **Molecularity of a reaction:** - The minimum number of reacting particles (molecules, atoms or ions) that come together or collide in a rate determining step to form product or products.

Example

a	$\text{PCl}_5 \rightarrow \text{PCl}_3 + \text{Cl}_2$	(Unimolecular)
b	$2\text{HI} \rightarrow \text{H}_2 + \text{I}_2$	(Bimolecular)
c	$2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$	(Termolecular)
d	$\text{NO} + \text{O}_3 \rightarrow \text{NO}_2 + \text{O}_2$	(Bimolecular)
e	$2\text{CO} + \text{O}_2 \rightarrow 2\text{CO}_2$	(Termolecular)
f	$2\text{FeCl}_3 + \text{SnCl}_2 \rightarrow \text{SnCl}_2 + 2\text{FeCl}_2$	(Termolecular)

- ❖ **Rate determining step:** - The slowest step in the sequence of steps leading to the formation of products. In other words, the slowest step in the mechanism determines the overall rate of the reaction.

NOTE: - In elementary steps, the reaction order, Molecularity and Stoichiometric Coefficient are numerically the same also they have different concepts except pseudo order rxns

Molecularity: - Cannot be zero, negative, fractions and infinite

Example

1. The reaction $\text{NO}_2(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{CO}_2(\text{g})$ is thought to occur by the following mechanism: the experimentally rate law is $\text{Rate} = k[\text{NO}_2]^2$

Step 1: $\text{NO}_2(\text{g}) + \text{NO}_2(\text{g}) \rightarrow \text{NO}_3(\text{g}) + \text{NO}(\text{g})$ Slow

Step 2: $\text{NO}_3(\text{g}) + \text{CO}(\text{g}) \rightarrow \text{NO}_2(\text{g}) + \text{CO}_2(\text{g})$ Fast

Check whether the proposed reaction mechanism is acceptable or unacceptable

1. The sum of these reaction steps must give the overall balanced equation:
2. Mechanism must agree with the experimental rate law

For the above equation

The first requirement is full filled because the sum of overall reaction is correct

The second requirement is full filled because the rate from slowest or rate determining step is agreed with experimentally rate law. So the proposed mechanism is acceptable

3. The proposed steps for a catalyzed reaction between X^{4+} and Z^+ are represented below.

step 1: $\text{X}^{4+} + \text{Y}^{2+} \rightarrow \text{X}^{3+} + \text{Y}^{3+}$

step 2: $\text{X}^{4+} + \text{Y}^{3+} \rightarrow \text{X}^{3+} + \text{Y}^{4+}$

step 3: $\text{Y}^{4+} + \text{Z}^+ \rightarrow \text{Z}^{3+} + \text{Y}^{2+}$

The catalyst in this process is

- A. X^{4+} B. X^{3+} C. Y^{4+} D. Y^{2+}

Grade-11 Chemistry Short Note With Practice Questions

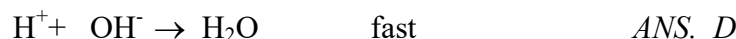
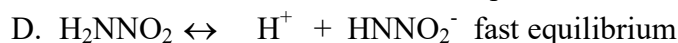
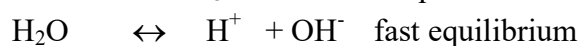
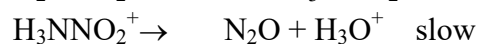
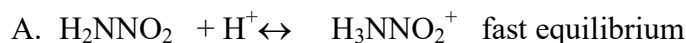
✓ *ANS. D.* A catalyst is a chemical that gets used up during an early step, but returns as a product during a later step

4. From question number 3 which species is intermediate? *ANS. Y^{4+} , Y^{3+}*

Because Intermediates are produced in one step and then consumed in a later step

5. The rate equation for the decomposition of nitramide, $H_2NNO_2 \longrightarrow N_2O + H_2O$, is

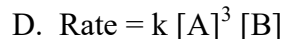
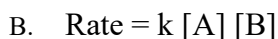
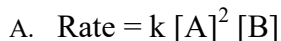
Rate = $k [H_2NNO_2][H^+]^{-1}$. Which of the following mechanisms is consistent with this rate equation?



6. The proposed mechanism for a given chemical reaction is shown below.

Step 1: $A + B \rightleftharpoons F$	(fast)
Step 2: $F + A \rightarrow C + G$	(slow)
Step 3: $G + A \rightleftharpoons C + D$	(fast)

Which is the correct rate law?



✓ *ANS. A* Since the second step is the slow step, cross off the intermediate, and gather all the reactants in the first and second steps.

7. A proposed mechanism for the following reaction, $H_2 + I_2 \rightarrow 2HI$ is

$I_2 \leftrightarrow 2I$	fast, equilibrium
$I + H_2 \leftrightarrow H_2I$	fast, equilibrium
$H_2I + I \rightarrow 2HI$	slow

A. Show that this mechanism predicts the correct rate law, rate = $k[H_2][I_2]$.

From step 3 (the slow, rate-determining step):



Grade-11 Chemistry Short Note With Practice Questions

$$\text{rearranging, } [I] = \sqrt{\frac{k_{1f}[I_2]}{k_{1r}}}$$

From step 2, $\text{Rate}_{2f} = \text{Rate}_{2r}$ or $k_{2r}[I][H_2] = k_{1r}[IH_2]$

$$\text{rearranging, } [IH_2] = \frac{k_{2f}[I][H_2]}{k_{2r}}$$

Substituting the expression for [I] from step 1 into this latter expression,

$$[IH_2] = \frac{k_{2f} \sqrt{\frac{k_{1f}[I_2]}{k_{1r}}} [H_2]}{k_{2r}}$$

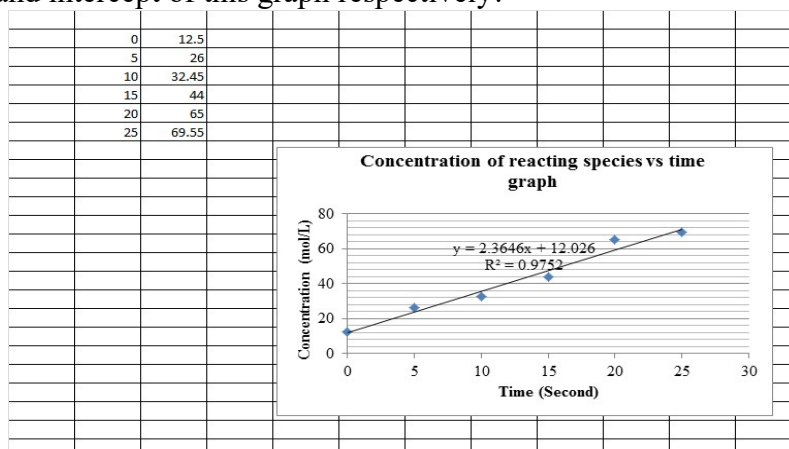
Substituting this expression for [I] and $[IH_2]$ into the rate law deduced from step 1, and the combining terms,

$$\text{Rate} = k_3 \times \frac{k_{2f} \sqrt{\frac{k_{1f}[I_2]}{k_{1r}}} [H_2]}{k_{2r}} \times \sqrt{\frac{k_{1f}[I_2]}{k_{1r}}} = \frac{k_3 k_{2f} k_{1f}}{k_{2r} k_{1r}} [H_2][I_2]$$

$$\text{Rate} = k[H_2][I_2]$$

The mechanism is **consistent** with the observed rate law.

8. Each of the choices below gives an elementary reaction. Which one could **not** be true regarding the molecularity of individual step?
- A. $A + B + C \rightarrow P$: Trimolecular C. $2A \rightarrow P$: Bimolecular
 B. $A \rightarrow P$: Unimolecular D. $A + B \rightarrow P$: Unimolecular Ans. D
9. Which of the following statements are applicable to a balanced chemical equation of an elementary reaction?
- A. Order is same as molecularity.
 B. Order is less than the molecularity.
 C. Order is greater than the molecularity. Ans. A.
 D. Molecularity can never be positive value
10. The kinetics of the reaction below were studied and it was determined that the reaction rate did not change when the concentration of B was tripled. The reaction is ____ order in B:
 ($A + B \rightarrow P$)
 A. First B. zero C. second D. third Asn. B.
11. **Refer the fig. below.** Which order of the reaction would be represented by this graph, what is the slope and intercept of this graph respectively?

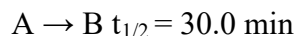


- A. First order, 2.3646 and 0.9752 C. Third order, -2.3646 and -12.026
 B. Second order, 2.3646 and 12.026 D. Zero order, 0.9752 & 2.3646

Ans. B

Grade-11 Chemistry Short Note With Practice Questions

12. For the first-order reactions of different substances A and X



- A. Doubling the concentration of A will have $\frac{1}{2}$ the effect on half-life that doubling the concentration of B will have on its half-life.
- B. A certain number of grams of A will react twice as fast as the same number of grams of X.
- C. A certain number of grams of X will react twice as fast as the same number of grams of A.
- D. The rate constant for $A \rightarrow B$ is lower than the rate constant of $X \rightarrow Y$.
- E. 3 moles of A will react more rapidly than 3 moles of X

✓ Answer: (E) It takes half as much time for A to form B as for X to form Y, as seen by the smaller half-life. Note that option "B" would be incorrect as the grams of A and the grams of X are not the same number of moles.

13. A reaction is first order with respect to [X] and second order with respect to [Y]. When [X] is 0.20 M and [Y] = 0.20 M the rate is $8.00 \times 10^{-3} \text{ M/min}$. The value of the rate constant, including correct units, is

- A. 1.00 M min^{-1} B. $1.00 \text{ M}^{-2} \text{ min}^{-1}$ C. $2.00 \text{ M}^{-1} \text{ min}^{-1}$ D. $2.00 \text{ M}^{-2} \text{ min}^{-1}$

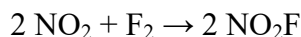
✓ Answer: (B) From these data, it follows that the rate law is $\text{Rate} = k[X]Y^2$. Solving for the rate constant and substituting data for this reaction: $K = \text{Rate}/[X][Y]^2$
 $= 8.00 \times 10^{-3} \text{ M/min}/(0.200 \text{ M})(0.200 \text{ M})^2$
 $= 0.008 \text{ M/min}/0.008 \text{ M}^3 = 1.00 \text{ M}^{-2}\text{min}^{-1}$

14. A reaction mechanism

- A. Is the sum of all steps in a reaction except the rate determining step.
- B. Has a ΔH equal to the ΔH for the most demanding step.
- C. Always has a rate determining step.
- D. May be absolutely proven from the rate law.
- E. Is determined from the balanced expression only.

✓ Answer: (C) In all series of steps (the mechanism for the overall reaction), there must always be a slowest step. The rate determining step is the slowest step.

15. The balanced equation for the reaction of nitrogen dioxide and fluorine is



The proposed mechanism is

step 1: $\text{NO}_2 + \text{F}_2 \rightarrow \text{NO}_2\text{F} + \text{F}$	(slow)
step 2: $\text{F} + \text{NO}_2 \rightarrow \text{NO}_2\text{F}$	(fast)

Which of the following are correct?

- I. The mechanism supports an experimentally determined rate law of $\text{rate} = k[\text{NO}_2]^2[\text{F}_2]$
 - II. F is an intermediate
 - III. The reaction is first order in F_2
- A. I only B. I and II only C. I and III only D. II and III

✓ ANS. D Since the first step is the slow step, the stoichiometry of that step results in the order of the reaction, thus it is first order for both NO_2 and F_2 . Also, F is an intermediate because it is a product of one step, but gets used up in a subsequent step.

Grade-11 Chemistry Short Note With Practice Questions

16. If tripling the concentration of a single reactant R in a multi-reactant reaction (while leaving all other conditions unchanged) leads to a nine-fold increase in rate, it can be deduced
- A. That the complete rate law is $\text{Rate} = k[\text{R}]$
 - B. That the complete rate law is $\text{Rate} = k[\text{R}]^2$
 - C. That the complete rate law is $\text{Rate} = k[\text{R}]^3$
 - D. Only that the order with respect to [R] is 2 ASN. D
- ✓ We cannot know the entire rate law, only that the order is second with respect to the reactant
17. The greatest increase in the rate for the reaction between X and Z with rate $r = k[\text{X}][\text{Z}]^2$ will be caused by:
- A. Doubling the concentration of Z
 - C. Doubling the concentration of X
 - B. Tripling the concentration of X
 - D. Lowering the temperature ASN. A
18. If a reaction proceeding by the mechanism $\text{A} + \text{B} \rightarrow \text{C} + \text{D}$ occurs at a rate x , and if the concentrations of A and B are both doubled, what will be the new rate of reaction?
- A. x
 - B. $2x$
 - C. $4x$
 - D. $8x$
 - E. $16x$ ANS. C
19. Consider the rate law: $\text{rate} = k[\text{Y}]^m[\text{Z}]^n$. How are the exponents m and n determined?
- A. By using the balanced chemical equation
 - B. By using the subscripts of the chemical formulas
 - C. By using the coefficients of the chemical formulas
 - D. By educated guess
 - E. By experiment ASN. E
20. For a reaction the time required for the 20% completion is 40 min and 40% completion is 80 min what is the order of the reaction?
- A. 1
 - B. 2
 - C. 3
 - D. 0
21. Which of the following statement is true about order and molecularity?
- A. Order is defined for both elementary as well as complex reactions
 - B. Molecularity is defined for both elementary as well as complex reactions
 - C. Order cannot be zero, negative or fractions
 - D. Molecularity can be zero, negative or fractions ASN. A
22. Which of the following statement is not true?
- A. Bimolecular, involves the collision of two particles
 - B. Termolecular, involves the simultaneous collision of three particles all at one time.
 - C. Unimolecular, involves the collision of a single particle
 - D. Termolecular collision is easier than bimolecular collision

Ans C. Unimolecular is decomposition or rearrangement of a particle not only collision

UNIT 5

Chemical Equilibrium and Phase Equilibrium

Main topics

- Chemical Equilibrium
- Phase Equilibrium

5.1 Chemical Equilibrium

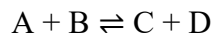
Having looked at factors that affect the rate of a reaction, we now need to ask some important questions. Does a reaction always proceed in the same direction or can it be reversible? In other words, is it always true that a reaction proceeds from reactants to products, or is it possible that sometimes, the reaction will reverse and the products will be changed back into the reactants? And does a reaction always run its full course so that all the reactants are used up, or can a reaction reach a point where reactants are still present, but there does not seem to be any further change taking place in the reaction?

Reversible and Irreversible Reactions

Reversible reactions

The reactions in which the products formed during the reaction combine to give back the reactants are called **reversible reactions**.

It is indicated by writing two half headed arrow marks between the reactants and product (\rightleftharpoons)



Question: what do arrow marks indicate?

Answer: the arrow head pointing to the right represents the forward reaction and the arrow head pointing to the left represents the backward reaction.

Conditions for a reversible reaction:

- The reaction should be carried out in a closed vessel.
- The products should not be removed from the vessel.
- Temperature and pressure should be kept constant.

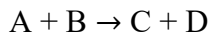
Examples for reversible reactions

- $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$
- $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
- $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$

The reactions in which products formed do not combine to give back the reactants are called **irreversible reactions**.

- ✓ The irreversible reactions proceed only in one direction.
- ✓ They are indicated by writing an arrow between reactants and products as:

Grade-11 Chemistry Short Note With Practice Questions



Question: can irreversible reactions be made reversible ?

Answer: an irreversible reaction can not be made reversible even if it is carried out in a closed vessel.

Examples for irreversible reactions:

- a. $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2 \uparrow$ c. $Mg + 2HCl \rightarrow MgCl_2 + H_2 \uparrow$
b. $AgNO_3 + KCl \rightarrow AgCl \downarrow + KNO_3$

Attainment and characteristics of chemical equilibrium

Equilibrium is a state in which there are no observable changes as time goes by.

Question: what is dynamic chemical equilibrium?

Answer: any reaction will be in dynamic equilibrium if it's reversible and the rates of the forward and reverse reactions are equal.

An example of dynamic equilibrium is: $NO_2(g) + CO(g) \rightleftharpoons NO(g) + CO_2(g)$.

Nitrogen dioxide (NO_2) reacts with carbon monoxide (CO) to form nitrogen oxide (NO) and carbon dioxide (CO_2), and, in the reverse reaction, nitrogen oxide and carbon dioxide react to form nitrogen dioxide and carbon monoxide.

Another example of dynamic equilibrium is, say that you prepare a solution that is saturated with an aqueous solution of $NaCl$. If you then add solid crystals of $NaCl$, the $NaCl$ will be simultaneously dissolving and recrystallizing within the solution.

The process, $NaCl(s) \rightleftharpoons Na^+(aq) + Cl^-(aq)$, will be in dynamic equilibrium when the rate of the dissolution of the $NaCl$ equals the rate of recrystallization.

Below is a table showing the key differences between dynamic and static equilibrium.

Dynamic Equilibrium	Static Equilibrium
➤ Reversible	➤ Irreversible
➤ Reaction is still occurring	➤ Reaction has stopped
➤ Rate of forward rxn = Rate of reverse rxn	➤ Both rxn rates are zero
➤ Occurs in a closed system	➤ Can occur in an open or closed system

Chemical equilibrium

When a chemical reaction has reached the equilibrium state, the concentrations of reactants and products remain constant over time, and there are no visible changes in the system. However,

Grade-11 Chemistry Short Note With Practice Questions

there is much activity at the molecular level because reactant molecules continue to form product molecules while product molecules react to yield reactant molecules, it is a dynamic situation.

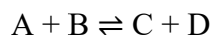
Chemical equilibrium is the state of the reaction when the macroscopic properties like temperature, pressure, volume and concentration of the reaction do not change with time.

Conditions for attainment of chemical equilibrium

Analogous to the physical systems chemical reactions also attain a state of equilibrium.

- The reactions can occur both in forward and backward directions.
- The rates of the forward and reverse reactions become equal
- The concentrations of the reactants and the products remain constant.
- The equilibrium is dynamic in nature as it consists of a forward reaction in which the reactants give product(s) and reverse reaction in which product(s) gives the original reactants

For a better comprehension, let us consider a general case of a reversible reaction,



With passage of time, there is accumulation of the products C and D and depletion of the reactants A and B.

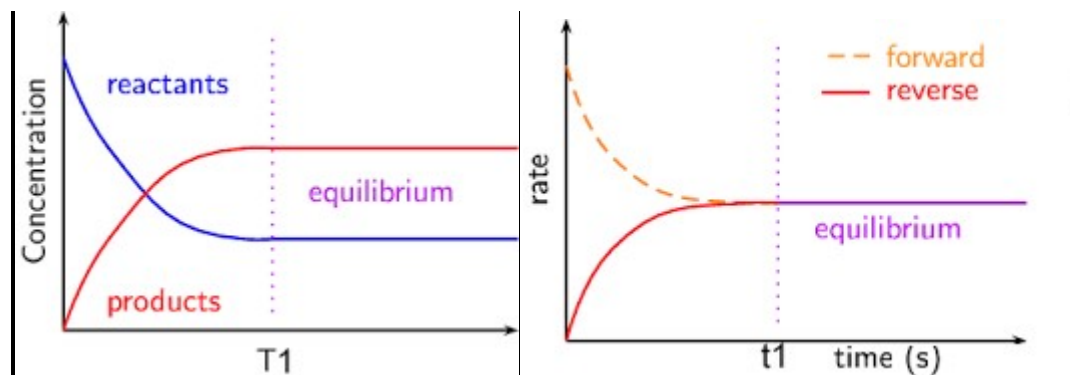


Fig.

This leads to a decrease in the rate of forward reaction and an increase in the rate of the reverse reaction, eventually, the two reactions occur at the same rate and the system reaches a state of equilibrium. We know from "Factors affecting reaction rates" the rate of the forward reaction is directly proportional to the concentration of the reactants. In other words, as the concentration of the reactants increases, so does the rate of the forward reaction. This can be shown using the following equation:

Rate of forward reaction $\propto [A][B]$ or Rate of forward reaction $= k_1[A][B]$. Similarly,

Grade-11 Chemistry Short Note With Practice Questions

Rate of the reverse reaction is directly proportional to the concentration of the products. This can be shown using the following equation:

Rate of reverse reaction $\propto [C][D]$ or Rate of reverse reaction $= k_2[C][D]$

At equilibrium, the rate of the forward reaction is equal to the rate of the reverse reaction. This can be shown using the following equation:

$$k_1 [A][B] = k_2 [C][D] \text{ or}$$
$$\frac{k_1}{k_2} = \frac{[C][D]}{[A][B]}$$

or, if the constants k_1 and k_2 are simplified to a single constant, the equation becomes:

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

A more general form of the equation for a reaction at chemical equilibrium is:



$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

It is important to note that if a reactant or a product in a chemical reaction is in either the liquid or solid phase, the concentration stays constant during the reaction. Therefore, these values can be left out of the equation to calculate K_c .

For example, in the following reaction: $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$

$$K_c = \frac{[CO][H_2]}{[H_2O]}$$

Characteristics of chemical equilibrium

- Equilibrium is achieved in a closed system. A closed system prevents exchange of matter with the surroundings, so equilibrium is achieved where both reactants and products can react and recombine with each other.
- The rate of the forward reaction is equal to the rate of the backward reaction.
- Equilibrium is dynamic. The reaction hasn't stopped but both forward and backward reactions are still occurring.
- The concentrations of reactants and products remain constant at equilibrium. They are being produced and destroyed at an equal rate.
- At equilibrium there's no change in macroscopic properties.
- Equilibrium can be reached from either direction.
- The free energy change of the reaction is zero at equilibrium

Equilibrium expression and equilibrium constant

The value of the equilibrium constant expression, K_c , is constant for a given reaction at equilibrium and at a constant temperature.

Grade-11 Chemistry Short Note With Practice Questions

The equilibrium constant for a general reaction, $aA + bB \rightleftharpoons cC + dD$ is expressed as,

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Where $[A]$, $[B]$, $[C]$ and $[D]$ are the equilibrium concentrations of the reactants and products. Since pressure is proportional to concentration for gases in a closed system, the equilibrium expression can also be written as, $PV = nRT, \Rightarrow P = \frac{n}{V} RT$. Here, P is the pressure in Pa, n is the number of moles of the gas, V is the volume in m^3 and T is the temperature in Kelvin. Therefore, n/V is concentration expressed in mol/m^3 . If concentration c , is in mol/L or mol/dm^3 , and P is in bar then $P = cRT$. We can also write $p = [gas]RT$. Here, $R = 0.0831$ bar litre/mol K. At constant temperature, the pressure of the gas is proportional to its concentration i.e., $P \propto [gas]$.

For a general reaction; $aA + bB \rightleftharpoons cC + dD$

$$K_p = \frac{P_C^c \times P_D^d}{P_A^a \times P_B^b}$$

Example 1: Write K_p expression for reaction below $A + 2B \rightleftharpoons 3C$

$$K_p = P_C^3 / (P_A \times P_B^2)$$

Example 2: Consider: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$

$$K_p = \frac{(P_{HI})^2}{(P_{H_2})(P_{I_2})}, \quad K_p = K_c(RT)^{\Delta n},$$

$\Delta n = \text{mol gas phase product} - \text{mol gas phase reactant}$

$$K_p = K_c \text{ when } \Delta n = 0$$

Exercise: CH_4 reacts with H_2S to form H_2 and CS_2 : $CH_4(g) + 2H_2S(g) \rightleftharpoons CS_2(g) + 4H_2$

(g). Determine the value of K_p for this reaction at 1000 K if the following equilibrium pressures were measured: $P_{CH_4} = 0.20$ atm, $P_{H_2S} = 0.25$ atm, $P_{CS_2} = 0.52$ atm and $P_{H_2} = 0.10$ atm. Determine the value of K_c for this reaction at 1000 K.

Relationship between K_c and K_p

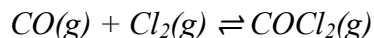
From the Ideal Gas Law we know that: $PV = nRT$ and $P = (n/V)RT = cRT$. Plugging this into the expression for K_p for each substance, the relationship between K_c and K_p becomes:

$$K_p = K_c(RT)^{\Delta n}, \text{ where}$$

$$\Delta n = (\text{moles of gaseous product}) - (\text{moles of gaseous reactant})$$

Example 1: The equilibrium concentrations for the reaction between carbon monoxide and molecular chlorine to form $COCl_2(g)$ at $74^\circ C$ are $[CO] = 0.012$ M, $[Cl_2] = 0.054$ M, and $[COCl_2] = 0.14$ M. Calculate the equilibrium constants K_c and K_p .

Grade-11 Chemistry Short Note With Practice Questions

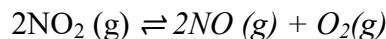


$$K_c = \frac{[\text{COCl}_2]}{[\text{CO}][\text{Cl}_2]^2} = \frac{0.14}{0.012 \times 0.054} = 220,$$

$$K_p = K_c(RT)^{\Delta n}, \Delta n = 1 - 2 = -1, R = 0.0821, T = 273 + 74 = 347 \text{ K}$$

$$K_p = 220 \times (0.0821 \times 347)^{-1} = 7.7$$

Example 2: The equilibrium constant K_p for the reaction given below is 158 at 1000K. What is the equilibrium pressure of O_2 if the $P_{\text{NO}} = 0.400 \text{ atm}$ and $P_{\text{NO}_2} = 0.270 \text{ atm}$?



$$K_p = \frac{(P_{\text{NO}})^2(P_{\text{O}_2})}{(P_{\text{NO}_2})^2}, \text{ this implies } P_{\text{O}_2} = K_p \times \frac{(P_{\text{NO}})^2}{(P_{\text{NO}_2})^2} = 158 \times \frac{(0.400)^2}{(0.270)^2} = 347 \text{ atm}$$

Example 3: For the equilibrium, $2\text{NOCl(g)} \rightleftharpoons 2\text{NO(g)} + \text{Cl}_2\text{(g)}$ the value of the equilibrium constant, K_c is 3.75×10^{-6} at 1069 K. Calculate the K_p for the reaction at this temperature.

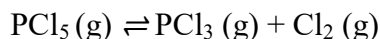
Solution

we know that, $K_p = K_c(RT)^{\Delta n}$

For the above reaction, $\Delta n = (2+1) - 2 = 1, K_p = 3.75 \times 10^{-6} (0.0831 \times 1069)$

$$K_p = 0.033$$

Exercise: The equilibrium constant K_p for the decomposition of phosphorus pentachloride to phosphorus trichloride and molecular chlorine is found to be 1.05 at 250°C . If the equilibrium partial pressures of PCl_5 and PCl_3 are 0.875 atm and 0.463 atm, respectively, what is the equilibrium partial pressure of Cl_2 at 250°C ?



Homogeneous equilibria: Occur when all reactants and products are in the same phase.

Examples:



Heterogeneous equilibria: Occur when reactant or product in the equilibrium is in a different phase.

Examples;



Concentrations of pure solids and liquids are fixed by their density and molar mass (both constants) and do not vary with the amount. The value used for the concentration of a pure substance is always 1. Therefore, the concentrations of solids and liquids do not appear in the equilibrium expression.

- Thus, Omit concentration terms for solids and liquids from K_c and K_p expressions; only include terms for gases (g) and aqueous substances (aq).

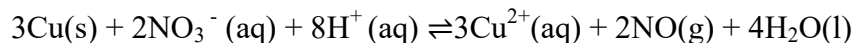
Grade-11 Chemistry Short Note With Practice Questions

Example 1: Write the K_c expression for $\text{CaCO}_3(s) \rightleftharpoons \text{CaO}(s) + \text{CO}_2(g)$. $K_c = [\text{CO}_2]$

Example 2: $\text{PbCl}_2(s) \rightleftharpoons \text{Pb}^{2+}(aq) + 2 \text{Cl}^-(aq)$

$$K_c = [\text{Pb}^{2+}][\text{Cl}^-]^2$$

Example 3: Write the K_c expression for the following reaction:

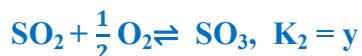


$$K_c = \frac{[\text{Cu}^{2+}]^3 [\text{NO}]^2}{[\text{NO}_3^-]^2 [\text{H}^+]^8}$$

Modifying Equilibrium Constant Expressions:

Example: $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$, $K_1 = x$

A. Changing stoichiometric coefficients:



The equilibrium constant of a reaction that has been multiplied by a number, is the equilibrium constant raised to a power that is equal to that number. Therefore,

$$K_2 = (K_1)^{1/2}, y = (x)^{1/2}$$

B. Reversing the reaction:



The equilibrium constant of a reaction in the reverse direction is the reciprocal of the equilibrium constant of the forward reaction. Therefore,

$$K_3 = 1/K_1, z = 1/x$$

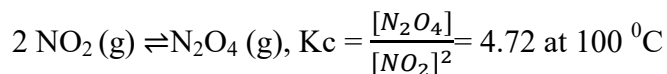
C. Adding equations for multiple equilibrium reactions:



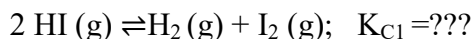
The equilibrium constant for a net reaction made up of two or more steps is the product of the equilibrium constants for the individual steps. Therefore,

$$K_3 = K_1 \times K_2 = xy = z$$

Example 1: $\text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g)$, $K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.212$ at 100°C



Example 2: Consider, $\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)$; $K_C = 54$ and:



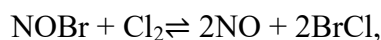
$$K_c = \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} = 54; K_{C1} = \frac{[\text{H}_2][\text{I}_2]}{[\text{HI}]^2} = (K_C)^{-1} = 0.018$$

Grade-11 Chemistry Short Note With Practice Questions

Example 3: $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g}), K_c = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 0.212 \text{ at } 100^\circ\text{C}$

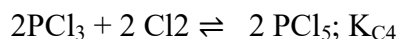
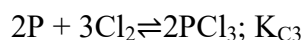
$$\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 4\text{NO}_2(\text{g}), K_c = \frac{[\text{NO}_2]^4}{[\text{N}_2\text{O}_4]^2} = (0.212)^2 \text{ at } 100^\circ\text{C}$$

Example 4: $2\text{NOBr} \rightleftharpoons 2\text{NO} + \text{Br}_2, K_1 = 0.014$



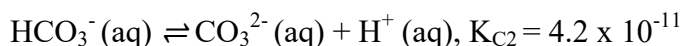
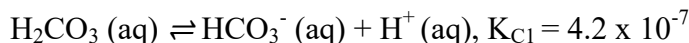
$$K_3 = K_1 \times K_2 = 0.014 \times 7.2 = 0.10$$

Example 5: Consider combining the following equations:



$$K_c = K_{C3} \cdot K_{C4}; K_c = \frac{[\text{PCl}_5]^2}{[\text{P}]^2[\text{Cl}_2]^5} = \frac{[\text{PCl}_3]^2}{[\text{P}]^2[\text{Cl}_2]^3} \times \frac{[\text{PCl}_5]^2}{[\text{PCl}_3]^2[\text{Cl}_2]^2}$$

Example 6: The following equilibrium constants have been determined for carbonic acid at 25°C :



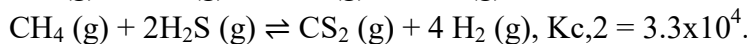
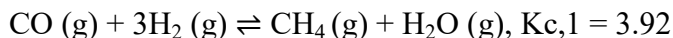
The overall reaction is the sum of the two reactions: $\text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons 2\text{H}^+(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$

$$K_c = K_{C1} \times K_{C2} = (4.2 \times 10^{-7})(4.2 \times 10^{-11}) = 1.8 \times 10^{-17}$$

Exercises:



b. Consider the following reactions at 1200 K.



Use the above reactions to determine the equilibrium constant (K_c) for the following reaction at 1200 K.



Table: Relations between equilibrium constants for a general reaction and its multiples.

Chemical equation	Equilibrium constant
$aA + bB \rightleftharpoons cC + dD$	K_c
$cC + dD \rightleftharpoons aA + bB$	$K_{C1} = (1/K_c) = (K_c)^{-1}$

Grade-11 Chemistry Short Note With Practice Questions

$naA + nbB \rightleftharpoons ncC + ndD$	$K_{C2} = (K_c)^n$
sum of 2 reactions with K_{C3} & K_{C4}	$K_C = (K_{C3})(K_{C4})$

The relationship between chemical kinetics and chemical equilibrium

For the reaction: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$

Rate of forward reaction = Rate of reverse reaction

$$k_f[N_2O_4] = k_r[NO_2]^2$$

Rearrange; $K_C = \frac{[NO_2]^2}{[N_2O_4]} = \frac{K_f}{K_r}$

Thus, the equilibrium constant is simply the ratio of the forward and reverse rate constants which are both constant values at a given temperature.

- ❖ If $k_f > k_r$, $K_C > 1$, forward reaction is favored equilibrium position lies to the right.
- ❖ If $k_f < k_r$, $K_C < 1$ reverse reaction is favored equilibrium position lies to the left

Application of equilibrium constant

What does the equilibrium constant tell us?

- ✓ If $K_C \gg 1$, the reaction is product-favored; product predominates at equilibrium.
- ✓ If $K_C \ll 1$, the reaction is reactant-favored; reactant predominates at equilibrium.

Predicting the Extent of a Reactions

The numerical value of the equilibrium constant for a reaction indicates the extent of the reaction. But it is important to note that equilibrium constant does not give any information about the rate at which the equilibrium is reached. The magnitude of K_c or K_p is directly proportional to the concentrations of products (as these appear in the numerator of equilibrium constant expression) and inversely proportional to the concentrations of the reactants (these appear in the denominator). This implies that a high value of K is suggestive of a high concentration of products and vice-versa.

We can make the following generalizations concerning the composition of equilibrium mixtures:

- If $K_C > 10^3$, products predominate over reactants, i.e., if K_C is very large, the reaction proceeds nearly to completion.
Consider the following examples:
 - The reaction of H_2 with O_2 at 500 K has a very large equilibrium constant, $K_C = 2.4 \times 10^{47}$.
 - $H_2(g) + Cl_2(g) \rightleftharpoons 2HCl(g)$ at 300K has $K_C = 4.0 \times 10^{31}$.
 - $H_2(g) + Br_2(g) \rightleftharpoons 2HBr(g)$ at 300 K, $K_C = 5.4 \times 10^{18}$.
- If $K_C < 10^{-3}$, reactants predominate over products, i.e., if K_C is very small, the reaction proceeds rarely.

Consider the following examples:

- The decomposition of H_2O into H_2 and O_2 at 500 K has a very small equilibrium constant, $K_C = 4.1 \times 10^{-48}$.
- $N_2(g) + O_2(g) \rightleftharpoons 2NO(g)$, at 298 K has $K_C = 4.8 \times 10^{-31}$.

Grade-11 Chemistry Short Note With Practice Questions

Note: If K_c is in the range of 10^{-3} to 10^3 , appreciable concentrations of both reactants and products are present. Consider the following examples:

- For reaction of H_2 with I_2 to give HI , $K_c = 57.0$ at $700K$.
- Also, gas phase decomposition of N_2O_4 to NO_2 is another reaction with a value of $K_c = 4.64 \times 10^{-3}$ at $25^\circ C$ which is neither too small nor too large.

Hence, equilibrium mixtures contain appreciable concentrations of both N_2O_4 and NO_2 .

Predicting the direction of reaction:

Equilibrium constant can be used to predict the direction of net reaction. For a reaction of known K_c value, the direction of net reaction can be predicted by calculating the reaction quotient, Q_c .

Q_c is called the reaction quotient, where for a reaction such as: $aA + bB \rightleftharpoons cC + dD$

$$Q_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

Q_c has the same expression as K_c , but it is calculated using concentrations that are not necessarily at equilibrium.

- If $Q_c < K_c$, the reaction will go to the right.

The ratio of products over reactants is too small & the reaction will move toward equilibrium by forming more products.

- If $Q_c = K_c$, the reaction mixture is already at equilibrium, so no shift occurs.

Example1: At $448^\circ C$, $K = 51$ for the reaction: $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$. Predict the direction the reaction will proceed, if at $448^\circ C$ the pressures of HI , H_2 , and I_2 are 1.3, 2.1 and 1.7 atm., respectively.

Solution: $Q = \frac{[P_{HI}]^2}{[P_{H_2}] [P_{I_2}]} = \frac{(1.3)^2}{(2.1)(1.7)} = 0.47$

$0.47 < 51$, imply system are not at equilibrium. Numerator (products) must increase and denominator (reactants) must decrease. Consequently reaction must shift to the right.

Example2: Consider the gaseous reaction of H_2 with I_2 , $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$; $K_c = 57.0$ at $700 K$. Suppose we have molar concentrations $[H_2]_t = 0.10M$, $[I_2]_t = 0.20 M$ and $[HI]_t = 0.40 M$. (the subscript t on the concentration symbols means that the concentrations were measured at some arbitrary time t, not necessarily at equilibrium).

Solution: Thus, the reaction quotient, Q_c at this stage of the reaction is given by:

$$Q_c = [HI]_t^2 / [H_2]_t [I_2]_t = (0.40)^2 / (0.10)(0.20) = 8.0$$

Now, in this case, Q_c (8.0) does not equal K_c (57.0), so the mixture of $H_2(g)$, $I_2(g)$ and $HI(g)$ is not at equilibrium; that is, more $H_2(g)$ and $I_2(g)$ will react to form more $HI(g)$ and their concentrations will decrease till $Q_c = K_c$

Grade-11 Chemistry Short Note With Practice Questions

Example 3: The value of K_c for the reaction $2A \rightleftharpoons B + C$ is 2×10^{-3} . At a given time, the composition of reaction mixture is $[A] = [B] = [C] = 3 \times 10^{-4} \text{ M}$. In which direction the reaction will proceed?

Solution: For the reaction, the reaction quotient Q_c is given by, $Q_c = [B][C] / [A]^2$ as $[A] = [B] = [C] = 3 \times 10^{-4} \text{ M}$, $Q_c = (3 \times 10^{-4})(3 \times 10^{-4}) / (3 \times 10^{-4})^2 = 1$ as $Q_c > K_c$ so the reaction will proceed in the reverse direction.

Exercises: For the reaction, $B \rightleftharpoons 2A$, $K_c = 2$.

Suppose 3.0 moles of A and 3.0 moles of B are introduced into a 2.00 L flask.

- In which direction will the reaction proceed to attain equilibrium?
- Will the concentration of B increase, decrease or remain the same as the system moves towards equilibrium?

Calculating Equilibrium Concentrations

In case of a problem in which we know the initial concentrations but do not know any of the equilibrium concentrations, the following three steps shall be followed:

Step1. Write the balanced equation for the reaction.

Step2. Under the balanced equation, make a table that lists for each substance involved in the reaction:

- the initial concentration,
- the change in concentration on going to equilibrium, and
- the equilibrium concentration.

In constructing the table, define x as the concentration (mol/L) of one of the substances that reacts on going to equilibrium, then use the stoichiometry of the reaction to determine the concentrations of the other substances in terms of x .

Step3. Substitute the equilibrium concentrations into the equilibrium equation for the reaction and solve for x . If you are to solve a quadratic equation choose the mathematical solution that makes chemical sense.

Step4. Calculate the equilibrium concentrations from the calculated value of x .

Step5. Check your results by substituting them into the equilibrium equation.

Example1. Use K to find conc. of reactants and products at equilibrium.

Given $K = 2.7 \times 10^{-7}$ and initial amounts of $[A] = 1.0$ $[B] = 0$ and $[C] = 0$ then find concentrations at equilibrium for reaction $2A \rightleftharpoons B + 3C$

- Make table of initial, change, and final in symbolic form

	$2A$	\rightleftharpoons	B	+	$3C$
Initial	1.00		0		0
Change	-2x		+x		+3x
Equilibrium	$1.00-2x$		x		3x

- Write Equilibrium expression and substitute in symbols

$$K_c = [B][C]^3/[A]^2 = [x] [3x]^3/[1-2x]^2$$

Grade-11 Chemistry Short Note With Practice Questions

c. Substitute in for K and solve math exactly or simplify to approximate

$$2.7 \times 10^{-7} = (27)(x^4) / [1-2x]^2$$

Since K_c is small then x is small so approximate $1-2x \sim 1$

$$2.7 \times 10^{-7} = (27)(x^4) / [1-2x]^2, \quad 1.0 \times 10^{-8} = x^4, \quad (1.0 \times 10^{-8})^{1/4} = (x^4)^{1/4}, \quad 1.0 \times 10^{-2} = x$$

Example2: $\text{SO}_2(\text{g}) + \text{NO}_2(\text{g}) \rightleftharpoons \text{SO}_3(\text{g}) + \text{NO}(\text{g}); K_c = 3.75$

1.24 mol $\text{SO}_2(\text{g})$ and 0.750 mol $\text{NO}_2(\text{g})$ are combined in a 2.50 L flask, and the reaction is allowed to reach equilibrium. Determine the concentration of each species present at equilibrium.

	$\text{SO}_2(\text{g})$	+	$\text{NO}_2(\text{g})$	\rightleftharpoons	$\text{NO}(\text{g})$	+	$\text{SO}_3(\text{g})$
initial []	0.496 M		0.300 M		0 M		0 M
Δ []	- x		- x		+ x		+x
equil []	(0.496 - x) M		(0.300 - x) M		x M		x M

$$K_c = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]} = \frac{(x)(x)}{(0.496-x)(0.3-x)} = \frac{x^2}{0.149 - 0.796x + x^2}$$

Using the quadratic formula: $x = 0.847$ (extraneous) or 0.240

So at equilibrium: $[\text{SO}_2] = 0.496 - x = 0.256 \text{ M}$

$[\text{NO}_2] = 0.300 - x = 0.060 \text{ M}$

$[\text{SO}_3] = [\text{NO}] = x = 0.240 \text{ M}$

Example 3: The value of $K_c = 4.24$ at 800K for the reaction, $\text{CO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}_2(\text{g}) + \text{H}_2(\text{g})$ Calculate equilibrium concentrations of CO_2 , H_2 , CO and H_2O at 800 K, if only CO and H_2O are present initially at concentrations of 0.10M each.

Solution:

	$\text{CO}(\text{g})$	+	$\text{H}_2\text{O}(\text{g})$	\rightleftharpoons	$\text{CO}_2(\text{g})$	+	$\text{H}_2(\text{g})$
initial []	0.1M		0.1M		0		0
Δ []	- x		- x		+ x		+x
equil []	(0.1-x) M		(0.1-x) M		x M		x M

Hence, equilibrium constant can be written as, $K_c = x^2 / (0.1-x)^2 = 4.24$

$$3.24x^2 - 0.848x + 0.0424 = 0$$

Using the quadratic formula; $x = 0.067$ or 0.194. The value 0.194 should be neglected because it will give concentration of the reactant which is more than initial concentration. Hence the equilibrium concentrations are, $[\text{CO}_2] = [\text{H}_2] = x = 0.067 \text{ M}$

$[\text{CO}] = [\text{H}_2\text{O}] = 0.1 - 0.067 = 0.033 \text{ M}$

Example4: 13.8g of N_2O_4 was placed in a 1L reaction vessel at 400K and allowed to attain equilibrium. $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$. The total pressure at equilibrium was found to be 9.15 bar. Calculate K_c , K_p and partial pressure at equilibrium.

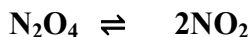
Solution:

Grade-11 Chemistry Short Note With Practice Questions

we know $pV = nRT$, total volume (V) = 1 L, molecular mass of $N_2O_4 = 92$ g, number of moles = $13.8g/92 g = 0.15$ of the gas (n), Gas constant (R) = $0.083 \text{ bar L mol}^{-1}\text{K}^{-1}$, Temperature (T) = 400 K.

$$pV = nRT = p \times 1L = 0.15 \text{ mol} \times 0.083 \text{ bar L mol}^{-1}\text{K}^{-1} \times 400 \text{ K}$$

$$p = 4.98 \text{ bar}$$



Initial pressure: 4.98 bar 0

At equilibrium: $(4.98 - x)$ bar $2x$ bar

Hence, P_{total} at equilibrium = $P_{N_2O_4} + P_{NO_2}$

$$9.15 = (4.98 - x) + 2x$$

$$9.15 = 4.98 + x$$

$$x = 4.17 \text{ bar}$$

Partial pressures at equilibrium are;

$$P_{N_2O_4} = 4.98 - 4.17 = 0.81 \text{ bar}, P_{NO_2} = 2x = 2 \times 4.17 = 8.34 \text{ bar}$$

$$K_p = \frac{[P_{NO_2}]^2}{[P_{N_2O_4}]} = (8.34)^2 / 0.81 = 85.87$$

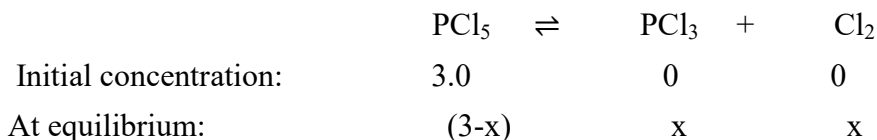
$$K_p = K_c(RT)^{\Delta n}$$

$$85.87 = K_c(0.083 \times 400)^1$$

$$K_c = 2.586 = 2.6$$

Example 5: 3.00 mol of PCl_5 kept in 1L closed reaction vessel was allowed to attain equilibrium at 380K. Calculate composition of the mixture at equilibrium. $K_c = 1.80$

Solution:



$$K_c = [PCl_3][Cl_2]/[PCl_5]$$

$$1.8 = x^2 / (3 - x)$$

$$x^2 + 1.8x - 5.4 = 0$$

Using quadratic formula, $x = 1.59$

$$[PCl_5] = 3.0 - x = 3 - 1.59 = 1.41 \text{ M}, [PCl_3] = [Cl_2] = x = 1.59 \text{ M}$$

Changing equilibrium conditions-Le Chatelier's principle

Factors affecting equilibria

Equilibrium constant, K_c is independent of initial concentrations. But if a system at equilibrium is subjected to a change in the concentration of one or more of the reacting substances, then the system is no longer at equilibrium; and net reaction takes place in some

Grade-11 Chemistry Short Note With Practice Questions

direction until the system returns to equilibrium once again. Similarly, a change in temperature or pressure of the system may also alter the equilibrium. In order to decide what course the reaction adopts and make a qualitative prediction about the effect of a change in conditions on equilibrium we use:

Le Chatelier's principle: States, when a stress is applied to a system at equilibrium, the equilibrium position will shift in the direction that relieves the applied stress.

Stresses include:

- Changes in concentration or pressure of reactants or products by the addition or removal
- Changes in pressure as a result of changes in volume of container
- Changes in temperature (value of K will change)
- Addition of a catalyst

A. Effect of Changes in Concentration on Equilibrium Position

In general, when equilibrium is disturbed by the addition/removal of any reactant/ products, Le Chatelier's principle predicts that:

- ❖ The concentration stress of an added reactant/product is relieved by net reaction in the direction that consumes the added substance
- ❖ The concentration stress of a removed reactant/product is relieved by net reaction in the direction that replenishes the removed substance.

Example: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$

0.683 M 8.80 M 1.05 M

Increase the concentration of NH_3 to 3.65M, the position of equilibrium shifts to the left

$$Q_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(3.65)^2}{(0.683)(8.80)^3} = 0.0286$$

$$\text{But } K_c = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3} = \frac{(1.05)^2}{(0.683)(8.80)^3} = 0.02$$

$Q_c > K_c$, a net reverse reaction (equilibrium shifted to the left)

B. Effect of Changing Pressure on Equilibrium Position

- ✚ A pressure change obtained by changing the volume can affect the yield of products in case of a gaseous reaction where the total number of moles of gaseous reactants and total number of moles of gaseous products are different.
- ✚ In applying Le Chatelier's principle to a heterogeneous equilibrium the effect of pressure changes on solids and liquids can be ignored because the volume (and concentration) of a solution/liquid is nearly independent of pressure.

Example 1: How does the position of equilibrium change as the pressure is increased?



Grade-11 Chemistry Short Note With Practice Questions

three moles of gaseous reactant two moles of gaseous product

An increase in pressure causes an increase in products.

Example 2: Consider the reaction, $\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$

Here, 4 mol of gaseous reactants ($\text{CO} + 3\text{H}_2$) become 2 mol of gaseous products ($\text{CH}_4 + \text{H}_2\text{O}$).

Suppose equilibrium mixture (for above reaction) kept in a cylinder fitted with a piston at constant temperature is compressed to one half of its original volume. Then, total pressure will be doubled (according to $PV = \text{constant}$).

The partial pressure and therefore, concentration of reactants and products have changed and the mixture is no longer at equilibrium.

The direction in which the reaction goes to re-establish equilibrium can be predicted by applying the Le Chatelier's principle.

Since pressure has doubled, the equilibrium now shifts in the forward direction, a direction in which the number of moles of the gas or pressure decreases (we know pressure is proportional to moles of the gas).

This can also be understood by using reaction quotient, Q_c .

Let $[\text{CO}]$, $[\text{H}_2]$, $[\text{CH}_4]$ and $[\text{H}_2\text{O}]$ be the molar concentrations at equilibrium for methanation reaction.

When volume of the reaction mixture is halved, the partial pressure and the concentration are doubled.

We obtain the reaction quotient by replacing each equilibrium concentration by double its value. As $Q_c < K_c$, the reaction proceeds in the forward direction

Example 3: In reaction $\text{C(s)} + \text{CO}_2\text{(g)} \rightleftharpoons 2\text{CO(g)}$, when pressure is increased, the reaction goes in the reverse direction because the number of moles of gas increases in the forward direction.

C. Effect of Inert Gas Addition

If the volume is kept constant and an inert gas such as argon is added which does not take part in the reaction, the equilibrium remains undisturbed. It is because the addition of an inert gas at constant volume does not change the partial pressures or the molar concentrations of the substance involved in the reaction. The reaction quotient changes only if the added gas is a reactant or product involved in the reaction.

D. Effect of Temperature Change

Whenever an equilibrium is disturbed by a change in the concentration, pressure or volume, the composition of the equilibrium mixture changes because the reaction quotient, Q_c no longer equals the equilibrium constant, K_c . However, when a change in temperature occurs, the value of equilibrium constant, K_c is changed.

In general, the temperature dependence of the equilibrium constant depends on the sign of ΔH for the reaction.

Grade-11 Chemistry Short Note With Practice Questions

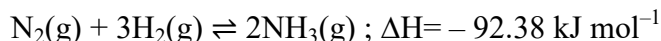
- ✓ The equilibrium constant for an exothermic reaction (negative ΔH) decreases as the temperature increases
- ✓ The equilibrium constant for an endothermic reaction (positive ΔH) increases as the temperature increases.

Note: temperature changes affect the equilibrium constant and rates of reactions.

Example: $\text{N}_2\text{O}_4(\text{g}) + 58 \text{ KJ} \rightleftharpoons 2\text{NO}_2(\text{g})$. Endothermic reaction

Increase temperature: Favours forward reaction

Example: Production of ammonia according to the reaction:



According to Le Chatelier's principle, raising the temperature shifts the equilibrium to left and decreases the equilibrium concentration of ammonia. In other words, low temperature is favourable for high yield of ammonia, but practically very low temperatures slow down the reaction and thus a catalyst is used

E. Effect of a Catalyst

- ✚ A catalyst enhances the rate of a reaction by lowering the reaction's activation energy.
- ✚ A catalyst lowers the activation energy of the forward reaction and the reverse reaction to the same extent. We can therefore conclude that:
- ✚ The presence of a catalyst does not alter the equilibrium constant, nor does it shift the position of an equilibrium system.
- ✚ Adding a catalyst to a reaction mixture that is not at equilibrium will simply cause the mixture to reach equilibrium sooner.
- ✚ The same equilibrium mixture could be obtained without the catalyst, but we might have to wait much longer for it to happen.

Note: catalyst lower the activation energy of reactions and make the reaction faster but does not change the equilibrium constant (K_{eq}).

Example: Let us consider the formation of NH_3 from dinitrogen and dihydrogen which is highly exothermic reaction and proceeds with decrease in total number of moles formed as compared to the reactants. Equilibrium constant decreases with increase in temperature.

Chemical equilibrium and industry

A. Haber process for the manufacture of ammonia



For reaction above equilibrium constants are shown below

T	K
398K	57
298K	812
198K	1.6×10^5

Grade-11 Chemistry Short Note With Practice Questions

Equilibrium favors products at high pressure and low temperature. However, speed of reaction favored at high temperature. At low temperature rate decreases and it takes long time to reach at equilibrium, whereas high temperatures give satisfactory rates but poor yields.

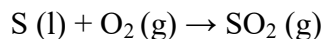
German chemist, Fritz Haber discovered that a catalyst consisting of iron catalyse the reaction to occur at a satisfactory rate at temperatures, where the equilibrium concentration of NH_3 is reasonably favourable. Since the number of moles formed in the reaction is less than those of reactants, the yield of NH_3 can be improved by increasing the pressure. Optimum conditions of temperature and pressure for the synthesis of NH_3 using catalyst are around $380 - 450^\circ\text{C}$ and 200 atm.

$\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g}) ; \Delta H = -92 \text{ KJ/mol}$		
Typical conditions	Pressure	20000 Kpa (200 atm)
	Temperature	$380 - 450^\circ\text{C}$
	Catalyst	Iron
Equilibrium theory Favors	Low temperature	Exothermic reaction-higher yield at low T
	High pressure	Decrease in number of gaseous molecules
Kinetic theory favors	High temperature	Greater average energy + more frequent collision
	High pressure	More frequent collisions for gaseous molecules
	Catalyst	Lower activation energy
Compromise conditions	Which is better?	A low yield in a shorten time or A high yield over a longer period

B. Contact process for the manufacture of sulphuric acid

Most of the sulfuric acid manufactured is produced using the Contact Process, a process involving the catalytic oxidation of sulphur dioxide, SO_2 , to sulphur trioxide, SO_3 . This process is called the Contact Process (because the SO_2 and O_2 gases must come into contact with a catalyst).

1. Liquid sulphur is sprayed with excess dry air at atmospheric pressure.



2. The oxidation of SO_2 to SO_3 is the most difficult part of making sulfuric acid. The reaction is as follows:



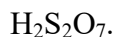
This reaction is an exothermic reaction (energy is released), so by Le Chatelier's Principle, higher temperatures will force the equilibrium position to shift to the left of the equation favouring the production of sulphur dioxide. Lower temperatures would favour the production of the product sulphur trioxide and result in a higher yield. However, the rate of reaching equilibrium at the lower temperatures is extremely low. Higher temperature means equilibrium is established more rapidly but the yield of sulphur trioxide is lower. A temperature of 450°C is a compromise whereby a faster reaction rate results in a slightly lower yield.

Grade-11 Chemistry Short Note With Practice Questions

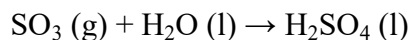
Similarly, at higher pressures, the equilibrium position shifts to the side of the equation in which there are the least numbers of gaseous molecules. On the left hand side of the reaction there are 3 moles of gaseous reactants, and the right hand side there are 2 moles of gaseous products, so higher pressure favours the right hand side, by Le Chatelier's Principle. Higher-pressure results in a higher yield of sulphur trioxide. A vanadium catalyst is also used in this reaction in order to speed up the rate of the reaction.

In summary, the conditions used to convert SO_2 to SO_3 is pressure a little above atmospheric, small excess of oxygen, a catalyst of vanadium oxide supported on silica and temperatures of catalyst beds of 550°C (for high rate) and 400°C (for high conversion).

- Any unreacted gases from the above reaction are recycled back into the above reaction.
- Sulphur trioxide, SO_3 (g) is dissolved in 98% (18M) sulfuric acid, H_2SO_4 , to produce disulfuric acid or persulfuric acid, also known as fuming sulfuric acid or oleum,



This is done because when water is added directly to sulphur trioxide to produce sulfuric acid the reaction is slow and tends to form a mist of sulfuric acid droplets in residual gasses (N_2 with some O_2) in which the particles refuse to coalesce.



- Water is added to the disulfuric acid, $\text{H}_2\text{S}_2\text{O}_7$, to produce sulfuric acid, H_2SO_4
$$\text{H}_2\text{S}_2\text{O}_7 (\text{l}) + \text{H}_2\text{O} (\text{l}) \rightarrow 2\text{H}_2\text{SO}_4 (\text{l})$$

5.2 Phase equilibrium

Common terms: phase, components and degree of freedom

Phase: “The physically distinct, homogenous and mechanically separable part of a system are called phases”.

Examples

- A gaseous mixture constitutes a single phase since gases are completely miscible. Air is a mixture of N_2 , O_2 , CO_2 , water vapour etc. Which constitute a single phase.
- Two or more liquids which are miscible with one another constitute a single phase as there is no bounding surfaces separating the different liquids. e.g., water and alcohol, benzene & chloroform constitute one phase system.
- A system consisting of a liquid in equilibrium with its vapour constitute a two phase system

Component : The number of component of a system at equilibrium is defined as the minimum number of independently variable constituents which are required to express the composition of each phase in the system. In a chemically reactive system, the number of components is given by

$$C = N - E$$

Grade-11 Chemistry Short Note With Practice Questions

where C = components, N = Number of chemical species, E = Number of independent equations relating the concentrations of the N species. Each independent chemical equilibrium involving the constituents counts as one equation. The condition that a solution be electrically neutral also counts as one equation if ions are considered as constituents.

Examples

- i. **Sulphur system:** Consists of four phases namely monoclinic sulphur, rhombic sulphur, liquid sulphur and sulphur vapour. The composition of each phase of the system can be expressed in terms of sulphur only, so, it is a one component system.
- ii. **Water system:** It is a one component system because the composition of each of the three phases present can be expressed as H_2O .
- iii. **Na_2SO_4 + water system:** Certain salts are capable of existing as hydrates with different number of water molecules of crystallization. These hydrates correspond to different solids and hence two different phases. The system is a two component, because the composition of each phase of the hydrates is completely described in terms of the anhydrous salt and water alone. e.g., Na_2SO_4 + water

Degrees of Freedom : The degree of freedom or variance of a system is defined as the minimum number of variable factors such as temperature, pressure and concentration which should be arbitrarily fixed in order to define the system completely.

Examples

- i. For a given sample of any gas $PV = nRT$. Any two of the three variables P , V , T define the system completely. Hence the system is bivariant or it has two degrees of freedom.
- ii. A gaseous mixture say N_2 and O_2 gases (mixed 50% each), is completely defined when three variables temperature, pressure and concentration are specified. Thus, the degrees of freedom is three or the system is trivariant.

Note:

- ❖ The greater the number of components in a system, the greater is the degree of freedom for a given number of phases.
- ❖ The greater the number of phases, the smaller is the number of degrees of freedom.
- ❖ The number of phases is maximum when the number of degrees of freedom = Zero, for a given number of components.

Phase Rule

The phase rule, also known as the Gibbs phase rule, relates the number of components and the number of degrees of freedom in a system at equilibrium by the formula:

$$F = C - P + 2$$

Where: F equals the number of degrees of freedom or the number of independent variables,

Grade-11 Chemistry Short Note With Practice Questions

C equals the number of components in a system in equilibrium and

P equals the number of phases

The digit 2 stands for the two variables, temperature and pressure.

Example: In the reaction involving the decomposition of calcium carbonate on heating, there are three phases – two solid phases and one gaseous phase.



There are also three different chemical constituents, but the number of components is only two because **any two constituents completely define the system in equilibrium**. Any third constituent may be determined if the concentration of the other two is known.

Substituting into the phase rule we can see that the system is univariant, since:

$$F = C - P + 2 = 2 - 3 + 2 = 1.$$

Therefore only one variable, either temperature or pressure, can be changed independently.

(The number of components is not always easy to determine at first glance, and it may require careful examination of the physical conditions of the system at equilibrium.)

Application of Gibbs Phase Rule to One Component System:

From the mathematical expression, $F = C - P + 2$

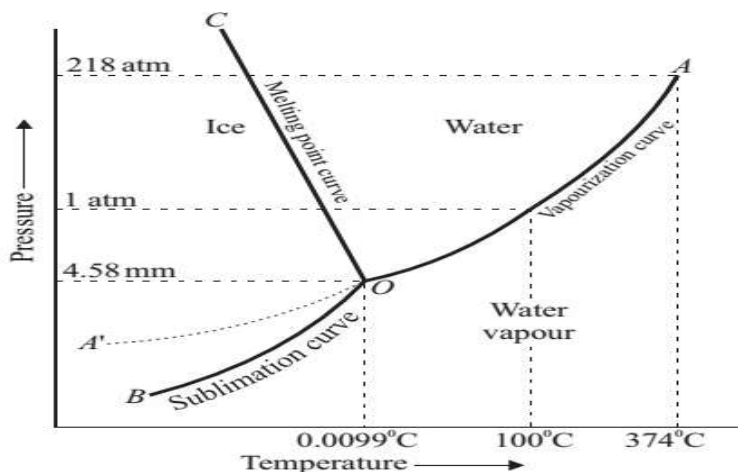
When $C = 1$, $P = 1$, $F = 1 - 1 + 2 = 2$

Hence, all one component systems can be completely described graphically by stating only two variables, pressure and temperature on appropriate axis.

The Water System : It is a one component system. Water exists in three possible phases viz. ice (solid), water (liquid), and vapour (gas). These three single phases may form four possible equilibria.

- Solid \rightleftharpoons Liquid
- Liquid \rightleftharpoons Vapour
- Solid \rightleftharpoons Vapour
- Solid \rightleftharpoons Liquid \rightleftharpoons Vapour

Phase Diagram of water system



Grade-11 Chemistry Short Note With Practice Questions

The phase diagram consists of :

- i. Stable curves: three OB, OA and OC
- ii. Metastable curve: one OA'
- iii. Areas: three AOB, COB and AOC
- iv. Triple point: One O

Stable curves

OA : It is known as vapour pressure curve of water.

- ❖ The curve OA starts from point O i.e., freezing point of water, 0.0098°C under 4.579 mm of Hg pressure and ends at A, the critical temperature (374°C at 218 atm.).
- ❖ Above critical temp. on the vapour phase exists whatever may be the value of pressure
- ❖ The vapour pressure of water increases with increase in temperature.
- ❖ The rate of increase of its vapour pressure with temperature is relatively higher at higher temperatures and therefore the curve OA slants upwards and slopes away from the temperature axis.
- ❖ From phase rule, $F = C - P + 2 = 1 - 2 + 2 = 1$
- ❖ The water vapour system is univariant

OB : It is the sublimation curve of ice.

- ✓ Along this curve, solid ice is in equilibrium with its vapour.
- ✓ This curve is not the prolongation of curve A but falls off more steeply.
- ✓ Curve OB starts from the temperature 0.0098°C above which solid water i.e., ice cannot exist.
- ✓ The curve terminates at B i.e., absolute zero (-273°C). At this temperature, no vapour can exist and, therefore, only ice is left. But on other points of the curve OB, ice is in equilibrium with vapour.

Hence, there are two phases.

- ✓ According to phase rule, $F = C - P + 2 = 1 - 2 + 2 = 1$.
- ✓ Thus, the system is univariant. This means that for each temperature; there may be one pressure and for each pressure there may be one temperature.

OC : This curve is the melting point curve or fusion curve of ice.

- Along this curve two phases, ice and water are in equilibrium.
- The inclination of OC line towards the pressure axis indicates that the melting point of ice is slightly lowered by increase of pressure.
- (According to Le Chatelier's principle the increase in pressure causes the water - ice equilibrium to shift in such a direction that there is a decrease in volume.).
- As the melting point of ice is accompanied by decrease in volume, it should be lowered by the increase of pressure.
- The curve OC starts from point O but there is no limit for this curve. It goes upto a point corresponding to 2000 atm. and -20°C .
- According to phase rule, $F = C - P + 2 = 1 - 2 + 2 = 1$.

Grade-11 Chemistry Short Note With Practice Questions

- Thus, the system is univariant. This means that for any given pressure, melting point must have one fixed value.

Metastable Curve

OA' : It is a metastable curve shown in continuation of AO.

- ✚ When water is cooled below its freezing point (when it is not vigorously stirred) without separation of ice, the water is said to be super cooled.
- ✚ The vapour pressure curve of liquid water AO extends below O as shown by the dotted curve OA'.
- ✚ Along curve OA' liquid water coexists with vapour and vapour pressures are different than over the solid.
- ✚ This equilibrium is called metastable equilibrium as slight disturbance brings it to the stable region OB of the phase diagram.

Areas

- The areas give the conditions of temperature and pressure under which single phase - ice (solid), water (liquid) and vapour (gas) can exist.
- It is necessary to specify both temperature and pressure to define a system within this area.
- In the area BOC, AOC and AOB exists, ice (solid), water (liquid) and vapour (gas) respectively.
- In these areas, the degrees of freedom for the system is **two** or they are **bivariant**.

The Triple point O

- The point O at which the curves AO, BO and CO meet is called the triple point.
- At this point all the three phases viz , ice, water and vapour co-exist.
- Thus, $P = 3$.
- According to phase rule, at triple point O, $F = C - P + 2 = 1 - 3 + 2 = 0$
- Thus, the degree of freedom at triple point is zero, which indicates that there is only one set of variables P, T at which all the three phases coexist.
- If any of the variables is changed, then the number of phases decreases.

For example, if the temperature is raised, heat causes more and more of the solid (ice) to melt but no change in temperature or pressure of the system occurs till the whole of the solid has completely changed into liquid (water) and the system becomes a two phase system.

On applying the pressure to the system, vapours start condensing to liquid or solid phase. As long as there are three phases, temperature and pressure remains same.

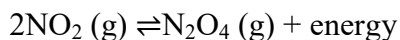
The triple point O is a self defined point corresponding to 0.0075°C temperature and 4.579 mm of Hg pressure (difference from the ordinary freezing point, 0.000°C at 760mm pressure) because freezing temperature is lower than triple point temperature due to effect of high pressure and dissolved air).

Grade-11 Chemistry Short Note With Practice Questions

Review Exercise

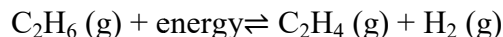
Choose the Correct answer for the following questions.

1. Consider the following equilibrium:



The equilibrium will shift to the left as a result of

- A. Adding a catalyst. C. Increasing the volume.
B. Removing some N_2O_4 . D. Decreasing the temperature.
2. Ethene, C_2H_4 , can be produced in the following industrial system:



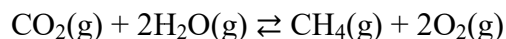
The conditions that are necessary to maximize the equilibrium yield of C_2H_4 are:

- A. Low temperature and low pressure. C. Low temperature and high pressure
B. High temperature and low pressure. D. High temperature and high pressure.
3. Consider the following equilibrium: $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$.

The volume of the equilibrium system is increased and a new equilibrium is established. Compared to the rates in the original equilibrium, which of the following describes the rates of the forward and reverse reactions as a new equilibrium is being established?

	FORWARD RATE	REVERSE RATE
A.	decreased	decreased
B.	increased	increased
C.	decreased	increased
D.	remained constant	remained constant

4. Consider the following equilibrium:



Which of the options below indicates that the reactants are favored?

- A. K_{eq} is zero. C. K_{eq} is very small
B. K_{eq} is slightly less than 1 D. K_{eq} is slightly greater than 1.
5. The value of K_{eq} changes when:
- A. A catalyst is added C. The temperature changes
B. The surface area changes. D. The concentration of reactants changes
6. Which of the following is true for a chemical reaction at equilibrium?
- A. Only the forward reaction stops
B. Only the reverse reaction stops
C. Both the forward and reverse reactions stop
D. The rate constants for the forward and reverse reactions are equal
E. The rates of the forward and reverse reactions are equal
7. Which of the following is true regarding the concentration of products, for a chemical reaction that is already at equilibrium, assuming no disruptions to the equilibrium?
- A. The concentrations of products will not change because there are no more reactants.

Grade-11 Chemistry Short Note With Practice Questions

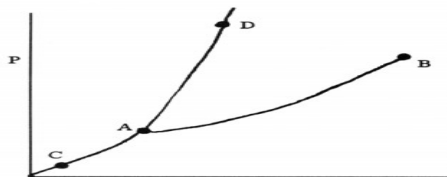
- B. The concentrations of products will not change because the limiting reagent is gone.
C. The concentrations of products will not change because the forward and reverse rates are equal.
D. The concentrations of products will change continually because of reversibility.
8. Which of the following are equal for a chemical system at equilibrium?
A. The concentrations of reactant and products are equal
B. The rate constants for the forward and reverse reactions are equal
C. The time that a particular atom or molecule spends as a reactant and product are equal
D. The rate of the forward and reverse reaction
9. An equilibrium that strongly favors products has:
A. A value of $K \ll 1$.
B. A value of $K \gg 1$
C. A value of $Q \ll 1$
D. A value of $Q \gg 1$
10. If the reaction quotient Q has a smaller value than the related equilibrium constant, K :
A. The reaction is at equilibrium.
B. The reaction is not at equilibrium, and will make more products at the expense of reactants.
C. The reaction is not at equilibrium, and will make more reactants at the expense of products.
D. The value of K will decrease until it is equal to Q
11. If the equilibrium is established by initially adding 0.10 mol each of A and B to a 1L container, then which of the following must be true once the mixture achieves equilibrium?
- $A + 2B \rightleftharpoons 2C, K = 320$
- A. $[A] = [B]$
B. $[B] = 2[C]$
C. $[A] = [B] = [C]$
D. $[A] > [B]$
12. For the chemical equilibrium: $A + 2B \rightleftharpoons 2C$, the value of the equilibrium constant, K , is 10. What is the value of the equilibrium constant for the reaction written in reverse?
 $2C \rightleftharpoons A + 2B, K = ???$ Given that: $A + 2B \rightleftharpoons 2C, K = 10$
A. 0.10 B. 100 C. 10 D. -10 E. 1
13. In which of the following does the reaction go farthest to completion?
A. $K = 10^5$ B. $K = 10^{-5}$ C. $K = 1000$ D. $K = 100$ E. $K = 1$
14. Consider the equilibrium $N_2(g) + 3 H_2(g) \rightleftharpoons 2 NH_3(g)$ at a certain temperature. An equilibrium mixture in a 4.00-L vessel contains 1.60 mol NH_3 , 0.800 mol N_2 , and 1.20 mol H_2 . What is the value of K_c ?
A. 9.00 B. 29.6 C. 3.37 D. 17.1 E. 7.41
15. For the reaction system: $2HI(g) \rightleftharpoons H_2(g) + I_2(g)$, $K_c = 0.020$ at 720 K. If the initial concentrations of HI, H_2 , and I_2 are all $1.50 \times 10^{-3} M$ at 720 K, which one of the following statements is correct?
A. The system is at equilibrium.
B. The concentrations of HI and I_2 will increase as the system is approaching equilibrium.
C. The concentrations of H_2 and HI will decrease as the system is approaching equilibrium.
D. The concentration HI will increase as the system is approaching equilibrium.
E. The concentrations of H_2 and I_2 will increase as the system is approaching equilibrium.

Grade-11 Chemistry Short Note With Practice Questions

16. Exactly 1.0 mol N_2O_4 is placed in an empty 1.0-L container and is allowed to reach equilibrium described by the equation: $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2 \text{NO}_2(\text{g})$. If at equilibrium the N_2O_4 is 20% dissociated, what is the value of the equilibrium constant (in units of mol/L) for the reaction under these conditions?
- A. 0.05 B. 0.2 C. 0.5 D. 20 E. 400
17. For the equilibrium: $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$; $K_c = 4.0$ at 228°C . Pure PCl_5 is added to the reaction system. At equilibrium there is 0.040 M PCl_5 . What is the concentration of PCl_3 in this system?
- A. 0.16 M B. 0.20 M C. 0.40 M D. 0.80 M E. 1.6 M
18. For which of the following reactions are the numerical values of K_p and K_c the same?
- A. $2 \text{CO}(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2 \text{CO}_2(\text{g})$ C. $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
B. $\text{N}_2(\text{g}) + 3 \text{H}_2(\text{g}) \rightleftharpoons 2 \text{NH}_3(\text{g})$ D. $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2 \text{HI}(\text{g})$
19. Which of the following equilibria would not be affected by volume changes at constant temperature?
- A. $2 \text{NO}_2(\text{g}) \rightleftharpoons \text{N}_2\text{O}_4(\text{g})$ C. $2 \text{NO}(\text{g}) + 3 \text{F}_2(\text{g}) \rightleftharpoons 2 \text{F}_3\text{NO}(\text{g})$
B. $\text{F}_3(\text{g}) + \text{NH}_3(\text{g}) \rightleftharpoons \text{H}_3\text{NBF}_3(\text{s})$ D. $\text{O}_3(\text{g}) + \text{NO}(\text{g}) \rightleftharpoons \text{NO}_2(\text{g}) + \text{O}_2(\text{g})$
20. Gibbs phase rule for general system:
- A. $P+F=C-1$ B. $P+F=C+1$ C. $P+F=C-2$ D. $P+F=C+2$
21. In a single-component condensed system, if degree of freedom is zero, maximum number of phases that can co-exist:
- A. 0 B. 1 C. 2 D. 3
22. Based on the equilibrium constants given, in which of these reactions are the products most favoured over the reactants?
- A. $K_{eq} = 0.002$ B. $K_{eq} = 0.0$ C. $K_{eq} = 3.5$ D. $K_{eq} = 6.0 \times 10^{-4}$
23. Which of the statements regarding the effect of reactant or product concentration changes on equilibrium is false?
- A. Adding a reactant always shifts a reversible reaction in the direction of reactants.
B. Removing a reactant always shifts a reversible reaction in the direction of reactants.
C. Adding a product causes a reversible reaction to shift in the direction of the formation of the reactants.
D. Removing a product shifts a reversible reaction in the direction of formation of products.
24. Which of the following statements correctly describes the effect of an increase in temperature on the position of equilibrium for an exothermic reaction?
- A. The position of equilibrium does not change if you change the temperature.
B. According to Le Châtelier's Principle, the position of equilibrium will move to the left if the temperature is increased.
C. If you increase the temperature, the position of equilibrium will move in such a way as to increase the temperature again.
D. Equilibrium will shift to favour the reaction which releases heat in this context.

Grade-11 Chemistry Short Note With Practice Questions

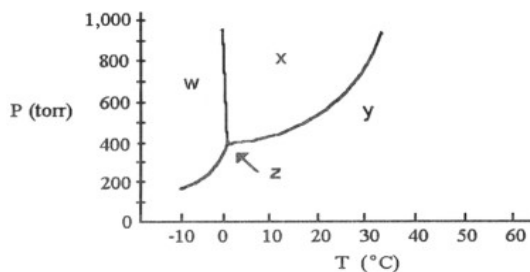
25. Which explanation best describes why a catalyst does not affect the position of equilibrium?
- A catalyst will increase the activation energy in both directions; therefore, no shift in the position of equilibrium will result.
 - The position of equilibrium is not affected by adding a catalyst since a catalyst speeds up both the forward and back reactions by exactly the same amount.
 - Equilibrium constants are not affected by adding a catalyst since no shift in the position of equilibrium occurs.
 - Not all chemical reactions respond to catalysts.
26. Which one of the following statements regarding a dynamic equilibrium is false?
- At equilibrium, there is no net change in the system
 - At equilibrium, the concentration of reactants and products stays the same
 - At equilibrium, the forward and back reactions cease to occur
 - At equilibrium, the rates of the forward and back reactions are identical
27. K_1 and K_2 are the velocity constants of the forward and back ward reactions. The equilibrium constant K of the reaction is:
- $K_1 \times K_2$
 - $K_1 - K_2$
 - K_1 / K_2
 - $\frac{K_1 + K_2}{K_1 - K_2}$
28. The rate of Kp/Kc for the reaction: $\text{CO (g)} + \frac{1}{2} \text{O}_2 \rightleftharpoons \text{CO}_2 \text{ (g)}$ is:
- 1
 - $(RT)^{1/2}$
 - RT
 - $(RT)^{-1/2}$
29. Equilibrium constant greater than 1 means:
- Equilibrium constant lies to the right , products predominate.
 - Equilibrium constant lies to the left, reactants predominate.
 - Equilibrium constant lies to the right, reactants predominate.
 - Equilibrium constant lies to the left, products predominate
30. Equilibria in which all substances are in the same phases are called
- Homogeneous equilibria.
 - Equilibrium constant
 - Heterogeneous equilibria.
 - Equilibrium constant expression
31. Which of the following statements is NOT true in relation to the triple point on a single component phase diagram?
- The point at which the solid, liquid and gaseous phases for a substance co-exist
 - The triple point exists for a substance occurs at a specific temperature and pressure
 - The triple point exists at a single temperature and is independent of pressure
 - The system must be enclosed so that no vapour can escape
32. On the phase diagram shown below, segment _____ corresponds to the conditions of temperature and pressure under which the solid and the gas of the substance are in equilibrium.



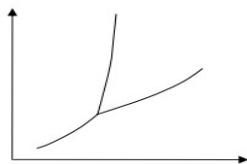
- AD
 - BC
 - AC
 - AB
33. On the phase diagram shown above (Question number 32), the coordinates of point correspond to the critical temperature and pressure.
- A
 - B
 - C
 - D

Grade-11 Chemistry Short Note With Practice Questions

34. According to the phase diagram shown below, the normal boiling point of this substance is:

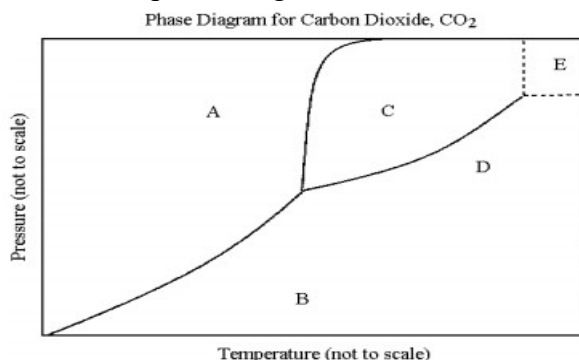


- A. 10 B. -3 C. 29 D. 38 E. 0
35. The phase diagram of oxygen (O_2) includes the following points:
- Critical point: $T_c = 154.5 \text{ K}$, $P_c = 49.3 \text{ atm}$
 - Triple point: $T_{t.p.} = 54.33 \text{ K}$, $P_{t.p.} = 0.00150 \text{ atm}$
 - Normal melting point: $T_{fus} = 54.8 \text{ K}$
 - Normal boiling point: $T_{vap} = 90.2 \text{ K}$
 - You may wish to use the sketch below to help you answer the question.



Under which one of the following conditions is oxygen a liquid?

- A. $T = 50. \text{ K}$, $P = 0.80 \text{ atm}$ D. $T = 60. \text{ K}$, $P = 0.50 \text{ atm}$
 B. $T = 90. \text{ K}$, $P = 0.20 \text{ atm}$ E. $T = 100. \text{ K}$, $P = 0.050 \text{ atm}$
 C. $T = 150. \text{ K}$, $P = 1.0 \text{ atm}$
36. Consider the phase diagram for CO_2 .



In going from phase A to phase B,

- A. Gaseous CO_2 becomes solid. D. Solid CO_2 becomes gaseous.
 B. Solid CO_2 becomes liquid. E. Liquid CO_2 becomes solid.
 C. Liquid CO_2 becomes gaseous.
37. Consider the phase diagram for CO_2 above (question number 36).
 The point on the graph labelled E represents CO_2 as:
- A. Gas. C. Supercritical fluid.
 B. Solid. D. CO_2 cannot exist at that temperature and pressure.

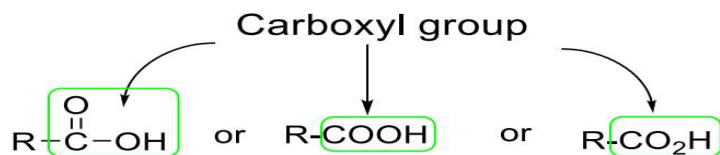
UNIT 6

Carboxylic Acids, Esters, Fats and Oils

Main topics

- Carboxylic Acids
- Esters
- Fats and Oils

Carboxylic acids are organic compounds containing the carboxyl group (-COOH), wherein the hydroxyl group (-OH) is directly attached to the carbonyl (C=O) group.



Three representations of a carboxylic acid

Nomenclature of Carboxylic Acids

The common names of some basic carboxylic acids are derived from Latin names that indicate the first original natural source of the carboxylic acid.

Structure of Acid	Natural Source	Common Name
$\text{H}-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Ants (Formica)	Formic acid
$\text{CH}_3-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Vinegar (Acetum)	Acetic acid
$\text{CH}_3\text{CH}_2-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Basic Fat (Propio)	Propionic acid
$\text{CH}_3\text{CH}_2\text{CH}_2-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Rancid butter (Butyrum)	Butyric acid
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Present in a Valerian herb	Valeric acid
$\text{CH}_3\text{CH}_2\text{CH}_2\text{CH}_2\text{CH}_2-\overset{\text{O}}{\parallel}{\text{C}}-\text{OH}$	Goat (Caper)	Caproic acid

Common Names of Carboxylic Acids

The common name of a carboxylic acid (R-COOH) is derived by adding the suffix -ic acid to a prefix representing the chain length of the carboxylic acid.

# of Carbons	Prefix	Common Name of Acid
1	Form-	Formic acid
2	Acet-	Acetic acid
3	Propion-	Propionic acid
4	Butyr-	Butyric acid
5	Valer-	Valeric acid
6	Capro-	Caproic acid
Aromatic acid	Benzo-	Benzoic acid

Grade-11 Chemistry Short Note With Practice Questions

IUPAC Nomenclature of Aliphatic Carboxylic Acids

IUPAC names of straight chain aliphatic carboxylic acids are derived by adding the suffix –oic acid to the systematic name of the parent hydrocarbon. They are named as alkanoic acids

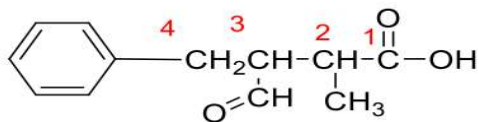
# of Carbons	Structure & IUPAC Name of Alkane	Structure & IUPAC Name of Acid
1	$\begin{array}{c} \text{H} \\ \\ \text{H}-\text{C}-\text{H} \\ \\ \text{H} \end{array}$ Methane	$\begin{array}{c} \text{O} \\ \\ \text{H}-\text{C}-\text{OH} \end{array}$ Methanoic acid
2	CH_3-CH_3 Ethane	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3-\text{C}-\text{OH} \end{array}$ Ethanoic acid
3	$\text{CH}_3\text{CH}_2-\text{CH}_3$ Propane	$\begin{array}{c} \text{O} \\ \\ \text{CH}_3\text{CH}_2-\text{C}-\text{OH} \end{array}$ Propanoic acid

Systematic Nomenclature of Substituted Carboxylic Acids

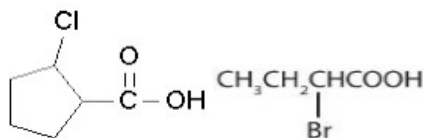
The systematic names of substituted aliphatic carboxylic acids are derived by:

- First identifying the parent chain that contains most, if not all, the carboxyl groups.
- Number the parent chain from the carbon of the carboxyl group i.e the carboxyl carbon is C-1.
- Identify the substituents and assign each substituent a locator/address number (2,3,4...etc.) consistent with the numbering in the parent chain.
- Arrange the names of the substituents in alphabetical order in the systematic name of the poly-substituted carboxylic acid.

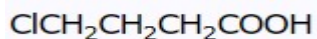
Examples:



3-Formyl-2-methyl-4-phenylbutanoic acid



2-chlorocyclopentanoic acid



4-chlorobutanoic acid

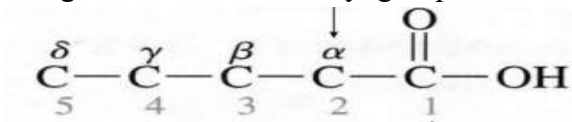
Grade-11 Chemistry Short Note With Practice Questions

Naming using Greek letters α β γ δ ϵ

There are two ways to identify substituent carbons in carboxylic acid:

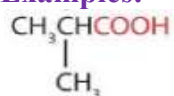
- ✓ numbers or
- ✓ Greek letters.

Using numbers, the carboxyl group carbon is given the number one.

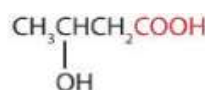


Greek letters are used to designate the position of substituent relative to the carbon of the carboxyl group. The carbon of the carboxyl group is NOT given a Greek letter.

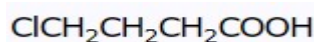
Examples:



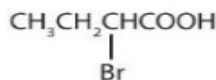
α -Methylpropionic acid



β -Hydroxybutyric acid



γ -chlorobutyric acid

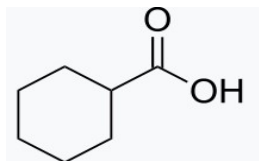


α -bromobutyric acid

Systematic Nomenclature of Cyclic Carboxylic Acids

The systematic name of a carboxylic acid in which the COOH group is attached directly to a ring is derived by adding a suffix – carboxylic acid to the name of the attached cycloalkane or cycloalkene or arene.

Example:

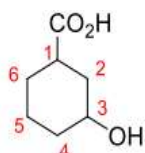


Cyclohexanecarboxylic acid

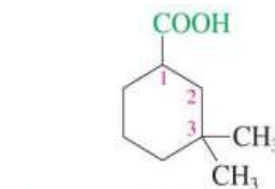
When such carboxylic acids are substituted, the carbon of the COOH group is itself not numbered, but it is, by convention, taken to be attached to C-1 of the ring.

Systematic Nomenclature of Substituted Cyclic Carboxylic Acids

For substituted carboxylic acids with the carboxyl group attached directly to a ring, the carbon of the COOH group is itself not numbered, but it is, by convention, taken to be attached to C-1 of the ring.



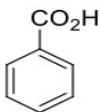
3-Hydroxycyclohexanecarboxylic acid



3,3-dimethylcyclohexanecarboxylic acid

Grade-11 Chemistry Short Note With Practice Questions

Under this system of nomenclature, benzoic acid would be named as benzenecarboxylic acid.

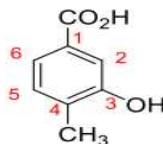


Benzenecarboxylic acid (benzoic acid)

IUPAC Nomenclature of Substituted Aromatic Carboxylic Acids

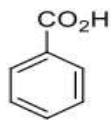
Substituted aromatic acids with one carboxyl group are named as derivatives of benzoic acid, with the position of substituents being cited using the locators (2,3etc) according to their position on the benzene ring relative to the carboxyl group. The carbon on which the carboxyl group is attached is by convention C-1

Example:

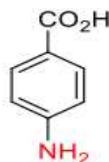


3-Hydroxy-4-methylbenzoic acid or 3-Hydroxy-4-methylbenzenecarboxylic acid

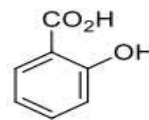
Typically aromatic acids of the form Ar-CO₂H are named as derivatives of benzoic acids, with ortho, meta and para indicating the location relative to the carboxyl group. (Recall that this is non-IUPAC).



benzoic acid



p-aminobenzoic acid



o-hydroxybenzoic acid

Systematic Nomenclature of Dicarboxylic Acids

The systematic name of an open chain aliphatic dicarboxylic acid is derived by adding a suffix -dioic acid to the name of the parent hydrocarbon i.e. alkanedioic acid.

Systematic Name	Common Name	Structure
Ethanedioic acid	Oxalic acid	HO ₂ C-CO ₂ H
Propanedioic acid	Malonic acid	HO ₂ CCH ₂ CO ₂ H
Butanedioic acid	Succinic acid	HO ₂ C(CH ₂) ₂ CO ₂ H
Pentanedioic acid	Glutaric acid	HO ₂ C(CH ₂) ₃ CO ₂ H
Hexanedioic acid	Adipic acid	HO ₂ C(CH ₂) ₄ CO ₂ H
Heptanedioic acid	Pimelic acid	HO ₂ C(CH ₂) ₅ CO ₂ H

Systematic Nomenclature of Substituted Dicarboxylic Acids

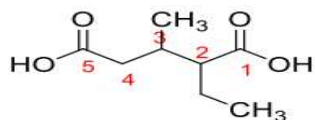
The systematic name of an aliphatic dicarboxylic acid is derived by:

- First identifying the parent chain that contains the two carboxylic acid groups and then adding a suffix -dioic acid to the name of the parent hydrocarbon.

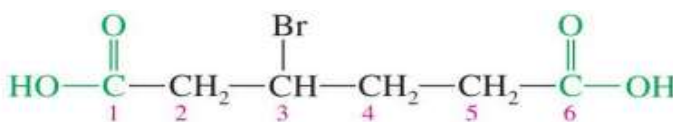
Grade-11 Chemistry Short Note With Practice Questions

- ii. The parent chain is numbered from the end that gives the substituents in the chain the lowest possible address number.
- iii. The substituents are arranged in alphabetical order in the full name of the dicarboxylic acid.

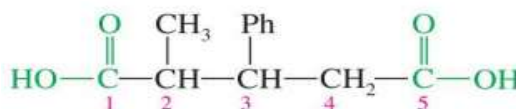
Examples:



2-Ethyl-3-Methylpentanedioic acid



3-bromohexanedioic acid



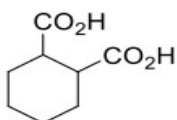
2-methyl-3-phenylpentanedioic acid

Systematic Nomenclature of Cyclic Dicarboxylic Acids

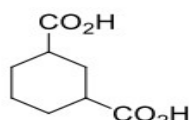
The systematic name of a cyclic aliphatic or aromatic dicarboxylic acid is derived by adding a suffix -dicarboxylic acid to the name of the parent cycloalkane or arene i.e.

cycloalkanedicarboxylic acid or arenedicarboxylic acid.

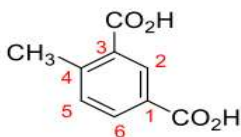
Examples



Cyclohexane-1,2-dicarboxylic acid



Cyclohexane-1,3-dicarboxylic acid



4-Methylbenzene-1,3-dicarboxylic acid



4-Methylbenzene-1,2-dicarboxylic acid

The positions of the two carboxyl groups are cited with the lowest possible address numbers to differentiate between isomers.

Exercise

1. Draw the structure for each compound.
 - a. heptanoic acid
 - b. 2,3-dibromobenzoic acid
 - c. 3-methylbutanoic acid
 - d. β -hydroxybutyric acid
2. Draw the structure for each compound.
 - a. *o*-nitrobenzoic acid
 - b. 3-chloropentanoic acid
 - c. *p*-chlorobenzoic acid
 - d. α -chloropropionic acid

- ## Properties of Carboxylic Acids
- ### Physical Properties of Carboxylic Acids

Physical Properties of Carboxylic Acids

- $$\begin{array}{c} \text{O} \cdots \text{OH} \\ \diagup \quad \diagdown \\ \text{R}-\text{C} \quad \text{C}-\text{R} \\ \diagdown \quad \diagup \\ \text{OH} \cdots \text{O} \end{array}$$

- ### Exercise: Predicting Physical Properties

a. 1-pentanol b. hexane c. butanoic acid d. pentanal

b. Which member of each of the following pairs of compounds would you expect to have a higher solubility in water?

2-butanone or propanoic acid hexanoic acid or ethanoic acid

Grade-11 Chemistry Short Note With Practice Questions

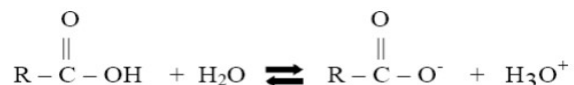
Reactions of Carboxylic Acids

The reactions of carboxylic acids can be directed to various sites on the carboxyl group

Reactions of carboxylic acids can be placed into four categories:

- Reactions at the acidic hydrogen on the carboxyl group.
- Reactions at the carbonyl group
- Reactions at the carboxylate oxygen
- Reactions that lead to loss of the carboxyl group as CO_2

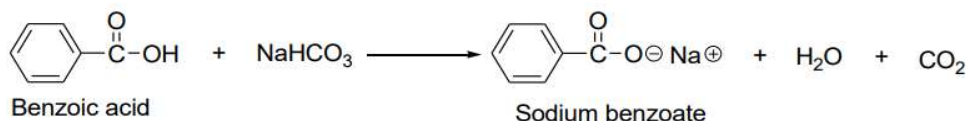
Reaction with water



Reaction of Carboxylic Acids with Sodium Bicarbonate

Most carboxylic acids ($\text{pK}_a \approx 5$) are stronger acids than carbonic acid (H_2CO_3) (pK_a 6.4). Consequently they displace carbonic acid from its salts (hydrogen carbonates). The most reliable test for carboxylic acids employs NaHCO_3 leading to evolution of CO_2 . This is commonly called the bicarbonate test for carboxylic acids.

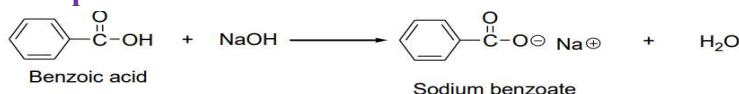
Example:



Reactions of Carboxylic Acids with Strong Bases

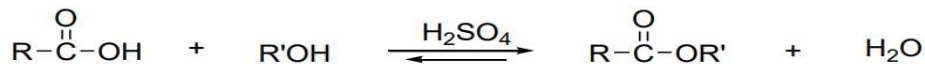
Bases such as metal hydroxides (NaOH and KOH) and amines abstract the acidic proton on carboxylic acids to form carboxylate salts.

Example

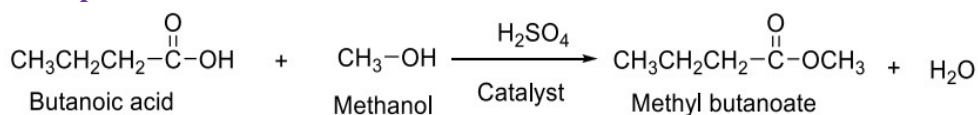


Acid-Catalysed Esterification of Carboxylic Acids

The traditional method for converting carboxylic acids to esters is through an acid-catalyzed esterification in the presence of an alcohol: Commonly referred to as the Fischer esterification.



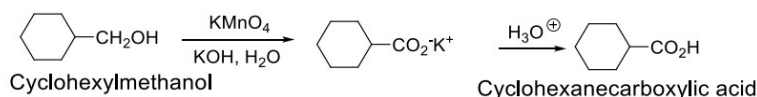
Example:



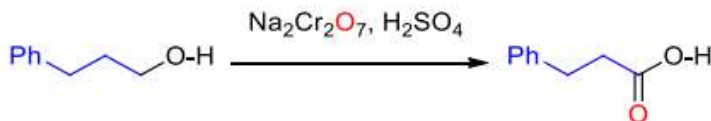
Synthesis of Carboxylic Acids

A. Oxidation of Primary Alcohols to Carboxylic Acids

The best conditions for the oxidation of primary alcohols to carboxylic acids is under the basic conditions employing potassium permanganate.

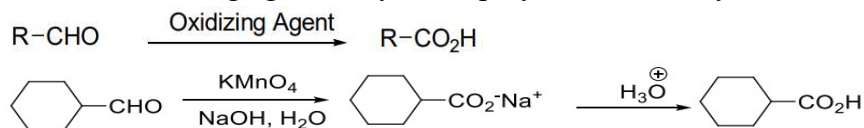


Grade-11 Chemistry Short Note With Practice Questions



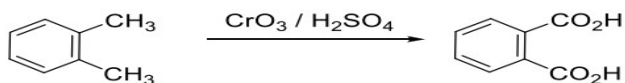
B. Oxidation of Aldehydes to Carboxylic Acids

Aldehydes can be oxidised to carboxylic acids by a variety of oxidizing agents. Both strong and mild oxidizing agents may be employed successfully.



C. Side-Chain Oxidation of Alkylbenzenes to Benzoic Acids

Oxidation of alkylbenzenes using strong oxidizing agents provides benzoic acids. The entire alkyl chain, regardless of its length, is oxidised to a carboxyl (-COOH) group



Fatty acids

- ✓ These are monocarboxylic organic acids, which usually contain an even number of carbon atoms.
- ✓ The general formula of fatty acid is R-COOH
- Example:** $\text{CH}_3 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{CH}_2 - \text{COOH}$
- ✓ They are further classified into: Saturated & Unsaturated fatty acids according to absence or presence of double bonds.

A. Saturated fatty acids

- They have no double bonds.
- They have the general formula : $\text{CH}_3 - (\text{CH}_2)_n - \text{COOH}$.
- They are further classified, according to the number of carbon atoms, into: Short chain & Long chain fatty acids.

	Short chain FA	Long chain FA
Number of carbon atoms	Less than 10 carbons (2C to 10C)	More than 10 carbons
Consistency at room temperature	Liquid	Solid
Volatility	Volatile	Nonvolatile
Solubility in water	Soluble	Insoluble
Examples	Acetic (2 C) Butyric (4 C)	Palmitic (16 C) Stearic (18 C)

Grade-11 Chemistry Short Note With Practice Questions

The following table shows the formula of the most common saturated fatty acids.

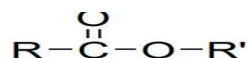
Common name	Formula
Acetic acid (2C)	$\text{CH}_3 - \text{COOH}$
Butyric acid (4C)	$\text{CH}_3 - (\text{CH}_2)_2 - \text{COOH}$
Caproic acid (6C)	$\text{CH}_3 - (\text{CH}_2)_4 - \text{COOH}$
Palmitic acid (16C)	$\text{CH}_3 - (\text{CH}_2)_{14} - \text{COOH}$
Stearic acid (18C)	$\text{CH}_3 - (\text{CH}_2)_{16} - \text{COOH}$
Arachidic acid (20C)	$\text{CH}_3 - (\text{CH}_2)_{18} - \text{COOH}$
Lignoceric acid (24C)	$\text{CH}_3 - (\text{CH}_2)_{22} - \text{COOH}$

B. Unsaturated fatty acids

- They have one or more double bonds.
- Most of the double bonds are of cis type

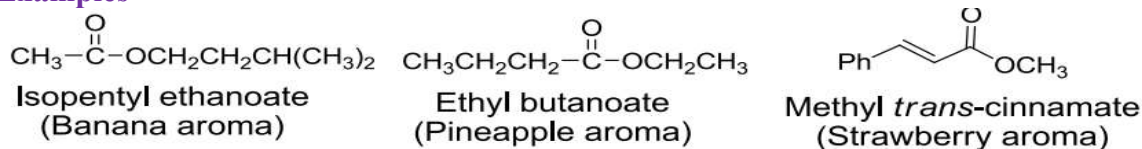
Esters

Structurally, an ester is a compound that has an alkoxy (OR) group attached to the carbonyl group.



R may be H, alkyl or aryl, while R' may be alkyl or aryl only. Esters are widespread in nature. Many of the fragrances of flowers and fruits are due to the esters present. Ethyl butanoate is the chief component that accounts for the pineapple-like aroma and flavour of pineapples.

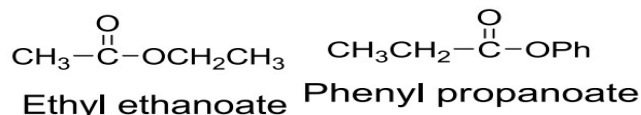
Examples



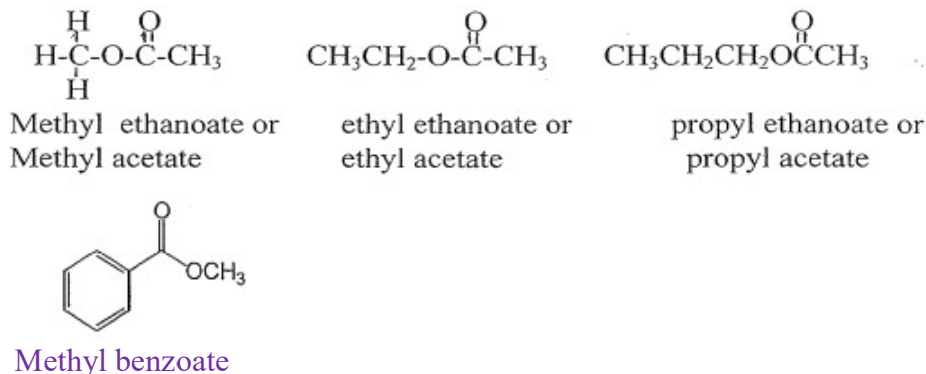
Nomenclature of Esters

- ✚ Names of esters consist of two words that reflect the composite structure of the ester. The first word is derived from the alkyl group of the alcohol component, and the second word from the carboxylate group of the carboxylic acid component of the ester.
- ✚ The name of the carboxylate portion is derived by substituting the -ic acid suffix of the parent carboxylic acid with the -ate suffix.
- ✚ The alkyl group is cited first followed by the carboxylate group separated by a space. An ester is thus named as an alkyl alkanoate.

Examples

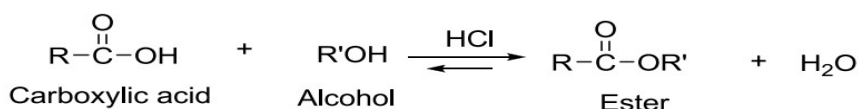


Grade-11 Chemistry Short Note With Practice Questions



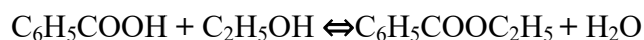
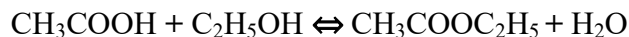
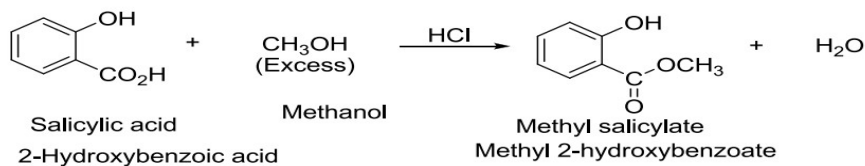
Synthesis of Esters

Acid-Catalysed Esterification of a Carboxylic Acid and an Alcohol



The acid-catalysed reaction of carboxylic acids and alcohols provides esters. Typically, a catalytic amount of a strong inorganic (mineral) acid such as H_2SO_4 , HCl and H_3PO_4 is used.

Examples:



Physical properties

- ❖ Most of the esters are liquids with much lower boiling point than those of the acids or alcohols of nearly equal weight, this is due to the absence of the polar hydroxyl group which found in alcohols and acids and leads to the association of the alcohol and carboxylic acid molecules with hydrogen bonds.
- ❖ Esters have a neutral effect on litmus.
- ❖ The boiling point of the ester is less than the boiling point of the acid and alcohol forming it due to the absence of polar hydroxyl group (presents in alcohols and acids) which has the ability to form hydrogen bonds between molecules and water.
- ❖ The solubility degree of ester in water is less than that of the corresponding acid due to the absence of polar hydroxyl group (presents in alcohols and acids) which has the ability to form hydrogen bonds between molecules and water.

Reactions of Ester

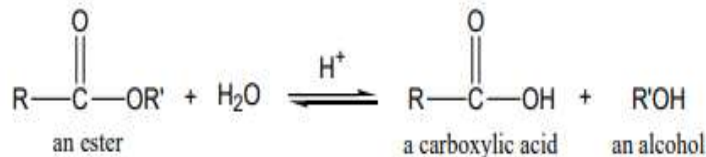
Hydrolysis of esters: Alcohol and acid are produced from the hydrolysis of ester, this reaction reverse to ester formation, Hydrolysis may take place by the use of dilute mineral acids as a as

Grade-11 Chemistry Short Note With Practice Questions

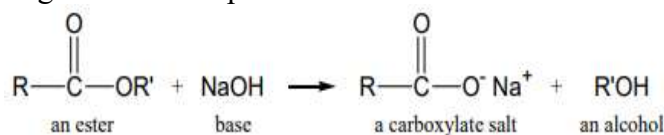
a catalyst and is called acid hydrolysis, Dilute mineral acid (H^+) is used to prevent the reversible reaction.

This is essentially the reverse reaction of the synthesis of esters from carboxylic acids and alcohols.

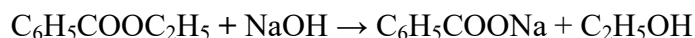
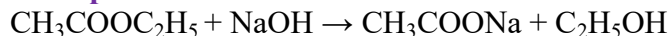
Example: $\text{CH}_3\text{COOC}_2\text{H}_5 + \text{H}_2\text{O} \rightarrow \text{CH}_3\text{COOH} + \text{C}_2\text{H}_5\text{OH}$



Hydrolysis of esters may also be carried out by heating with aqueous alkalis, to produce the alcohol and the salt of the acid, this is called alkaline-hydrolysis or saponification, (since soap is the sodium salt of high carboxylic acids), NaOH is added to react with the produced acid converting it to salt and prevent the reversible reaction.

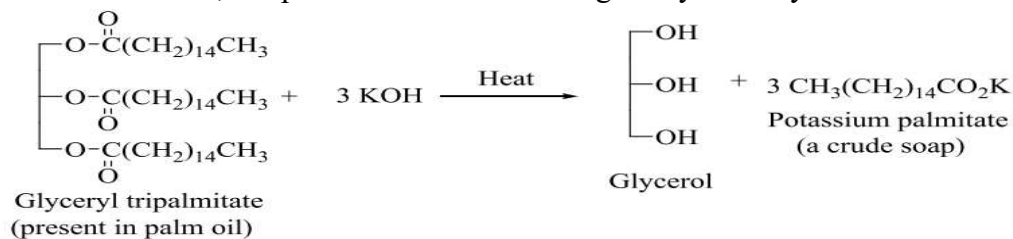


Examples:

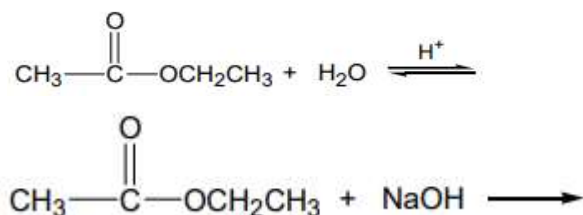


For historical reasons, ester hydrolysis in aqueous hydroxide (KOH or NaOH) is called saponification because it was used in the manufacture of soap through the reaction of oils or fats (triesters or triglycerides) with lye (contains mainly KOH).

- Hydrolysis of esters is the reaction of the ester with water (in an acidic medium) to form acid and alcohol
- Saponification is the heating of ester with aqueous alkalis to produce the alcohol and the salt of the acid, Soap is the sodium salt of high fatty carboxylic acids.



Exercise: Complete the following reactions

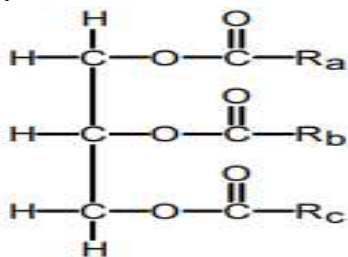


Fats and Oils

Fats and oils belong to a group of biological substances called lipids. Lipids are biological chemicals that do not dissolve in water. Fats differ from oils only in that they are solid at room

Grade-11 Chemistry Short Note With Practice Questions

temperature, while oils are liquid. Fats and oils share a common molecular structure, which is represented by the formula below.



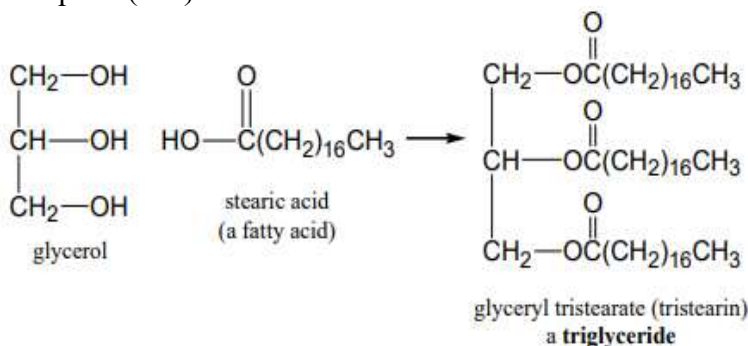
This structural formula shows that fats and oils contain three ester functional groups. Fats and oils are esters of the tri-alcohol, glycerol (or glycerine). Therefore, fats and oils are commonly called triglycerides, although a more accurate name is triacylglycerols. One of the reactions of triglycerides is hydrolysis of the ester groups.

Fats and oils:

- ❖ They are esters of fatty acids with glycerol.
- ❖ They are similar chemically, but they differ in physical properties as oils are liquid while fats are solid at room temperature.
- ❖ They are called triglycerides because they are triesters formed of glycerol and 3 fatty acids.

Triglycerides

- Fats and oils are both triglycerides; the only difference is that fats are solids at room temperature and oils are liquids at room temperature.
- Long-chain carboxylic acids, often referred to as fatty acids, are stored by living organisms by combining them with glycerol to produce tri-esters called **triglycerides**.
- Triglycerides at room temperature are usually either solids or semi-solids (fats), or viscous liquids (oils).

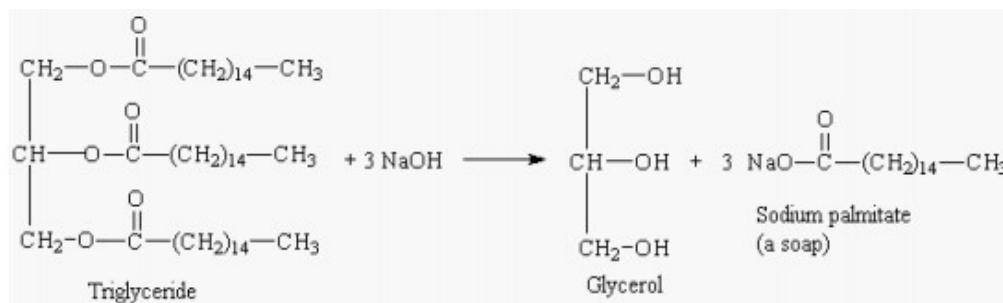


Hydrolysis of triglycerides

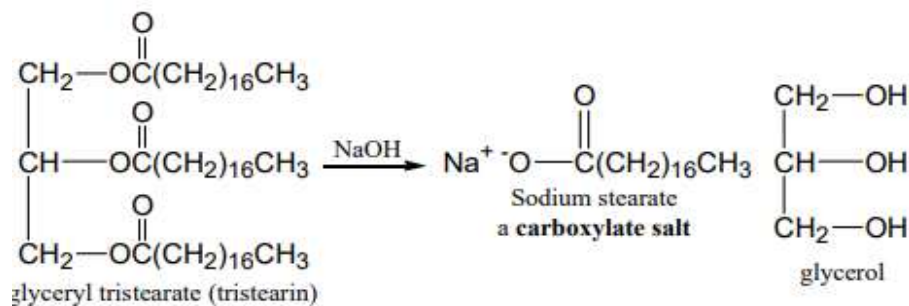
The hydrolysis of triglycerides can be carried out with base and this is the common way of doing it on an industrial scale. The hydrolysis and neutralization are carried out simultaneously and produces soap in a one step reaction called **saponification**.

Triglycerides can be broken apart under basic conditions (a **saponification reaction**) to produce long-chain carboxylate salts

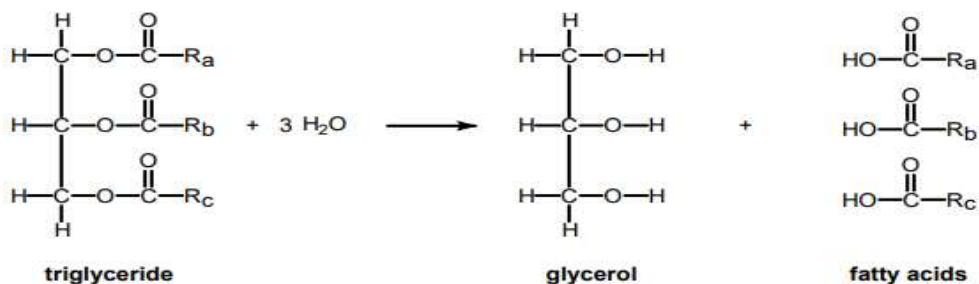
Grade-11 Chemistry Short Note With Practice Questions



Example:



Triglycerides (fats) can be hydrolyzed to produce glycerol and 3 fatty acids in the presence of acid and heat or with a suitable lipase enzyme under biological conditions.



This hydrolysis reaction produces glycerol and fatty acids, which are carboxylic acids derived from fats and oils. In the fatty acids, R_a , R_b , and R_c , represent groups of carbon and hydrogen atoms in which the carbon atoms are attached to each other in an unbranched chain.

The hydrolysis reaction is promoted by acids and by bases. When a strong base such as NaOH (lye) is used, the product contains salts of the fatty acids. These salts of fatty acids are the functional ingredient in soap. The ingredients lists of some soaps include sodium tallowate, a generic name for the mixture of fatty acid salts obtained from tallow (animal fat), and sodium cocoate, obtained from coconut oil.

Hardening of oils

Chemically, the hydrogenation of oils is the reduction of the double bonds in unsaturated fatty acids to single saturated bonds, by the reaction of hydrogen gas in the presence of a metal catalyst.

Rancidity

Rancidity is a condition in which fat attains a bad taste and disagreeable odour.

Grade-11 Chemistry Short Note With Practice Questions

Types of rancidity

There are 2 types of rancidity:

1. Hydrolytic rancidity
2. Oxidative rancidity

Hydrolytic rancidity: Fats are hydrolyzed into glycerol and fatty acids in presence of moisture, warm temperature and bacterial enzymes.

Oxidative rancidity: It occurs by oxidation of unsaturated fatty acids present in fats and oils forming lipid peroxides, fatty aldehydes, ketones and short chain fatty acids.

Rancidity is predisposed by: light, moisture, warm and temperature

Rancidity leads to:

- ✚ Fats and oils attain bad taste.
- ✚ Fats and oils attain disagreeable odour.
- ✚ Production of toxic compounds as lipid peroxides, aldehydes and ketones.

Prevention of rancidity:

- Avoid exposure of fats to light, moisture and high temperature.
- Addition of antioxidants to fats and oils.

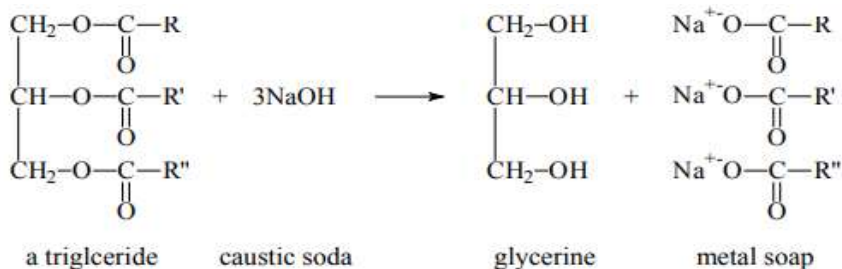
Soaps and detergents

Soaps are water-soluble sodium or potassium salts of fatty acids. Soaps are made from fats and oils, or their fatty acids, by treating them chemically with a strong alkali.

Saponification of fats and oils is the most widely used soapmaking process.

Soaps are the product of the reaction between a fat and sodium hydroxide:

fat + 3NaOH → glycerine + 3 soap



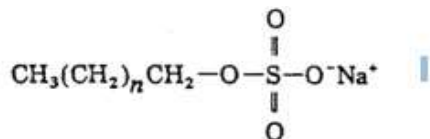
Detergent

Detergents are complex formulations that contain more than 25 different ingredients, which can be categorized into the following main groups:

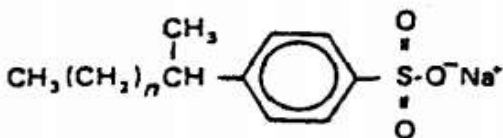
1. Surfactants
2. Bleaching agents
3. Builders
4. Additives

Detergents are structurally similar to soaps, but differ in the water-soluble portion. Three examples of detergents are shown below.

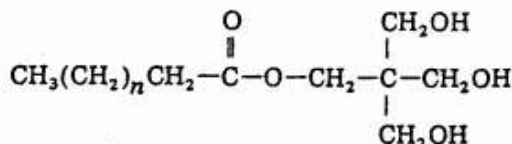
Grade-11 Chemistry Short Note With Practice Questions



a sodium alkyl sulfate



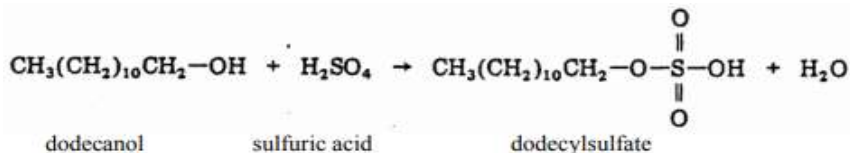
a sodium alkylbenzene sulfonate



a glycerol (a non-ionic detergent)

Preparation of a detergent

A synthetic detergent, a sodium alkyl sulfate called sodium dodecylsulfate, will be prepared by reacting dodecyl alcohol (dodecanol) with sulfuric acid.

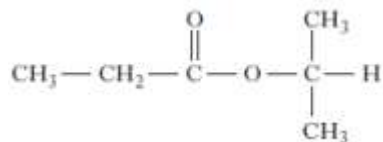


When a soap or detergent is added to water, a polar solvent, the molecules form clusters, known as **micelles**, in which the polar ends of the molecules are on the outside of the cluster and the non-polar ends are in the middle.

The cleaning action of both soaps and detergents results from their ability to emulsify or disperse water-insoluble materials (dirt, oil, grease, etc.) and hold them in suspension in water. This ability comes from the molecular structure of soaps and detergents. When a soap or detergent is added to water that contains oil or other water-insoluble materials, the soap or detergent molecules surround the oil droplets. The oil or grease is “dissolved” in the alkyl groups of the soap molecules while the ionic end allows the micelle to dissolve in water. As a result, the oil droplets are dispersed throughout the water (this is referred to as emulsification) and can be rinsed away.

Grade-11 Chemistry Short Note With Practice Questions

9. Which of the following is true about the chemical decomposition of Fats and Oils?
- A. Oxidative rancidity occurs when the double bond of an unsaturated fatty acid react chemically with hydrogen
 - B. Microbial rancidity refers to a process in which microorganisms use their enzymes to break down chemical structure in the fat or oil
 - C. Rancidity result in a pleasant odor and fruity smell
 - D. The spoilage of fat and oil is hardening of oil
10. The chemical makeup of fats is:
- A. Esters of glycerol with three identical saturated fatty acids.
 - B. Esters of glycerol with three predominantly saturated fatty acids.
 - C. Esters of glycerol with three predominantly unsaturated fatty acids.
 - D. Esters of glycerol with three identical unsaturated fatty acids.
11. Example of saturated fatty acid:
- A. Arachidonic acid B. Oleic acid C. palmitic acid D. Linoleic acid
12. Rancidity is a process of oxidation concerning :
- A. Short chain fatty acid C. Saturated fatty acid
 - B. Hydroxy fatty acid D. Unsaturated fatty acid
13. Which reaction can be used to convert oils into fats?
- A. Esterification C. Saponification
 - B. Dehydration D. Hydrogenation
14. Soaps are .
- A. Alkali metal salts of long chain carboxylic acids
 - B. Alkali metal salts of glycerol
 - C. Esters of a long chain fatty alcohol with a long chain fatty acid
 - D. Long chain fatty alcohols
15. An ester has the structural formula:



On hydrolysis, the ester would produce:

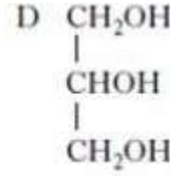
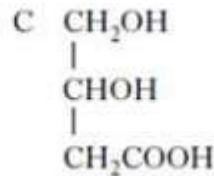
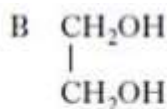
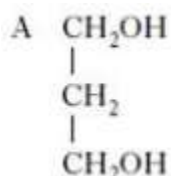
- A. Pthanoic acid and propan-1-ol C. Ethanoic acid and propan-2-ol
 - B. propanoic acid and propan-1-ol D. Propanoic acid and propan-2-ol.
16. An ester has the following structural formula, $\text{CH}_3\text{CH}_2\text{CH}_2\text{COOCH}_2\text{CH}_3$.
The name of this ester is:
- A. Propyl propanoate C. Ethyl butanoate
 - B. Butyl ethanoate D. Ethyl propanoate.
17. Esters are formed by the reaction between which two functional groups?
- A. A hydroxyl group and a carboxyl group
 - B. A hydroxyl group and a carbonyl group
 - C. A hydroxide group and a carboxyl group
 - D. A hydroxide group and a carbonyl group

Grade-11 Chemistry Short Note With Practice Questions

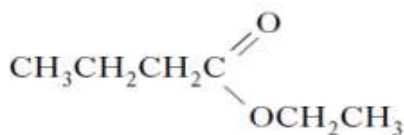
18. Oils are generally:

- A. Solid at room temperature and contain a high proportion of unsaturated molecules
- B. Solid at room temperature and contain a high proportion of saturated molecules
- C. Liquid at room temperature and contain a high proportion of unsaturated molecules
- D. Liquid at room temperature and contain a high proportion of saturated molecules.

19. What is the structural formula for glycerol?



20. Rum flavouring is based on the compound with the formula shown.



It can be made from:

- A. Ethanol and butanoic acid
- B. Butanol and methanoic acid
- C. Propanol and ethanoic acid
- D. Propanol and propanoic acid

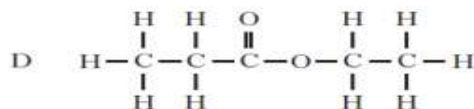
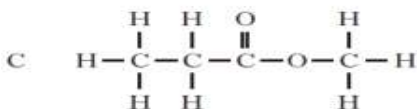
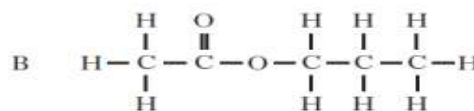
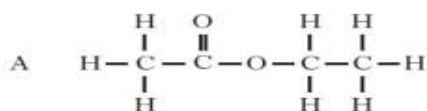
21. Which of the following reactions takes place during the „hardening“ of vegetable oil?

- A. Addition
- B. Hydrolysis
- C. Dehydration
- D. Oxidation

22. Fats have higher melting points than oils because comparing fats and oils

- A. Fats have more hydrogen bonds
- B. Fat molecules are more saturated
- C. Fat molecules are more loosely packed
- D. Fats have more cross-links between molecules

23. Hydrolysis of an ester gave an alcohol and a carboxylic acid both of which had the same molecular mass of 60. The structure of the ester was:



24. Fats are formed by the condensation reaction between glycerol molecules and fatty acid molecules. In this reaction the mole ratio of glycerol molecules to fatty acid molecules is:

- A. 1 : 1
- B. 1 : 2
- C. 1 : 3
- D. 3 : 1

25. Which of the following fatty acids is unsaturated?

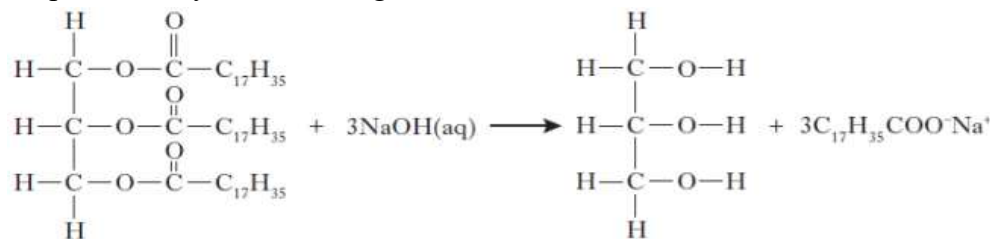
- A. $\text{C}_{19}\text{H}_{39}\text{COOH}$
- B. $\text{C}_{17}\text{H}_{31}\text{COOH}$
- C. $\text{C}_{21}\text{H}_{43}\text{COOH}$
- D. $\text{C}_{13}\text{H}_{27}\text{COOH}$

26. Which molecule is a fatty acid?

- A. $(\text{CH}_3)_2\text{CH}(\text{CH}_2)_3\text{COOH}$
- C. $\text{CH}_3(\text{CH}_2)_7\text{CH}=\text{CH}(\text{CH}_2)_7\text{COOH}$

Grade-11 Chemistry Short Note With Practice Questions

27. Triglycerides belong to which general class of organic molecules.
A. Ketones B. Esters C. Alcohols D. Carboxylic acids
28. The saponification reaction used to form soaps can be more specifically described as:
A. Dehydration C. Acid hydrolysis.
B. Hydrogenation D. Basic hydrolysis
29. In chemical terms, soaps can best be described as:
A. Simple esters of fatty acids C. Salts of carboxylic acids
B. Long chain acids D. Bases formed from glycerol
30. The hydrocarbon end of a soap molecule is
A. Hydrophobic and attracted to grease
B. Hydrophilic and attracted to water
C. Hydrophobic and attracted to water hydrophilic and attracted to grease
D. Neither hydrophobic nor hydrophilic
31. Which of the following is the residual product in the formation of soap?
A. Glyceraldehyde B. Glycerol C. Glycerine D. Acrylonitrile
32. Which of the following is a typical soap molecule?
A. Calcium stearate C. Potassium permanganate
B. Sodium bicarbonate D. Sodium stearate
33. Soaps are produced by the following reaction



This reaction is an example of:

- A. Condensation B. Esterification C. Hydrolysis D. oxidation

Grade-11 Chemistry Short Note With Practice Questions

Answers to Review Exercise

Unit 5

1. C	9. B	17. C	25. B	33. B
2. B	10. B	18. D	26. C	34. D
3. D	11. D	19. D	27. C	35. D
4. B	12. A	20. D	28. D	36. D
5. C	13. A	21. D	29. A	37. C
6. E	14. B	22. C	30. A	
7. C	15. D	23. A	31. D	
8. D	16. B	24. B	32. C	

Unit 6

1. D	9. B	17. A	25. B
2. C	10. B	18. C	26. C
3. A	11. C	19. D	27. B
4. C	12. D	20. A	28. D
5. D	13. D	21. A	29. C
6. D	14. A	22. C	30. A
7. D	15. D	23. B	31. B
8. A	16. C	24. C	32. D
			33. C

Dear Students:

KEEP SAFE, BE HEALTHY AND STAY AT HOME DURING THIS UNCERTAIN TIME.