CHEMISTRY

10th grade

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МИНИСТЕРСТВО НА ОБРАЗОВАНИЕТО И НАУКАТА

НАЦИОНАЛНА ПРОГРАМА

„Разработване на учебни помагала за обучение по общообразователни учебни предмети на чужд език, оценяване и одобряване на проекти на учебни помагала за подпомагане на обучението, организирано в чужбина, на проекти на учебници и на проекти на учебни комплекти“

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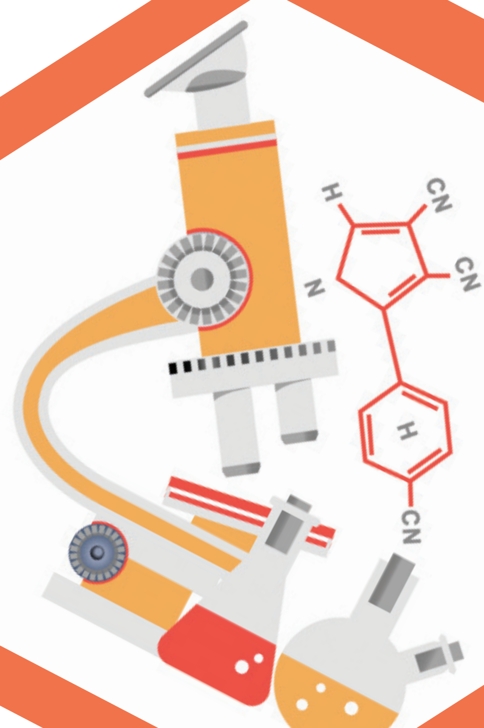
CHEMICAL PROCESSES



1 CHEMICAL PROCESSES

1. Chemical and physical changes of matter

Chemical and physical changes are two different types of changes in matter.

➢ Physical changes are changes in the shape, texture, size, in the states of matter (melting, freezing, evaporation, condensation, sublimation). Matter alters its forms, but not its chemical identity, it stays the same kind of matter. For example, when liquid water freezes, it changes to ice. However, the water molecules have the same composition as each water molecule contains two hydrogen atoms and one oxygen atom covalently bonded.

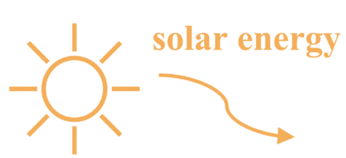
➢ Chemical changes occur when matter changes into a different kind of matter. For example, one substance is turned into one or more different substances (chemical decomposition) or two substances combine to form a new compound (chemical synthesis). Apart from these processes (also called chemical reactions), there are other types of chemical reactions, such as neutralisation, combustion, to name just a few. For example, the methane combustion is a chemical process in which the reactants (methane and oxygen) are converted to carbon dioxide and water. The process requires a flame or initiating spark. It involves breaking chemical bonds between the atoms in methane molecules (C - H bond) and oxygen molecules (O - O bond) and forming new bonds in the molecules of carbon dioxide (C - O bond) and water (H - O bond). The reaction can be expressed by the following chemical equation:

$$CH_4 + 2O_2 \rightarrow CO_2\uparrow + 2H_2O$$

2. Systems and surroundings

Chemical processes take place inside living organisms, in industrial plants, test tubes or flasks. They are accompanied by transfer of energy. Any chemical process can be defined as a system – the substance or collection of substances. The surroundings are everything around the system which is not directly involved in the reaction.

An example of a chemical system is the photosynthesis reaction. The system or the chemical reaction contains carbon dioxide, water, light, sugar, and oxygen. The surroundings are everything outside the system, the rest of the universe. The equation of the reaction is as follows.



$$6CO_2 + 6H_2O + light \xrightarrow{photosynthesis} C_6H_{12}O_6 + 6O_2$$

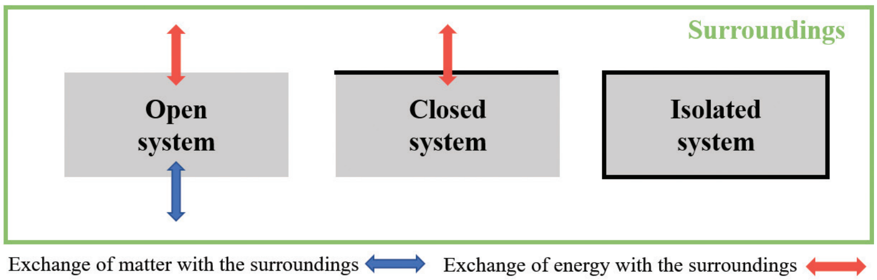
The exchange of energy between the system and its surroundings is in the form of heat. There are three types of systems:

I. **An open system** can exchange energy and matter (mass) with its surroundings. Most chemical reactions occur in an open system.

II. **A closed system** can exchange only energy with its surroundings, not matter. For example: a sealed flask of liquid, an unopened soda can, a closed bottle of water.

III. An isolated system can exchange neither energy, nor matter with its surroundings. For example, a hot tea in a thermos flask.

The systems are depicted in the figure below.



Test your knowledge:

1. **Which of the following is a chemical change?**
   a. Chopping woods.
   b. Lightn ing a match.
   c. Boiling water.
   d. Ice cream melting.
2. **Which of the following is a physical change?**
   a. Rusting of iron.
   b. Burning leaves.
   c. Cutting diamonds.
   d. Dropping a vitamin C tablet in water.
3. **Putting a lid on a saucepan is an example of:**
   a. A closed system
   b. An isolated system.
   c. An open system.
   d. None of them.
4. **Determine whether each of the following systems is open, closed, or isolated. Fill in the box to the right of the example**

№	An example	Open System, Closed System, or Isolated System
1	Boiling saucepan of water without a lid.	
2	A cold water in a thermos flask.	
3	A fish tank.	
4	A river.	
5	Biological organisms.	
6	A balloon filled with helium.	

5. **Why is the Earth regarded a closed system?**

1. Enthalpy

The word "enthalpy" comes from the Greek word "enthalpos" which means "to put heat inside". In simple terms, it is the heat content of a system at constant pressure because the reacting system is opened to the atmosphere. Enthalpy is the amount of energy within a system. It is denoted by the capital letter $\boldsymbol{H}$. When a chemical reaction proceeds, changes in the heat that passes into or out of the system correspond to the enthalpy change. $\Delta \boldsymbol{H}$ is the accepted symbol for the enthalpy change. Unlike the enthalpy ($\boldsymbol{H}$) that cannot be measured directly, the enthalpy change ($\Delta \boldsymbol{H}$) can be measured from the temperature change in a particular chemical reaction. It has units of *kJ* or *kJ/mol*.

2. Exothermic and endothermic reactions

Exothermic reactions transfer energy from the reacting molecules to the surroundings so that the temperature in the surroundings increases and they become warmer than before. The products of an exothermic reaction have less energy than the reactants because energy has been transferred from the reaction to the surroundings. The enthalpy change is negative ($\Delta \boldsymbol{H} < \boldsymbol{0}$). In the figure on the left below, the energy profile for an exothermic reaction is shown.

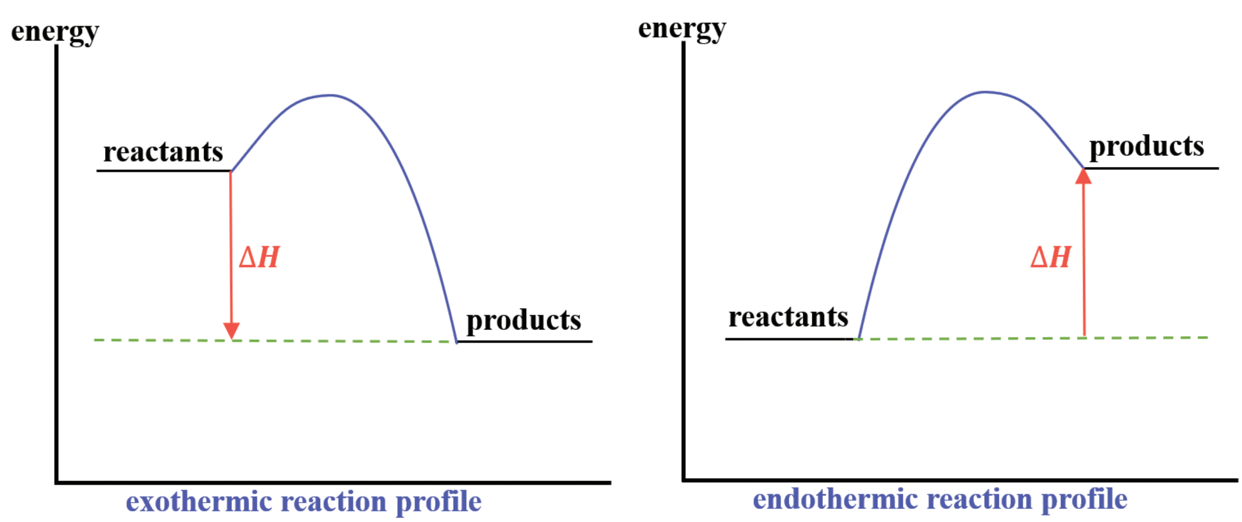


Figure 2.1 Energy profile (a reaction coordinate) diagrams for an exothermic and endothermic reactions

Endothermic reactions take in energy from their surroundings, so the temperature of the surroundings decreases, and they become colder than before. The products of an endothermic reaction have higher energy than the reactants. This is because energy has been taken from the surroundings. The enthalpy change is positive ($\Delta \boldsymbol{H} > \boldsymbol{0}$). The energy profile for an endothermic reaction is represented in the figure on the right above.

The energy exchange between any chemical system and surroundings is visualised in *Figure 2.2* below. The figure on the left shows the exchange that takes place when an exothermic reaction proceeds. Combustion, oxidation, and neutralisation, for example, are exothermic. The image on the right represents an endothermic process and its exchange. For example, melting ice cubes, evaporating liquid water, decomposition of calcium carbonate, photosynthesis, etc.

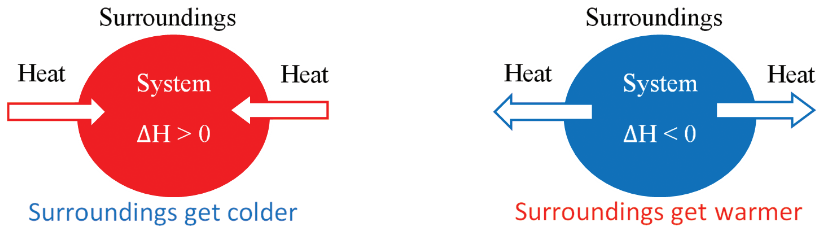


Figure 2.2

Test your knowledge:

1. Choose the correct word to complete the following text:

All chemical reactions involve ___________ (transfer/absorption) of heat. When a reaction ___________ (ends/occurs), the chemical bonds between the atoms of the (products/reactants) are broken first. The process is ___________ (exothermic/endothermic). Energy is ___________ (absorbed/released) when the new bonds between the ___________ (products/reactants) are formed. The process is ___________ (exothermic/endothermic). Whether the reaction is ___________ (exothermic/endothermic) or ___________ (exothermic/endothermic) depends on the difference between the ___________ (reaction/energy) needed for breaking bonds and the energy needed for forming new bonds.

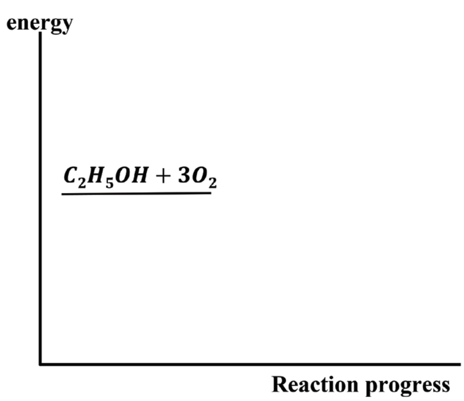
2. Decide whether each reaction is exothermic or endothermic and explain your choice. Fill in the table below.

Reaction (exothermic or endothermic)	Starting temperature °C	Final temperature °C
X	22	34
Y	25	20
Z	24	19

3. The ethanol combustion reaction releases energy and $\Delta H < 0$. The equation for the reaction is as follows:

$$C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$$

Complete the energy profile diagram for the reaction.



3 ENTHALPY CHANGES

Thermochemistry is a branch of chemistry that studies the energy changes that occur during chemical reactions.

1. Standard conditions

In chemistry, the standard conditions under which the reactions proceed are:

✓ a pressure of 100 *kPa* or 1 *atm*

✓ temperature of 298 *K* (25°C)

This set of conditions is important because the standard enthalpy of reaction is the change in enthalpy or heat of chemical reactions under these conditions.

2. Standard enthalpy changes:

- **Standard enthalpy of formation**

The enthalpy change is associated with the formation of one mol of a compound in its standard state from the elements composing that compound in their standard states. The symbol of standard enthalpy of formation is $\boldsymbol{\Delta H_f^{\ominus}}$. It is measured in *kJ/mol*.

✓ Δ = a change in enthalpy (indicates how much heat is released or absorbed by the system)

✓ $\ominus$ = the symbol indicates the standard conditions

✓ $\boldsymbol{f}$ = the indicates that the substance is formed from its elements

The sample values of standard enthalpies of formation of various compounds are provided in tables of the data booklets. We need to bear in mind that the standard enthalpy of formation of a pure element is zero. When a chemical element exists as more than one form then the more stable form of the element (with the lowest enthalpy) is chosen and determined for zero. For example, carbon exists as graphite and diamond. Graphite is more stable than diamond, so it has $\Delta H_f^{\ominus} = 0$.

- **Thermochemical equations**

Energy changes that take place during a chemical reaction are represented by thermochemical equations. The equations include the enthalpy change (ΔH) of a rection and the state of the reactants and products (whether they are solids, liquids, gases, or aqueous solutions). For example, carbon reacts with oxygen and a great amount of heat releases to the surroundings. The equation of the thermochemical reaction is as follows:

$$C_{(s)} + O_{2(g)} \rightarrow CO_{(g)} \ \Delta H_f^{\ominus} = -393.5 \ kJ/mol,$$

where: $_{(s)}$ stands for a solid; $_{(g)}$ stands for a gas.

Graphite (C) and oxygen (O_2) are in their elementally stable forms, so they have a standard enthalpy of formation zero. The enthalpy of formation of carbon dioxide is $\Delta H_f^{\ominus} = -393.5 \ kJ/mol$ which means that the reaction is exothermic.

- **Standard enthalpy change of combustion**

The standard enthalpy change of combustion of a substance, $\Delta H_C^{\ominus}$, is the enthalpy change that occurs when 1 *mol* of a substance completely burns in oxygen under standard conditions. It is measured in $kJ.mol^{-1}$ (kJ/mol).

For example, ethyne (also known as acetylene) burns with smoldering flame in air. Upon combustion in an oxygen atmosphere, the flame is at extremely high temperature (over 2700°C). This property of ethyne is used in the acetylene burners for cutting and welding metals.

The equation of the reaction is as follows:

$$2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(l)} \qquad \Delta H_c^{\ominus} = -890\ kJ/mol,$$

where: $_{(s)}$ stands for a solid, $_{(g)}$ stands for a gas; $_{(l)}$ stands for a liquid. The enthalpy of combustion of acetylene, - 890 kJ/mol, is the amount of heat produced when one mol C_2H_2 burns completely at 25°C and 1 atm.

Test your knowledge:

1. **What is the standard enthalpy of formation of any chemical element in its standard state?**
2. **Arrange the following compounds in order of increasing their thermal stability.**

The compound	$NaCl_{(s)}$	$NaHCO_{3(s)}$	$Na_2CO_{3(s)}$	$NaF_{(s)}$
$\Delta H_f^{\ominus}$ (kJ/mol)	- 411.0	-947.7	-1130.90	-571.0

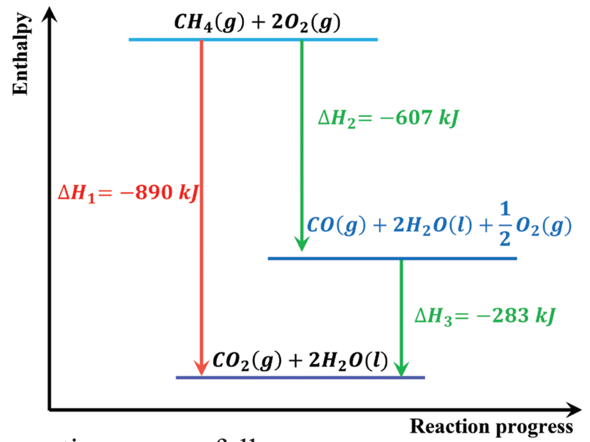
   a. $NaCl_{(s)} > NaF_{(s)} > NaHCO_{3(s)} > Na_2CO_{3(s)}$
   b. $Na_2CO_{3(s)} > NaHCO_{3(s)} > NaF_{(s)} > NaCl_{(s)}$
   c. $NaHCO_{3(s)} > Na_2CO_{3(s)} > NaCl_{(s)} > NaF_{(s)}$
   d. $Na_2CO_{3(s)} > NaHCO_{3(s)} > NaCl_{(s)} > NaF_{(s)}$
3. **Why is the standard enthalpy of formation of sodium equal to zero?**
   a. Because sodium is a reactive metal.
   b. Because sodium has 11 valent electrons in its valence shell.
   c. Because of the standard enthalpy of formation of sodium.
   d. Because the sodium is in its standard state and there is no need of a reaction to form it.

4 HESS'S LAW. APPLICATIONS OF HESS'S LAW

1. Hess's Law

Hess's Law states that if a reaction proceeds by more than one route, at the same initial and final conditions, the total enthalpy change is the same for each route.

For example, the energy level diagram below shows two routes for converting methane (CH_4) to carbon dioxide (CO_2).



The equations for the reactions are as follows:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{(g)} + 2H_2O(l) + \frac{1}{2}\ O_{2(g)}\quad \Delta H_2 = -607\ kJ$$

$$CO_{(g)} + 2H_2O_{(l)} + \frac{1}{2}\ O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}\quad \Delta H_3 = -283\ kJ$$

$$CH_{4(g)} + O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(l)}\ \text{(overall reaction)},\ \Delta H_1 = -890\ kJ$$

By Hess's Law, the total enthalpy change for the reaction is the sum of all changes.

$$\Delta H_1 = \Delta H_2 + \Delta H_3,\quad -890 = -607 + (-283).$$

2. Applications of Hess's Law

Hess's Law allows the enthalpy change for a reaction to be calculated, if the values for the standard enthalpies of formation, $\Delta H_f^{\ominus}$, of the reactants and products are previously determined.

The enthalpy change for a chemical reaction is the difference between the sum of the values for the standard enthalpies of formation of the products and the sum of the values for the standard enthalpies of formation of the reactants. The equation for the enthalpy change for a reaction is shown below:

(1) $\Delta H^{\ominus} = (\Sigma_n \cdot \Delta H_f^{\ominus})_{products} - (\Sigma_n \cdot \Delta H_f^{\ominus})_{reactants}$

To calculate $\Delta H^{\ominus}$ we need to follow several steps:

Step 1. Write a balanced thermochemical equation for the given reaction.

Step 2. Place the products and reactants in the equation (1), considering the moles of the compounds involved in the reaction.

Step 3. Look up the enthalpy of formation values for the products and reactants in a table that provides values of standard enthalpies of formation of various compounds. Pay attention to the states of matter.

Step 4. Plug the values into the equation described in step 2 and calculate the enthalpy for the given reaction.

4 HESS'S LAW. APPLICATIONS OF HESS'S LAW

Solved example: Calculate the standard enthalpy change for the combustion of ethyne that produces carbon dioxide and water using the data below.

$\Delta H_f^\ominus(C_2H_{2(g)}) = 226.7$ kJ/mol, $\Delta H_f^\ominus(CO_{2(g)}) = -393.5$ kJ/mol, $\Delta H_f^\ominus(H_2O_{(l)}) = -285.8$ kJ/mol.

Step 1. $2C_2H_{2(g)} + 5O_{2(g)} \rightarrow 4CO_{2(g)} + 2H_2O_{(l)}$

Step 2. We place the products and reactants in the expression (1) to have the following equation:

$\Delta H = (4.\Delta H_f^\ominus(CO_{2(g)}) + 2.\Delta H_f^\ominus(H_2O_{(l)}) - (2.\Delta H_f^\ominus(C_2H_{2(g)}) + 5.\Delta H_f^\ominus(O_{(g)})),$

where $\Delta H_f^\ominus(O_{(g)}) = 0$

Step 3. We use the given thermochemical data about the enthalpies of formation of the products, CO_2 and H_2O, and the reactant, C_2H_2 and plug them into the above equation.

$\Delta H^\ominus = (4.(-393.5) + 2.(-285.8)) - 2.(226.7) = (-1574 - 571.6) - (453.4) = -2145.6 - 453.4$
$= -2\,599$ kJ

$\Delta H^\ominus = -2\,599$ kJ

Test your knowledge:

1. Calculate the standard enthalpy change for the combustion of copper (I) sulphide that produces copper (I) oxide and sulphur dioxide. Use the data below and the thermochemical equation for the reaction.

$$2Cu_2S_{(s)} + 3O_{2(g)} \rightarrow 2Cu_2O_{(s)} + 2SO_{2(g)}$$

Compound	$\Delta H_f^\ominus$ (kJ/mol)
$Cu_2S_{(s)}$	-79.5
$Cu_2O_{(s)}$	-168.6
$SO_{2(g)}$	-296.81

(Answer: – 771.82 *kJ*)

2. Carbon tetrachloride (4) is produced industrially by the reacting chlorine (2) with methane (4). Calculate the standard enthalpy change for the reaction, using the data below and the thermochemical equation for the reaction

$$CH_{4(g)} + 4Cl_{2(g)} \rightarrow CCl_{4(g)} + 4HCl_{(g)}$$

Compound	$\Delta H_f^\ominus$ (kJ/mol)
$CH_{4(g)}$	-74.8
$HCl_{(g)}$	-92.31
$CCl_{4(g)}$	-95.98

(Answer: – 390.42 *kJ*)

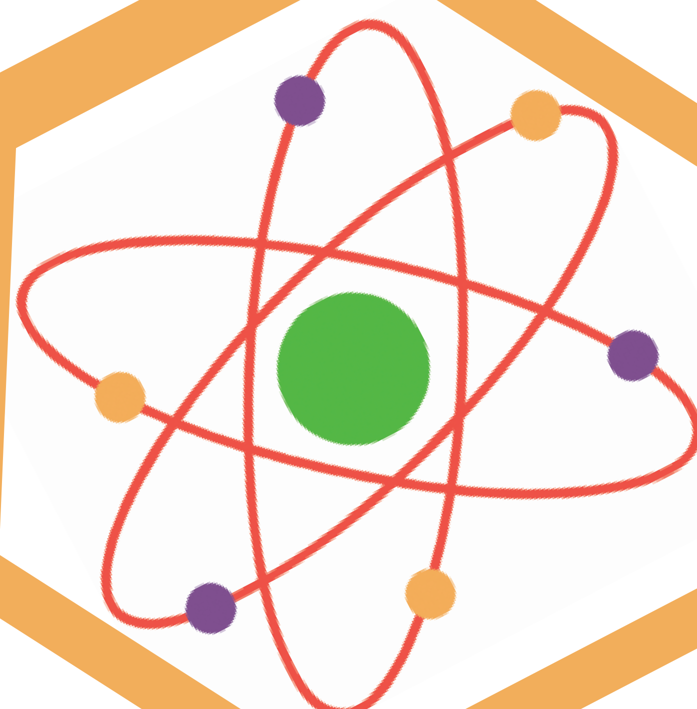
3. Ammonia reacts with oxygen to yield nitrogen monoxide and water vapour. Calculate the tandard enthalpy change of formation of ammonia, using the data below and the thermochemical equation for the reaction.

$$4NH_{3(g)} + 5O_{2(g)} \rightarrow 4NO_{(g)} + 6H_2O_{(g)} \quad \Delta H^\ominus = -893.6 \text{ kJ}$$

Substance	$\Delta H_f^\ominus$ (kJ/mol)
$H_2O_{(g)}$	-241.8
$NO_{(g)}$	+93.3

(Answer: –46 *kJ*)

CHEMICAL KINETICS



Chemical kinetics is the branch of physical chemistry that studies the rates of chemical reactions, the reaction mechanism by which they occur and the factors that change the rates.

1. Rates at chemical reactions.

The rate of a reaction is the speed at which the reaction takes place. It is defined as the change in concentration of a reactant or a product per unit time.

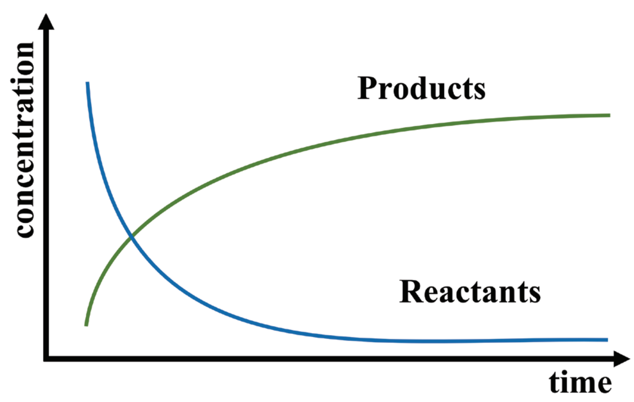
$$rate = \frac{change\ in\ concentration}{change\ in\ time}$$

The rate of a reaction is expressed by the following formula:

$\bar{v} = \pm\frac{\Delta c}{\Delta t}\left[\frac{mol}{L.s}\right]$, *where* $+\Delta c$ *is the change in the concentration(amount) of a product,* $-\Delta c$ *is the change in the concentration of a reactant per unit time* Δt.

The average rate of a chemical reaction is represented in a similar way: $\bar{v} = \pm\frac{\Delta c}{\Delta t} = \frac{c_2 - c_1}{t_2 - t_1}$

When a chemical reaction occurs, the concentrations of reactants decrease ($c_2 < c_1$), whereas the concentrations of products increase ($c_2 > c_1$). Therefore, the expressions for reactants are given a negative sign and for the products – a positive one.



For a general chemical reaction expressed by the equation: $a\,A + b\,B \rightarrow c\,C + d\,D$, *where* A *and* B *are reactants,* C *and* D *are products,* a, b, c *and* d *are their amounts, the reaction rate is determined by the following expression*:

$$v = -\frac{1}{a}.\frac{\Delta c(A)}{\Delta t} = -\frac{1}{b}.\frac{\Delta c(B)}{\Delta t} = \frac{1}{c}.\frac{\Delta c(C)}{\Delta t} = \frac{1}{d}.\frac{\Delta c(D)}{\Delta t}$$

It is accepted that reaction rate is a positive number although the concentration of reactants is negative as they decrease over time. The minus sign in front of the reactant concentrations is converted to a positive value for the reaction rate for this reaction.

The rate of a chemical reaction increases with increasing the concentration of reactants in a chemical system. It can be explained by collision theory.

- **Collision theory**

Collision theory is based on the principle that a chemical reaction takes place when two particles (atoms, molecules, ions) collide. The collisions between the particles will be **effective** if they have enough energy to react with each other. Moreover, they must collide with the correct orientation.

Thus, increasing concentration increases the number of particles in the same volume and their collisions will be more frequent.

The rate law for a chemical reaction expresses the relation between the reaction rate and concentrations of the reactants. For a general reaction, $\boldsymbol{a\,A + b\,B \rightarrow c\,C + d\,D}$, the rate law is given by the formula: $\boldsymbol{v = k.c^{a}(A).c^{b}(B)}$, where $\boldsymbol{c^{a}(A)}$.and $\boldsymbol{c^{b}(B)}$ are reactant concentrations (in units of moles per liter), $\boldsymbol{a}$ and $\boldsymbol{b}$ are exponents which vary for each chemical reaction, and $\boldsymbol{k}$ is the rate constant at a particular temperature, specific for a given reaction. The rate constant does not depend on the concentration of the reactants.

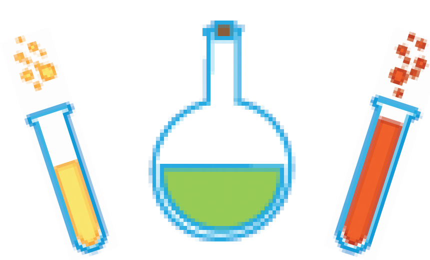
The concentrations of solid substances are not included in the expression for the rate law because they remain constant throughout the reaction.

For example, the rate law for the reaction: $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$, has the form $v = k.c(O_2)$, C is a solid substance.

Test your knowledge:

1. **Write the rate law for the following reactions:**
   a. $Zn_{(s)} + 2HCl_{(aq)} \rightarrow ZnCl_{2(s)} + H_{2(g)}$
   b. $2NO_{(g)} + H_{2(g)} \rightarrow N_{2(g)} + H_2O_{(g)}$
   c. $CaO_{(s)} + CO_{2(g)} \rightarrow CaCO_{3(s)}$
   d. $CO_{2(g)} + NO_{(g)} \rightarrow CO_{(g)} + NO_{2(g)}$
2. **According to collision theory what three criteria must be met for two molecules to react?**
3. **Balance the thermochemical equation below. Express the rate law for the given reaction. On the following set of axes, draw a curve showing the change in concentration of HCl over time.**

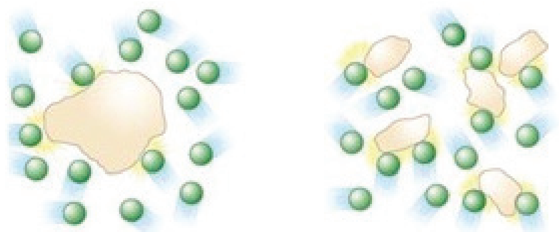
$2Na_{(s)} + 2HCl_{(aq)} \rightarrow 2NaCl_{(aq)} + H_{2(g)}$



6 OTHER FACTORS THAT AFFECT THE REACTION RATE

1. Surface area

For solids only the atoms on the surface of the solid can participate in the chemical reaction with a gas or liquid. If we increase the surface area of a solid reactant, we can increase the rate of the reaction. The surface of a solid can be increased by grinding it to a fine powder. A greater exposed surface area of a solid reactant means a greater chance of effective collisions.



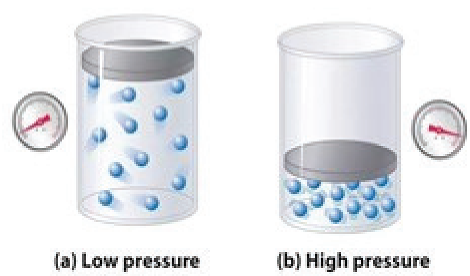
(a) *(b)*

The visualization shows one big lump of a solid (a) several small lumps (b) of the same solid reacting with particles of a gas or liquid. The figure (a) depicts a slow reaction, the figure (b) – a fast reaction.

https://slideplayer.com/slide/9906298/

2. Pressure (when the reactants are gases)

For a reaction that involves gases, increasing the pressure increases the reaction rate. When the pressure increases, the particles of the gases are forced to close together and decrease the amount of empty space between them. The volume of the gases decreases, which increases the collision frequency between reactants. As a result, thenumber of effective collisions also increases, and therefore increases the rate of the reaction.



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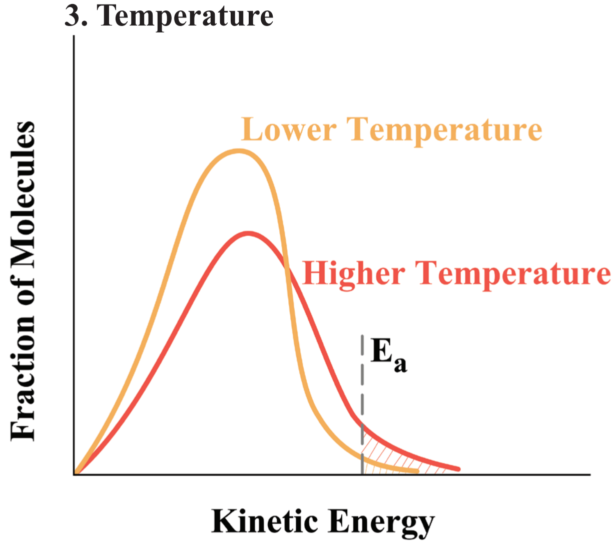
3. Temperature



As a rule of thumb, reactions speed up 2 to 4 times with a $10°C$ rise in temperature. This is true for both exothermic and endothermic reactions. Increasing the temperature of a reaction increases the average speed (kinetic energy) of the molecules, and therefore their collision frequency. Collision frequency cannot be the only factor affecting rate, in the vast majority of collisions, molecules rebound without reacting. Another consequence of the greater kinetic energy is that each collision is more energetic.

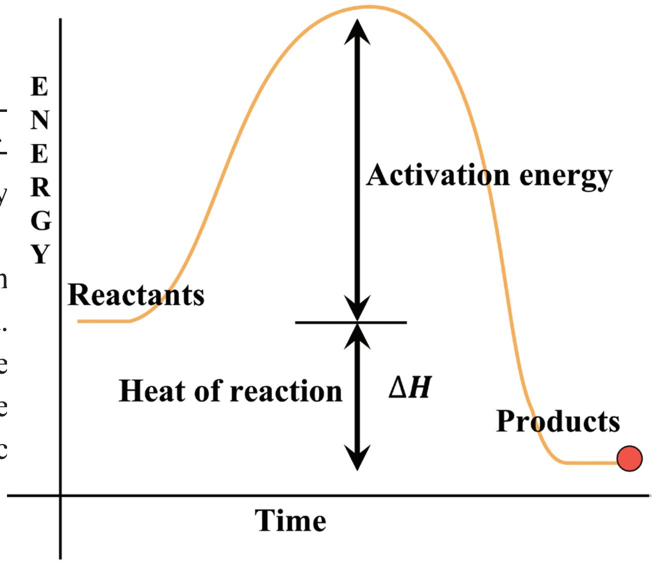
https://courses.lumenlearning.com/suny-introductory-chemistry/chapter/factors-that-affect-the-rate-of-reactions/

4. Activation energy, E_a

The minimum energy that colliding molecules need in order to react (to produce an effective collision) is called the activation energy of the reaction - E_a

If a given collision possesses energy less than E_a, the molecules will bounce apart unchanged.

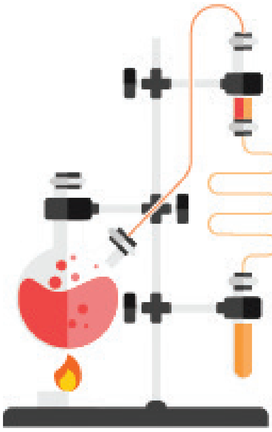
At higher temperatures, the particles have more kinetic energy, so more of them have energy $\geq E_a$. Thus, the number of energetic effective collisions increases.



http://www.dynamicscience.com.au/tester/solutions1/chemistry/rates/catalyst.htm

Test your knowledge:

1. **Write the rate equation for the reaction:**
   $2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$
   a. How does the reaction rate change if the pressure is increased two times?
   b. How does the reaction rate change if the container volume is increased three times?
   c. How does the reaction rate change if the concentration of sulfur dioxide, SO_2 is increased two times and that of oxygen is decreased four times?
2. **Write the rate equation for the reaction:**
   $2C_{(s)} + H_{2(g)} \rightarrow C_2H_{2(g)}$
   a. How does the reaction rate change if the concentration of carbon (C) is decreased two times and that of hydrogen (H_2) is increased two times?
   b. How does the reaction rate change if carbon powder is used instead of lump?



1. Basic concepts

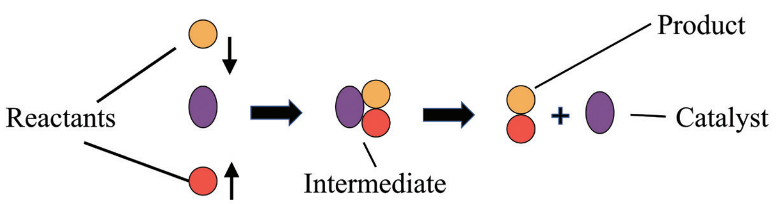
A substance that changes the rate of a chemical reaction without itself being consumed is called a **catalyst**. It only increases the reaction rate and is chemically unchanged at the end of the reaction. The phenomenon is known as **catalysis,** and the reactions with a catalyst as **catalytic.** The opposite of a catalyst is **an inhibitor** – a substance that decreases the rate of a chemical reaction.

2. Types of catalysis

- **Homogeneous Catalysis:** catalysis in which the reactants and the catalyst exist in the same phase (gaseous, liquid, or solid). The homogeneous catalysts are gases, liquids or soluble solids.

The oxidation of sulphur dioxide (SO_2) to sulphur trioxide (SO_3) in the presence of nitric oxide (NO) as a catalyst is an example of a homogeneous (gaseous) catalytic reaction.

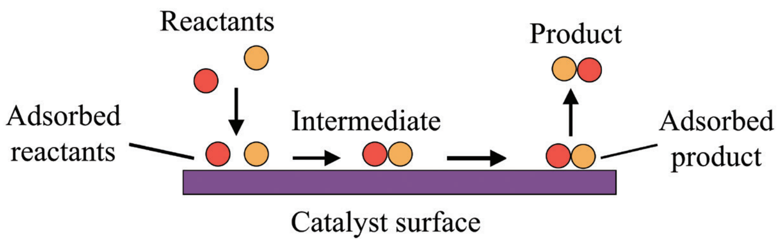
$$2SO_{2(g)} + O_{2(g)} \xrightleftharpoons{NO_{(g)}} 2SO_{3(g)}$$



As shown in the figure, the catalyst takes part in the reaction and affects its rate.

It has been experimentally proven that **the rate of a homogeneous reaction is dependent on the concentration of the catalyst. When we increase its concentration, the rate will also increase.**

- **Heterogeneous Catalysis:** catalysis in which the catalyst and the reactants are in different phases i.e., in different physical states. The catalyst is often a solid and the gaseous or liquid reactants are attracted and adsorbed onto its surface. As a result, the bonds between the atoms of the reactants become weak and break. The atoms rearrange and form new bonds to make the products. The reaction takes place and the new products are desorbed. Therefore, increasing the surface area of a catalyst can speed up the rate of a chemical reaction. **The greater the surface area of a catalyst, the faster the rate of a chemical reaction.**



An example of a heterogeneous catalytic reaction is the oxidation of sulphur dioxide (SO_2) to sulphur trioxide (SO_3) in the presence of a solid vanadium (V) oxide (V_2O_5) as a catalyst.

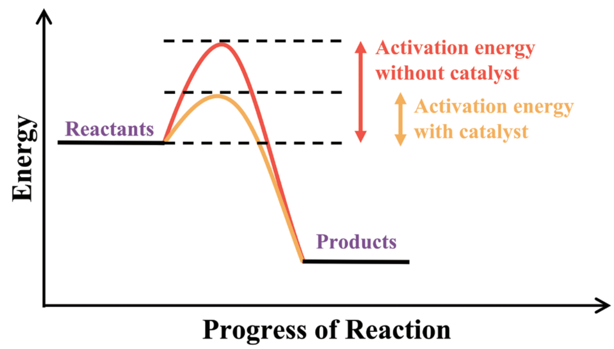
$$2SO_{2(g)} + O_{2(g)} \xrightleftharpoons{V_2O_{5(s)}} 2SO_{3(g)}$$

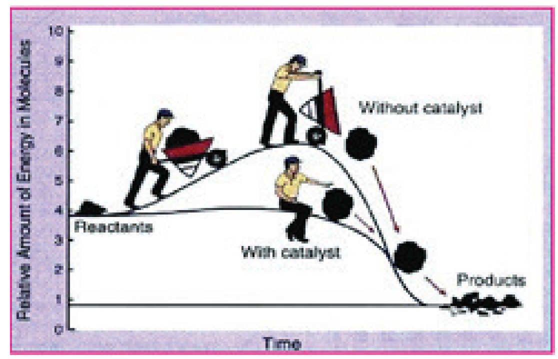
- Biocatalysis (Enzyme Catalysis)

Enzymes are chemical substances that are produced by living cells, they are proteins and act as catalysts in the human body. They regulate the rate of chemical reactions (biological processes) while not being changed themselves.

3. The mechanism of catalysis

Catalysts provide an alternative route for a reaction to occur and decrease the amount of activation energy needed to complete the reaction. Thus, more number of molecules become able to cross the energy barrier per unit time and the reaction proceeds faster and easier than uncatalysed one. The diagrams below represent the energy changes that happen during an exothermic chemical reaction with and without a catalyst. They show the potential energy of the system as a function of time.





https://www.researchgate.net/figure/Diagram-representing-catalyst-lowering-the-activation-energy

Test your knowledge:

1. **What do catalysts do? How do they affect a reaction?**
2. **Classify each of the following reactions as homogeneous or heterogeneous.**

a. $KClO_{3(s)} \xrightarrow{MnO_{2(s)}} 2KCl_{(s)} + 3O_{2(g)}$

b. $4NH_{3(g)} + 5O_{2(g)} \xrightarrow{Pt_{(s)}} 4NO_{(g)} + 6H_2O_{(l)}$

c. $H_2O_{2(l)} \xrightarrow{FeCl_{3(aq)}} 2H_2O_{(l)} + O_{2(g)}$

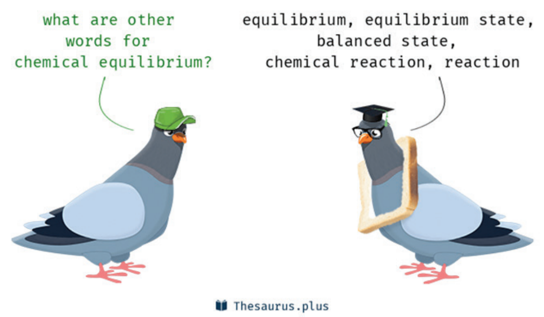
d. $C_{12}H_{22}O_{11(aq)} + H_2O \xrightarrow{H_2SO_{4(aq)}} C_6H_{12}O_{6(aq)} + C_6H_{12}O_{6(aq)}$

3. **Use the terms to correctly fill in the blanks:**

Terms: industry, increase, catalyst, living organisms, rate, alter, inhibitors, decreases, enzymes

The catalysts can (1) ______ the rate of a chemical reaction. They are positive when they (2) ______ the reaction. The iron, for example, is a positive (3) ______ in the production of ammonia. The negative catalysts slow down the (4) ______ of a reaction. Also, they are called (5) ______. For example, the barbituric acid (6) ______ the decomposition of hydrogen peroxide. Apart from (7) ______, catalysts known as (8) ______ play an essential role in the metabolic processes of all (9) ______ including humans, animals, plants and microorganisms.

8 CHEMICAL EQUILIBRIUM



1. Reversible and irreversible chemical reactions:

Reversible chemical reactions can go in both directions, from reactants to products and from products to reactants. They are indicated by arrows pointing in both directions (⇄).

- The forward reaction is one that goes to the right (⟶)
- The reverse (or backward) reaction is one that goes to the left (⟵)

<u>*For example:*</u> The reaction between nitrogen oxide and oxygen is reversible.

$$2NO_{(g)} + O_{2(g)} \rightleftarrows 2NO_{2(g)}$$

$$2NO_{(g)} + O_{2(g)} \xrightarrow{forward} 2NO_{2(g)} \qquad 2NO_{(g)} + O_{2(g)} \xleftarrow{reverse} 2NO_{2(g)}$$

Irreversible chemical reactions can take place only in one direction (⟶), the reactants convert to products, but the products cannot change back to the reactants.

Examples of irreversible reactions are combustion: $CH_4 + 2O_2 \longrightarrow CO_2 + 2H_2O$, neutralization between a strong acid (HCl) and a strong base ($NaOH$): $NaOH + HCl \longrightarrow NaCl + H_2O$ and other reactions that can go to completion.

2. Chemical Equilibrium

Chemical equilibrium is the state of a closed system in which the rate of the forward reaction is equal to the rate of the reverse reaction and the concentrations of reactants and products remain constant. The equation of a chemical reaction at equilibrium is written with a double arrow (⇄) which denotes a reversible reaction. The equilibrium state can only occur in a closed system because no substances can enter or leave it. Thus, there is no loss of products or reactants from the system.

Consider the reversible reaction of ammonia synthesis: $N_{2(g)} + 3H_{2(g)} \rightleftarrows 2NH_{3(g)}$

(A) Forward reaction: $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$ **(B)** Reverse reaction: $2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$

Rate Law: $v_{fwd} = k.c(N_2).c^3(H_2)$ Rate Law: $v_{rev} = k.c^2(NH_3)$

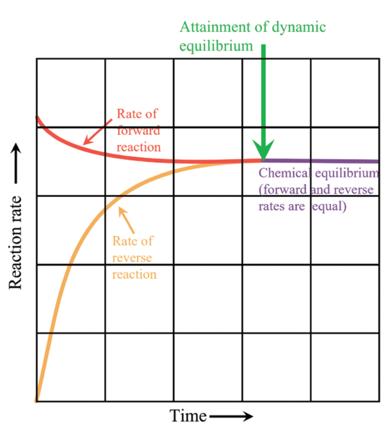


Figure 8.1

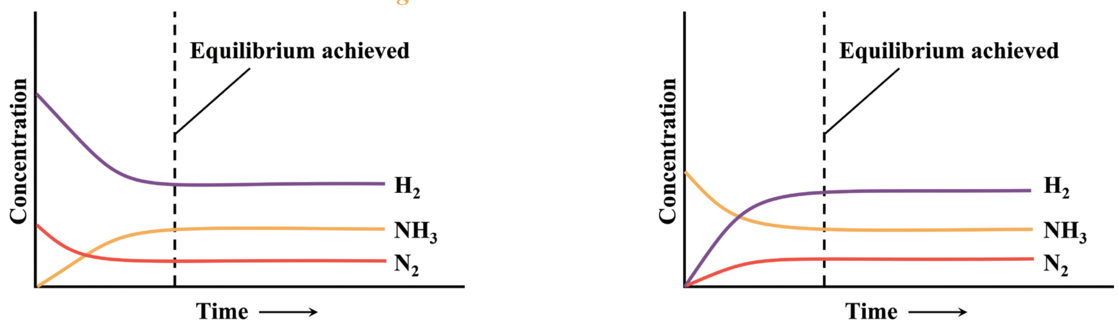
Therefore, at equilibrium $v_{fwd} = v_{rev}$

Figure 8. 1 visually shows the rate of the reaction over time.

Nitrogen reacts with hydrogen under high temperature, high pressure and in the presence of a catalyst. Initially, only the forward reaction takes place because of the great concentrations of N_2 and H_2. Gradually, the starting concentrations begin to decrease while theconcentration of the product - ammonia, NH_3, starts to increase. Over time, v_{fwd} decreases (falling curve), v_{rev} increases (rising curve) until they become equal. This is the point at which the system achieves chemical equilibrium (the curves ofthe two reactions merge into one common – the green line on the graph).

Once equilibrium is achieved the concentrations of the reactants and the product remain constant. Although there is no apparent change, the reaction does not stop. The forward and reverse reactions continue to occur simultaneously, but at identical rates. That's why chemical equilibrium is dynamic. The changes in concentrations of N_2, H_2 and NH_3 for forward (A) and reverse (B) reactions with time are presented in Figure 8.2. The unchanging horizontal lines indicate the equilibrium concentrations of the participants in the system.

Figure 8.2 Concentration vs time



(A) Forward reaction (B) Reverse reaction

3. The equilibrium constant, K_c

K_c is a number that can be derived from the rate laws for the forward and reverse reactions.

If we use the rate law expressions (1) and (2) for the above reaction of ammonia synthesis: $v_{fwd} = v_{rev}$, $k_1 . c(N_2) . c^3(H_2) = k_2 . c^2(NH_3)$, the rearranged equation is: $\frac{k_1}{k_2} = \frac{c^2(NH_3)}{c(N_2).c^3(H_2)}$, $K_c = \frac{k_1}{k_2}$, $K_c = \frac{c^2(NH_3)}{c(N_2).c^3(H_2)}$, where each concentration is raised to its stoichiometric coefficient in the balanced chemical equation. K_c depends on the temperature and the nature of reactants, but pure solids and pure liquids are not included in the equilibrium constant expression. If $K_c > 1$ or greater the equilibrium favours products, if $K_c < 1$ or less the equilibrium favours reactants. K_c represents the concentration in molarity expressed in moles per liter (mol/L).

Test your knowledge:

1. **Write the equilibrium constant expression for each reaction:**

a. $N_{2(g)} + O_{2(g)} \rightleftarrows 2NO_{2(g)}$ c. $4NH_{3(g)} + 5O_{2(g)} \rightleftarrows 4NO_{2(g)} + 6H_2O_{(l)}$

b. $C_{(s)} + CO_{2(g)} \rightleftarrows 2CO_{(g)}$ d. $CaCO_{3(s)} \rightleftarrows CaO_{(g)} + CO_{2(g)}$

2. **Calculate the equilibrium constant for the reaction at a certain temperature:**

$N_{2(g)} + I_{2(g)} \rightleftarrows 2NI_{2(g)}$ if the obtained equilibrium concentrations of the participants are: 0.1mol/L H_2, 0.05 mol/L I_2 and 0.5 mol/L NI. Does the equilibrium favour the products or reactants? (Answer: 50)

3. **The equilibrium constant for the reaction: $PCl_5 \rightleftarrows PCl_3 + Cl_2$ at 250°C is 0.042. The equilibrium concentrations of the reactants are: 0.01mol/L PCl_3 and 0.049 mol/L Cl_2. What is the equilibrium concentration for the product, PCl_5?**

(Answer: 0.012 mol/L)

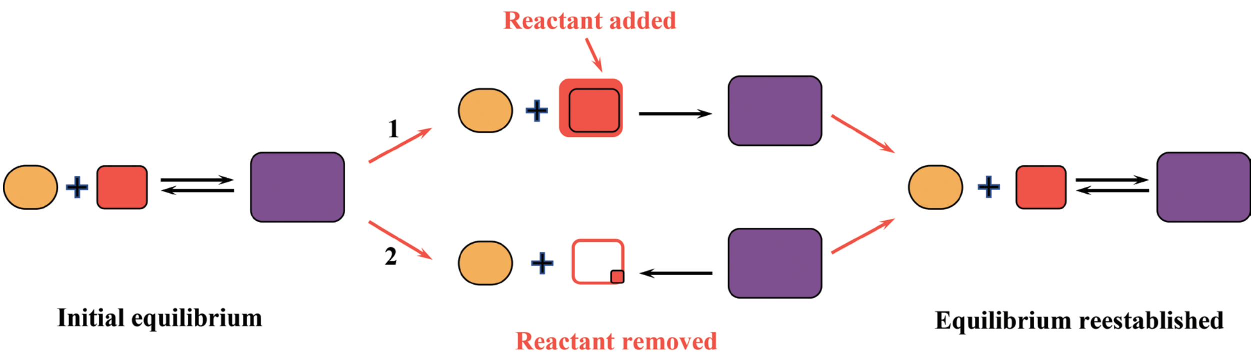
1. Le Chatelier – Braun Principle

"If a system at equilibrium is disturbed by a change in temperature, pressure, or concentration of one of the components, the system will shift its equilibrium position so as to counteract the effect of the disturbance." The principle can be used to select optimum conditions to form a substance.

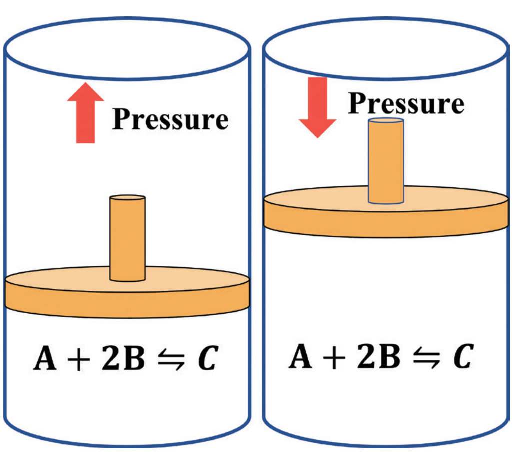
2. Changes in concentration.

If a **reactant** is **added** (1), the equilibrium shifts to the right as **the rate of the forward reaction increases** and reactants form products.

If a **reactant** is **removed** (2), the equilibrium shifts to the left as **the rate of the forward reaction decreases,** causing a shift to the reactants.



If a **product** is added, the equilibrium moves to the left, towards reactants as **the rate of the reverse reaction** increases and vice versa. If a **product** is **removed**, it causes a shift to the right as **the rate of the forward reaction increases.**



3. Changes in pressure – affect gases only

- **Increasing pressure (decreasing volume)** of a system at equilibrium causes a shift in the equilibrium towards a **fewer number of moles of gaseous component.**
- **Decreasing pressure (increasing volume)** of a system at equilibrium moves equilibrium towards **a greater number of moles of gaseous component**.

https://www.youtube.com/hashtag/chemicalequilibrium

Consider the oxidation of sulphur dioxide, SO_2 to sulphur trioxide, SO_3:

$2SO_{2(g)} + O_{2(g)} \rightleftarrows 2SO_{3(g)}$ If the pressure increases (the volume of the reaction vessel decreases), the equilibrium will shift to the right, towards fewer moles

(2 moles of SO_2 + 1 mole of O_2 $\rightleftarrows$ 2 moles of SO_3) in order to undo the stress.

- If a system at equilibrium has **the equal number of moles of gaseous reactants and products**, changing the pressure **does not cause the change** in the equilibrium position.

4. Changes in temperature

- **Increasing the temperature** of a system at equilibrium shifts the **equilibrium in the endothermic direction** because the endothermic reaction absorbs energy and decreases the temperature.
- **Decreasing the temperature** of a system at equilibrium shifts the equilibrium **in the exothermic direction** because the exothermic reaction releases heat(energy) and removes the stress as increases the temperature.

Consider the following reaction:

Exothermic reaction

$$N_2 + 3H_2 \rightleftarrows 2NH_3 \qquad \Delta H^{\ominus} = -91.8\ kJ/mol$$

Endothermic reaction

The reaction is reversible, and the production of ammonia is an exothermic process. If we add heat, the equilibrium will shift to the left (towards the reactants – nitrogen and hydrogen), in the endothermic direction that consumes energy. If we cool the system, the temperature will lower, and the equilibrium will move to the right (towards the product − ammonia) in an exothermic direction to compensate the stress.

5. Catalyst

Adding a catalyst **has no effect on a system at equilibrium**. It lowers the activation energy and speeds up the rates of both forward and reverse reactions and facilitates equilibrium to be established more quickly.

Test your knowledge:

1. Define the term dynamic equilibrium. What are the factors that affect the position of equilibrium?

2. During the synthesis of ammonia which reaction is favoured when nitrogen is added to the system?

a. the reverse reaction b. the forward reaction c. the reaction that adds N_2 to the system

3. A catalytic converter is a device that reduces pollutants a motor vehicle releases. It converts the poisonous carbon monoxide (in the exhaust steam) to carbon dioxide by oxidation. Use the equation of the reaction to complete the table below: $2CO_{(g)} + O_{2(g)} \rightleftarrows 2CO_{2(g)}$, $\Delta H^{\ominus} < 0$

An increase in:	causes a shift in the equilibrium towards the ___ reaction.	Does the value of KC change?
concentration of O_2		
pressure		
temperature		

4. Methanol (or methyl alcohol) is produced commercially by the exothermic reaction: $CO_{(g)} + 2H_{2(g)} \rightleftarrows CH_3OH_{(g)}$, in the presence of copper-zinc catalysts, $\Delta H^{\ominus} = -91.5\ kJ/mol$.

Write the equilibrium constant expression for this reaction.

How does the equilibrium position shift if?

a. H_2 is removed from the system.
b. Pressure is increased.
c. Temperature is decreased.
d. A catalyst is added.

10 SELF-CHECK 1

1. Indicate whether the following statements are true or false:

a. An isolated system can exchange energy with its surroundings.

b. The activation energy in a chemical reaction is the amount of energy needed for a reaction to occur.

c. In an exothermic process, the surroundings lose energy.

d. In an endothermic process, the system loses energy.

e. According to the collision theory, molecules must collide with each other with enough energy and correct orientation in order to react.

f. Increasing the concentration of the reactants decreases the rate of a chemical reaction.

2. Enthalpy changes can be calculated using enthalpy changes of formation. The table below shows some values for standard enthalpy changes of formation. Use these values to calculate the standard enthalpy change of the reaction: $CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$

Substance	$\Delta H_f^{\ominus} = kJ/mol$
$H_2O_{(g)}$	−241.8
$CO_{2(g)}$	−393.51
$CH_{4(g)}$	+74.81

$$\Delta H^{\ominus} = (\Sigma_n . \Delta H_f^{\ominus})_{products} - (\Sigma_n . \Delta H_f^{\ominus})_{reactants}$$

3. Consider the following reaction: $Zn_{(s)} + H_2SO_{4(aq)} \rightarrow ZnSO_{4(aq)} + H_{2(g)}$

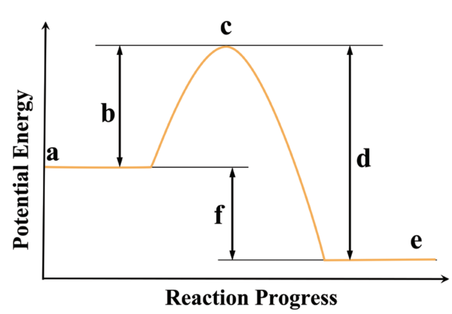
Write the equation rate for the rection.

How does the reaction rate change if the concentration of H_2SO_4 is increased?

How does the reaction rate change if the concentration of $ZnSO_4$ is removed from the vessel?

4. Identify the following terms with their correct letter from the potential energy diagram below

Letter				d		
Term	Activation energy	Energy of products	Enthalpy, ΔH	Energy of the activated complex	Activated complex	Energy of reactants



5. Lime, CaO is produced by heating calcium carbonate, $CaCO_3$:

$$CaCO_{3(s)} \rightleftarrows CaO_{(s)} + CO_{2(g)} \qquad \Delta H^{\ominus} = 178.1\ kJ/mol$$

a. Write the equilibrium constant expression for the reaction.

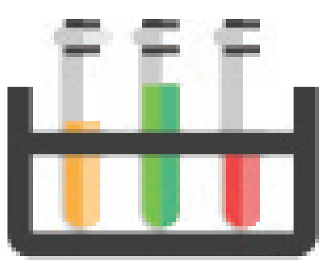
b. Predict the shift in the equilibrium system resulting in each of the following changes:

- an increase in the concentration of CO_2
- a decrease in the concentration of $CaCO_3$
- an increase in temperature
- the addition of a catalyst

6. Write the correct vocabulary term in the blanks. Use the circled letter in each term to find the hidden vocabulary word.

Clues	Vocabulary Terms
1. Another word for speed	___ ___ ___ ___
2. A reaction that absorbs energy from its surroundings	___ ___ ___ ___ ___ ___ ___ ___ ___ ___ ___
3. A substance that changes the rate of a chemical reaction	___ ___ ___ ___ ___ ___ ___
4. A reaction that releases energy to the surroundings	___ ___ ___ ___ ___ ___ ___ ___ ___
5. The substances that undergo changes in a chemical reaction	___ ___ ___ ___ ___ ___ ___ ___
6. Another word for change	___ ___ ___ ___
7. A system that can exchange both matter and energy with the surroundings	___ ___ ___
8. The power that is used to provide heat	___ ___ ___ ___ ___

Hidden term: ___ ___ ___ ___ ___ ___ ___ ___



SOLUTIONS AND REACTIONS IN AQUEOUS SOLUTIONS

SOLUTIONS

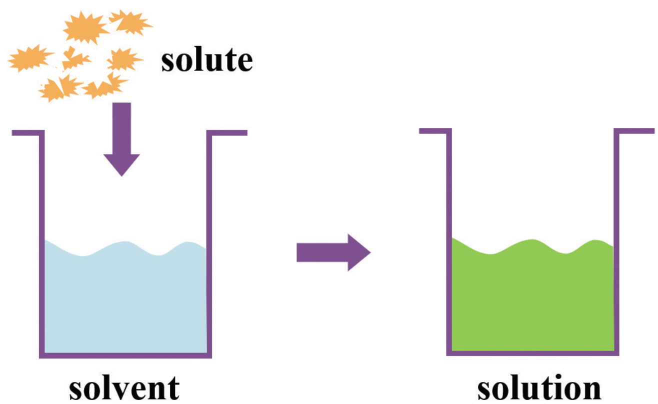
11 TYPES OF SOLUTIONS

1. Homogeneous and heterogeneous mixtures

- **Mixture** is a combination of two or more substances which are not chemically joined.
- **Homogeneous mixture** is a mixture which has a uniform (the same) composition throughout its volume. It consists of a single phase. There is no visible line of separation between its components. For example: tea, coffee, milk, vinegar, salt water, air, brass, etc.
- **Heterogeneous mixture** is a mixture which has a non-uniform (not the same) composition throughout its volume. It consists of two or more phases. Its components can be easily seen as they remain physically separate. For example: soil, rocks, sand in water, mud, oil in water, salad, pizza, etc.

2. Solutions.

Solutions are homogeneous mixtures that consist of two or more substances. Each solution has a **solute** and a **solvent** as its components. A solute is a substance dissolved in another substance known as a solvent. The solvent is present in a larger quantity than the solute.



3. Types of solutions.

Solutions can be in the form of solids, liquids, or gases. They can be classified on the bases of the physical state of their solvents and solutes.

Type 1. Gas solutions:

- **Gas in a gas**: the solute is a gas, the solvent is a gas, e.g., air (solute – oxygen, solvent – nitrogen)

Type 2. Liquid solutions:

- **Gas in a liquid**: the solute is a gas, the solvent is a liquid, e.g., sparkling water (solute – carbon dioxide, solvent – water)
- **Liquid in a liquid**: the solute is a liquid, the solvent is a liquid, e.g., antifreeze (solute – ethylene glycol, solvent – water)
- **Solid in a liquid**: the solute is a solid, the solvent is a liquid, e.g., sugar solution (solute – sugar, solvent – water)

Type 3. Solid solutions:

- **Gas in a solid**: the solute is a gas, the solvent is a solid, e.g., dissolved gases in rocks (solute – gases as CO_2, H_2, N_2, solvent – granite)
- **Liquid in a solid:** the solute is a liquid, the solvent is solid, e.g., amalgam of mercury with sodium (solute – sodium, solvent – mercury)
- **Solid in a solid:** the solute is a solid, the solvent is a solid, e.g., steel (solute – carbon, solvent – iron)

4. Characteristics of solutions:

- A solution is a homogeneous mixture with uniform particle distribution.
- A solution is composed of one phase (gaseous, liquid, or solid) with no visible boundary separating its components.
- A solution is very stable. The solute particles do not settle out if a solution is kept undisturbed indefinitely.
- A solution cannot be separated into its components by filtration because of very small size of its particles, less than $1\ nm(10^{-9}\ m)$ in diameter.
- A solution can be prepared by various solute/solvent combinations, each one with a different composition.

Test your knowledge:

1. **Define the term solution. Create your own example of a solution.**
2. **Why is nitrogen the solvent in the air? Indicate the solute(s) in the solution.**
3. **Classify the following mixtures as either homogeneous or heterogeneous by putting a thick in the box.**

Mixture	Homogeneous	Heterogeneous
Fruit juice		
Honey		
Pebbles in concrete		
Salad		
Milk shake		
Ice in soda		
Wine		

4. **Match the solution type with its corresponding example.**

a. Gas – gas solution	1. Vinegar
b. Liquid – solid solution	2. Coca cola
c. Solid – solid solution	3. Seawater
d. Gas – liquid solution	4. Bronze
e. Liquid – liquid solution	5. Heliox (helium-oxygen mixture)
f. Solid – liquid solution	6. Dental amalgam

5. **Indicate the solute(s) and the solvent in each of the following solutions:**

a. 20 g of $NaCl$ in 200 g of water: **solute ___________, solvent ___________.**

b. Laughing gas, which is 70% nitrous oxide, N_2O and 30% oxygen maximum.
solute _______________, solvent ______________.

c. Stainless steel contains 77% iron, 18% chromium and 5% nickel.
solute _______________, solvent ______________.

d. Household bleach is a water solution of $3-6\%$ sodium hypochlorite, $NaClO$.
solute _______________, solvent ______________.

12 CONCENTRATION OF SOLUTIONS

1. Solution concentration.

The concentration of a solution is a measure of the amount of solute dissolved in the given amount of solvent or solution. It is important to know the concentration of solutes to control the stoichiometry of reactants for solution reactions. Concentration can be expressed in different quantitative methods.

2. Molarity (c)

Molarity is also known as amount concentration or molar concentration. It is the amount of solute or the total number of moles of solute per liter of solution. Molarity is denoted by lowercase $\boldsymbol{c}$. The unit of molarity has a dimension mole per liter of solution (mol/L). To calculate the molarity of a solution we use the following formula:

$$c = \frac{amount\ of\ solute}{volume\ of\ solution}\ \frac{[mol]}{[L]},\ c = \frac{n}{V}\ \frac{[mol]}{[L]},$$

where $\boldsymbol{n}$ is the amount of the solute in moles, $\boldsymbol{V}$ is the volume of the solution.

The unit of molar concentration is symbolized by the capital letter $\boldsymbol{M}$, especially when the molarity is reported. For example, a solution labeled as $2M\ HNO_3$ is read as "2 molar nitric acid solution".

Solved example of molarity

Calculate the molarity of the following aqueous solution: $25\ g$ of table salt $(NaCl)$ in $150\ ml$ of solution.

Given: $m(NaCl) = 25g$ **To be found**: $c(NaCl) = ?$

$V = 150ml$

Solution:

✓ The molar mass of $NaCl$ is calculated by adding the relative atomic masses of sodium and chloride in the molecular formula and then multiplying by the molar mass constant $(1g/mol)$.
$M(NaCl) = (A_r\ (Na) + A_r\ (Cl)).1g/mol = (23 + 35.45).1\ g/mol = 58.45\ g/mol$

$$n(NaCl) = \frac{m(NaCl)}{M(NaCl)} = \frac{25g}{58.45g/mol} = 0.43\ mol$$

✓ The number of moles of $NaCl$ is calculated by the equation:

✓ The volume of the solution is $150\ ml$, in terms of liters: $V = \frac{150}{1000} = 0.15\ L$

✓ The molarity of the given solution is obtained by dividing the amount of salt by the volume:

$$c(NaCl) = \frac{n(NaCl)}{V(solution)} = \frac{0.43}{0.15} = 2.87\ mol/L$$

✓ The molarity of $25\ g$ of $NaCl$ in $150\ ml$ of water is $2.87 mol/L$ or $2.87M$.

3. Mass fraction (ω).

The amount of a solute present in a solution can be expressed as a percent by mass. It is denoted by the lowercase Greek letter omega, $\boldsymbol{\omega}$. The mass fraction is defined as:

$$\omega\% = \frac{mass\ of\ solute}{mass\ of\ solution}.100, mass\ of\ solution = mass\ of\ solute + mass\ of\ solvent$$

12 CONCENTRATION OF SOLUTIONS

$$\omega\%(A) = \frac{m(A)}{m}.100,$$

where $\omega\%(A)$ is the mass percent of a solute A dissolved in a solution with a given mass m

Solved example of mass fraction

Calculate the mass fraction of solute in the following solution: $45\ g$ of glucose ($C_6H_{12}O_6$) dissolved in $300\ g$ of water.

Given: $m(C_6H_{12}O_6) = 45g$ **To be found:** $\omega(C_6H_{12}O_6) = ?$

$m(H_2O) = 300g$

Solution:

✓ The mass of the solution:

$$m(solution) = m(C_6H_{12}O_6) + m(H_2O) = 45g + 300g = 345g$$

✓ The mass fraction of glucose is calculated by the equation:

$$\omega\%\,(C_6H_{12}O_6) = \frac{(m(C_6H_{12}O_6))}{(m(solution))} = \frac{45}{345}.100 = 13\%$$

Test your knowledge:

1. **Calculate the molarity of the following aqueous solutions:**
   a. 50 g of sodium hydroxide ($NaOH$) in 200 ml of solution
   b. 240 g of magnesium sulphate ($MgSO_4$) in 1000 ml of solution

(Answer: a) 6.25 M; b) 2 M)

2. **How many grams of calcium carbonate are contained in $1L$ of a $3M$ $CaCO_3$ solution?**

(Answer: 300 g)

3. **Potassium permanganate ($KMnO_4$) is used for several skin conditions because of its antiseptic properties. How would you prepare a solution of 0.250 M $KMnO_4$ in 3.0 L of water?**

(Answer: 118.5 g)

4. **Calculate the mass fraction of solute in the following solutions:**
   a. 7.5 g of potassium nitrate (KNO_3) in 95 g of water
   b. 46 g of ammonium chloride (NH_4Cl) in 500 g of water

(Answer: a) 7.3%; b) 8.4%)

5. **Which of the following answers represent the mass fraction of sodium sulphate (Na_2SO_4) in a solution made by dissolving 35 g of Na_2SO_4 in 300 g of water?**
   a. 11.7% b. 11.6% c. 10.4% d. 10.5%

13 THE SOLUTION PROCESS. AQUEOUS SOLUTIONS

1. The solution process (Dissolution)

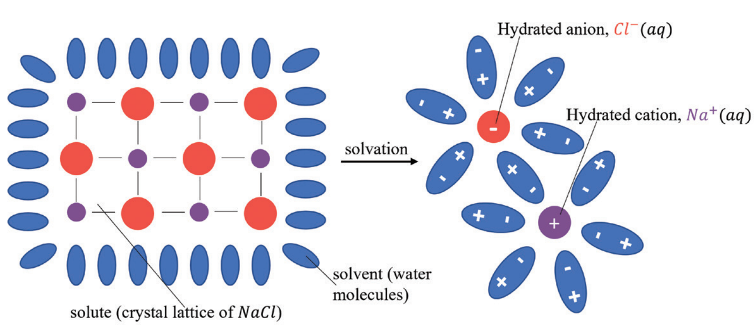
A solution is formed when one substance dissolves completely in another substance. The process is known as **dissolution** and involves the following three steps:

➢ **Separation of the solute particles**. The solute particles separate by overcoming electrostatic attractions between each other. This requires energy, so it is endothermic, $\Delta H_1^{\ominus} > 0$.

➢ **Separation of solvent particles**. The solvent molecules also separate by overcoming the intermolecular attractions which hold them together. They spread apart to provide space for solute particles. The process requires energy, so it is endothermic, $\Delta H_2^{\ominus} > 0$.

➢ **Interaction of solute-solvent particles.** The solvent particles surround the solute particles and form **solvates**. **Hydrates** are formed when the solvent is water. The process is called **solvation (hydration).** The solute-solvent interaction releases heat because of forming new bonds, so the process is exothermic, $\Delta H_3^{\ominus} < 0$.

The overall enthalpy change of dissolution process is a sum of enthalpy changes of these three steps. $\Delta H^{\ominus}(solution) = \Delta H_1^{\ominus} + \Delta H_2^{\ominus} + \Delta H_3^{\ominus}$. $\Delta H^{\ominus}(solution)$ can be either positive or negative depending on the amount of energy absorbed or released in each step. If $\Delta H^{\ominus}(solution) > 0$, an endothermic reaction takes place (the solution feels cold). If $\Delta H^{\ominus}(solution) < 0$, an exothermic process occurs, the solution feels hot because the heat is given off. The values of $\Delta H^{\ominus}(solution)$ of different solutes in inorganic and organic solvents can be found in different tables of data. For example: $\Delta H^{\ominus}solution\,(MgNO_3) = -91.2\,\frac{kJ}{mol}$, $\Delta H^{\ominus}solution\,(NH_4NO_3) = 26.4\,\frac{kJ}{mol}$,



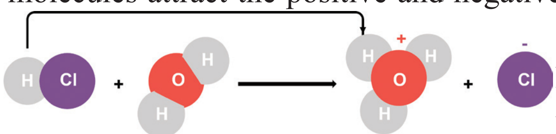
2. Ionic compounds as solutes

Ionic compounds consist of oppositely charged ions that are held together by electrostatic forces. Their ions are arranged in a crystal lattice. A common example of an ionic compound is table salt or sodium chloride ($NaCl$). When a small amount of sodium chloride is placed in water, it easily dissolved in it. The water molecules, which moved randomly in the container, collide with the crystal lattice of the salt as shown in the figure below. The positively charged sodium ions (**Na^+, the purple balls in the figure**) attract the negative end of water molecules (oxygen end) because they are partially negative. Thenegatively charged chloride ions (**Cl^-, the red balls in the figure**) attract the positive end of the water molecules (hydrogen end) as they are positive. Gradually, the crystal lattice is separated into ions. Sodium and chloride ions are surrounded by the solvent (water) molecules and form solvates or hydrates in this case. The equation of the dissolution $NaCl_{(s)} + H_2O_{(l)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)}$,

aq = aqueous = dissolved in water

3. Covalent compounds as solutes

Solutes that have highly polar covalent bonds can dissolve in water. The polar covalent bonds between the atoms in a molecule have unequal sharing of electrons causing partial charges to form on each of the bonding atoms. When a solute is being dissolved in water, the polar water molecules attract the positive and negative areas of the solute molecules, separate them to form



solvates (hydrates). For example, when gaseous hydrogen chloride (HCl) dissolves in water, it ionises to yield hydronium (H_3O^+) and chloride (Cl^-) ions in the solution. The hydrochloric acid is being formed. The hydronium ion is the hydrogen ion (of HCl) bonded to the water molecule in the aqueous solution. The equation of the process is:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

Most solutes that contain covalent bonds dissolve in water, but they do not react with it. When they dissolve, they break apart into molecules, not into ions. Their molecules stay together in solutions. Water molecules surround each entire covalent molecule and solve it. For example, when sugar dissolves in water, it does not separate into ions: $C_{12}H_{22}O_{11(s)} \rightarrow C_{12}H_{22}O_{11(l)}$

4. "Like dissolves like" – a general rule in chemistry (rule of thumb) which means that:

➢ Polar substances dissolve in polar substances (they have similar types of intermolecular forces). Polar solvents (water, ethanol) dissolve polar and ionic solutes.

➢ Nonpolar substances dissolve in nonpolar substances. The weak intermolecular attraction of a nonpolar solute can only be overcome by the weak intermolecular forces of a nonpolar solvent. Nonpolar solvents (benzene, hexane, chloroform, diethyl ether) can dissolve nonpolar solutes.

➢ For example, water can dissolve salts, methanol, but cannot dissolve oil.

➢ Some substances (ethanol, soap) can dissolve in both polar and nonpolar solvents because their molecules have polar and nonpolar regions.

Test your knowledge:

1. Which solvent, water or hexane, C_6H_{14} would you choose to dissolve each of the following? Put a thick in a column if a given solute dissolves in any of these two solvents.

solute / solvent	HF	Octane (C_8H_{18})	$(NH_4)_2SO_4$	Petrol	Iodine (I_2)	Methanol CH_3OH	H_2SO_4	Benzene (C_6H_6)
water								
hexane								

2. Solid potassium chloride, KCl is added to water. It dissolves because:

A) The K^+ ions are attracted to the

a. oxygen atom (δ^-) of water

b. hydrogen atom (δ^+) of water.

B) The Cl^- ions are attracted to the

a. oxygen atom (δ^-) of water

b. hydrogen atom (δ^+) of water.

3. Disposable instant cold packs with ammonium chloride, NH_4Cl are used as a first aid solution to minor injuries such as sprains, strains, broken bones, etc. NH_4Cl is stored in a sealed plastic bag surrounded by water. When you open the inner bag, the ammonium chloride comes into quickly? Describe the process that takes place.

14 SOLUBILITY

1. Solubility is the maximum amount of one substance (solute) that can be dissolved in a fixed amount of another substance (solvent) at a specified temperature. On the base of the dissolved amount of a solute, solutions can be classified as saturated, unsaturated, and supersaturated.

➢ **A saturated solution** contains the maximum amount of a solute that can be dissolved at a certain temperature. This type of solution is in equilibrium with excess undissolved solute. At equilibrium, the dissolved solute particles move throughout the solution with a rate equal to the rate at which some of them collide with the surface of particles of the undissolved solute, stick to this surface and crystallise. Two opposing processes occur simultaneously at the same rates – dissolution and crystallisation. The equilibrium is possible on the condition that excess solid is present. The achieved balance can be presented as follows: $solute \Leftarrow \frac{dissolution}{crystallisation} \Rightarrow solution$

➢ An unsaturated solution contains less than the maximum amount of solute that can be dissolved at a specified temperature.

➢ A supersaturated solution contains more solute than this for the saturated solution. It can be formed from a saturated solution by cooling it and filtering off the excess solute. This type of solution is unstable and can recrystallise if a "seed" crystal is added to it.

2. Factors that affect solubility:

2.1 Temperature effect (in aqueous solutions).

➢ For most of the solids, such as sodium nitrate ($NaNO_3$), potassium chloride ($CaCl_2$), etc., shown in *Figure 14.1.* below, the solubility increases with the increase of temperature. Since the average kinetic energy of the molecules that make up the solution increases with increasing temperature, the solvent molecules more easily separate the solute molecules, and the solute molecules dissolves more readily. Therefore, the dissolution is endothermic, $\Delta H^{\ominus} > 0$.

➢ For a few solids, solubility decreases with increasing temperature. Hence, the dissolution is exothermic, $\Delta H^{\ominus} < 0$. For example, the solubility of cerium (III) sulphate ($Ce_2(SO_4)_3$) that is illustrated in the same *Figure 14.1*, decreases with a rise in temperature.

➢ For all gases, the solubility decreases with increasing temperature. Gases in water become less soluble with rising temperature because gasses expand and escape from their solvents. The dissolution is exothermic, $\Delta H^{\ominus} < 0$.

2.2. Pressure effect (in aqueous solutions)

Pressure greatly affects the solubility of gases. Pressure increases the solubility of gases. The relationship between gas pressure and solubility is represented by Henry's law states that at constant temperature the amount of dissolved gas in a definite volume of liquid is proportional to the gas pressure above that liquid. It is expressed by the equation: $\boldsymbol{c_g = k.p_g}$, where $\boldsymbol{c_g}$ is a concentration of a dissolved gas in a particular solvent at a fixed temperature, mol/L, $\boldsymbol{k}$ is Henry's constant, specific for each gas-solvent combination at a given temperature, often in units $mol/L.atm$, $\boldsymbol{p_g}$ is the partial pressure of a gas, atm. Henry's Law is applicable only when there is not a chemical reaction between the solute and solvent. For example, the dissolution of ammonia gas in water does not obey Henry's Law because of a chemical reaction between ammonia and water molecules.

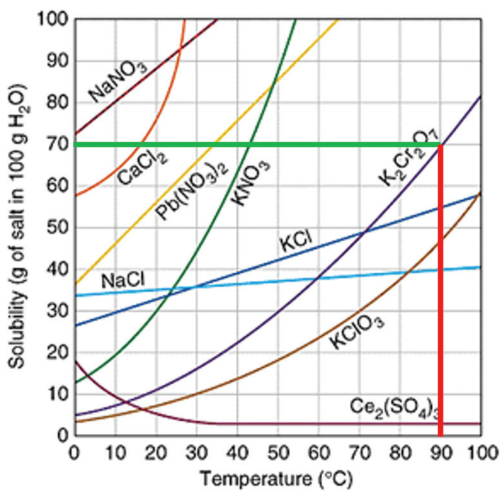


Figure 14.1

3. Reading a solubility curve.

The relationship between temperature and solubility is expressed in a graph called a solubility curve (Figure 14.1). Each curve shows the amount of a solute that can be dissolved in $100g$ of water at different temperatures. Any amount of solute below the line indicates that the solution is unsaturated, above the line shows the supersaturated solution. Any amount of solute on the line shows that the solution is saturated. For example, 70 $grams$ of $K_2Cr_2O_7$(on the y-axis) dissolves at 90°C (on the x-axis) form a saturated solution represented as the intersection point of the two lines (green and red) is on the curve. 80 $grams$ of $K_2Cr_2O_7$ at same temperature make up a supersaturated solution (the point is above the curve) whereas 50 $grams$ of $K_2Cr_2O_7$ at 90°C compose an unsaturated solution (the point is below the curve).

http://www.dynamicscience.com.au/tester/solutions1/chemistry/solutions/solubilitycurves.html

Test your knowledge:

For problems 1 – 5 use the graph above (*Figure 14.1.*).

1. State whether the following solutions are unsaturated, saturated or supersaturated.

a) $50g$ of KCl at $70°C$ d) $20g$ of $KClO_3$ at $40°C$

b) $90g$ of $NaNO_3$ at $20°C$ e) $30g$ of $NaCl$ at $60°C$

c) $10g$ of $Ce_2(SO_4)_3$ at $10°C$ f) $40g$ of $Pb(NO_3)_2$ at $30°C$

2. How many grams of the following solutes must be added to $100g$ of water to form a saturated solution (on the line) at that given temperature?

a) $KClO_3$ at $50°C$ b) $CaCl_2$ at $20°C$ c) $NaCl$ at $90°C$

3. Determine which of the solutions is more concentrated (more dissolved solute) at $40°C$: a saturated solution of $NaCl$ or a saturated solution of KCl?

4. A mass of $30\ g$ of $KClO_3$ is dissolved in $100\ g$ of water at $70°C$:

a) What solution is prepared?

b) The solution is cooled to $30°C$. How many grams of crystals are formed?

(Answer: a) saturated; b)10 g)

5. What amount of KNO_3 will dissolve in $50\ g$ of water at $40°C$? *(Answer: $31\ g$)*

(Hint: Use the proportion: given solubility $\frac{amount\ of\ solute}{amount\ of\ solvent} = \frac{amount\ of\ solute}{amount\ of\ solvent}$ (unknown solubility)

6.The partial pressure of CO_2 in a bottle of soda is $3.5\ atm$ at $25°C$. Henry's constant for CO_2 dissolved in water is $3.4.10-2\,\frac{mol}{L}.atm$ at $25°C$. Find the concentration of CO_2

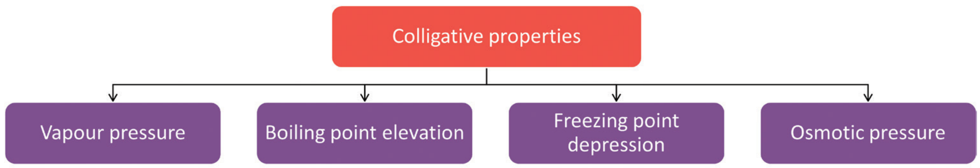
(Answer: $0.1\ mol/L$)

7. Which of the following dissolved gases in water will not obey Henry's Law: $H_2, H_2S, HCl, O_2, CO, SO_3, CO_2$?

15 COLLIGATIVE PROPERTIES OF SOLUTIONS

The word **colligative** originates from the Latin word ***colligatus*** meaning bound together.

In chemistry, **colligative properties** are physical properties of solutions that depend only on the number of solute particles, not on their identities.



Only properties of non-volatile and nonelectrolyte solutions are considered. By definition, a nonvolatile nonelectrolyte is a substance does not evaporate (does not change easily into a gas) nor separates into ions when dissolved. For example, glucose, sucrose, glycerin, ethylene glycol, etc.

- Vapour pressure

Vapour pressure is the pressure exerted by a vapour (gas) of a liquid or solid over its liquid surface, at a particular temperature. This is the pressure at which the liquid is converted to vapours. When a pure solvent as water or any volatile substance (evaporates easily) is in a closed container at a certain temperature, its molecules evaporate or condense at the same rate in a dynamic equilibrium: ($liquid \rightleftarrows gas$). This is because the molecules with high energy tend to escape from the liquid to the gaseous phase. The process is known as **evaporation** or **vaporisation**. At the same time, the molecules that are in vapours collide with the walls of the container or with the surface of the liquid and reenter the solution. These molecules condense and the process is **condensation**.

At equilibrium: $v_{evaporation} = v_{condensation}$

If a nonvolatile solute is added, the nonvolatile particles reduce the number of the solvent particles on the surface as they are solvated (hydrated). As a result, less solvent particles can evaporate from the surface which lowers the vapour pressure of the liquid. The rate of evaporation decreases although the condensation continues at the same rate. Since the rate of condensation is higher than the rate of vaporization, the vapour pressure is reduced. A new dynamic equilibrium is established at a lower pressure. The model illustrated in *Figure 15.1.* below is an example depicting why the vapour pressure of a solution of a nonvolatile substance as sucrose is less than the vapour pressure of pure water.

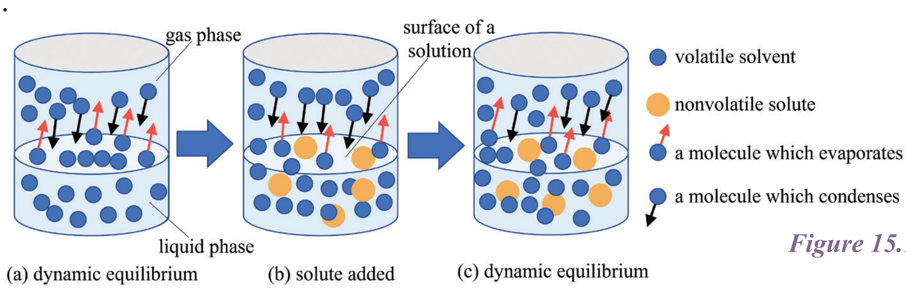


Figure 15.1

Raoult's Law. It states that the presence of a nonvolatile substance lowers the vapour pressure of a solvent. The extent to which a nonvolatile solute lowers the vapour pressure is proportional to itsconcentration. The vapour pressure of the solvent above the solution (p) is lower than the vapour pressure above the pure solvent (p_0), $p < p_0$.

- **Boiling point elevation** The temperature at which the vapour pressure of a liquid is the same as the pressure surrounding the liquid is called **boiling point**. It is the temperature at which the liquid boils. If the liquid is in a closed system, its pressure equals the internal pressure over the liquid. If the liquid is in an open system, its pressure equals atmospheric pressure. The units of atmospheric pressure can be $760\ mmHg$ (torr) or $101.3\ kPa$ or $1\ atm$. Since $p < p_o$, the boiling point of a solution increases or elevates with an increase in the number of solute particles. The elevation in boiling point is proportional to the concentration of solute: $\Delta T_b \propto c$. ***Increasing the concentration of a solute increases the boiling point of a solution. Hence, a solution with greater number of solute particles boils at a higher temperature than less concentrated solution.***

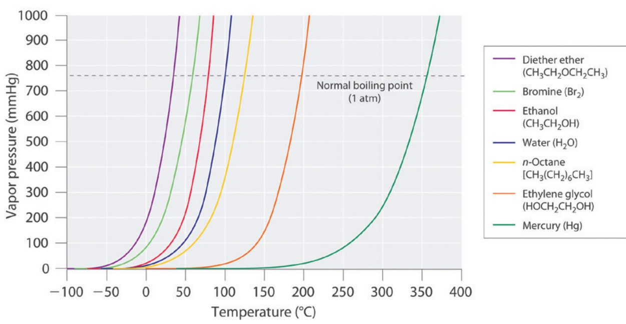


Figure 15.2 The vapour pressure of some liquids as a function of temperature

The temperature dependence of the vapour pressure of several liquids is depicted in *Figure 15.2* to the left. As the temperature rises, the volatile pressure of each liquid also increases until it reaches the point at which it boils. It is the intersection point of its curve and the dashed line. The dashed line is used to represent the atmospheric pressure ($1\ atm$)

https://chem.libretexts.org/Courses/Mount_Royal_University/Chem_1201/Unit_5%3A_Intermolecular_Forces/5.5%3A_Vapor_Pressure

- **Freezing point depression**

The temperature at which a liquid crystallises (freezes) to a solid is called freezing point. This is the point at which the vapour pressure of a liquid equals the vapour pressure of the solid phase, so both phases are in equilibrium. It is well known that water freezes at $0°C$. If a nonvolatile substance is added to water, the freezing point of the solution will be lower than pure water because of a lower vapour pressure. The freezing point of a solution is depressed due to the decreasing of its vapour pressure. The freezing point depression ΔT_f is proportional to the concentration of solute: $\Delta T_f \propto c$. ***Increasing the concentration of a solute decreases the freezing point of a solution. Hence, a solution with greater number of solute particles freezes at a lower temperature than less concentrated solution.***

Test your knowledge:

1. Why is the vapour pressure of an aqueous solution of glucose lower than that of water?
2. Ethylene glycol (antifreeze) is used as to keep the water in vehicles' engines from freezing. Why?
3. At sea level, atmospheric pressure is 1 atm or 760 $mmHg$ but it decreases with altitude. On the top of Mount Everest, it drops to over 228 $mmHg$ or 0.30 atm. Water boils at 100°C at sea level and at 68°C on Everest. Explain the difference in temperatures in terms of vapour pressure. Why is it difficult for climbers to prepare a decent cup of tea or coffee there?
4. Vapour pressure curves of diethyl ether, bromine, ethanol, and ethylene glycol are given on the graph above (*Figure 15.2*). Using the graph, arrange the liquids in order of increasing vapour pressure and decreasing boiling points.

- **Diffusion**

Diffusion is the process of movement of particles from an area of high concentration to an area of low concentration. The word "diffusion" is derived from the Latin word *"diffundere"* meaning "to spread out".

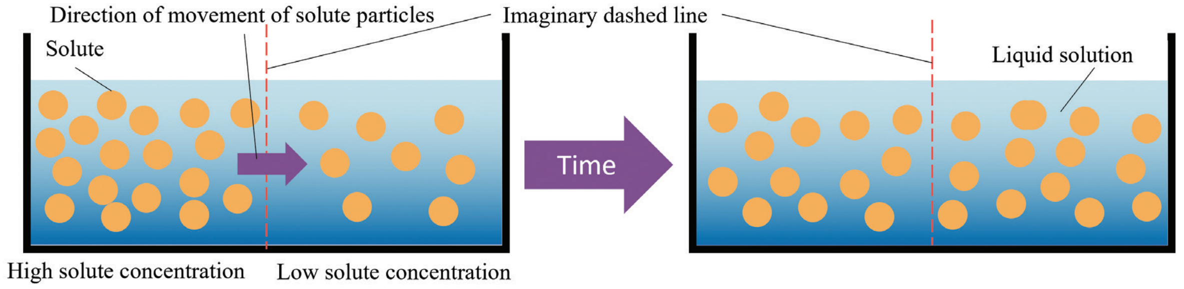


Figure 16.1

Diffusion occurs because of the difference in the concentration of solute particles moving randomly across the space they occupy. The process continues until the solute particles are spread evenly in their solvent as represented in *Figure 16.1* above. The two liquid-filled containers are divided inside by an imaginary dashed line. The container to the left contains two equal compartments with a different distribution of the solute particles which are in a state of random movement. The solute particles move from left to right, from an area where they are more concentrated to an area where they are less concentrated. They diffuse or spread out until they equalize the concentration throughout the solution. This occurs naturally because the system seeks a state of balance and therefore achieves equilibrium. At equilibrium, the particles in the solution continue moving in both directions. As shown in the figure above the container to the right consists of two compartments with the solution in a state of diffusion equilibrium.

Diffusion takes place in liquids and gases because their particles can move quickly and around each other and mix spontaneously. It is important for living organisms as some substances move in and out of cells.

- **Osmosis**

Osmosis is the movement of solvent participles (usually water) through a selectively semipermeable membrane from an area of low concentration to an area of higher concentration of solute. The semipermeable membrane is a thin layer of material which allows solvent molecules but block solute particles to pass through. It separates the solution from the pure solvent. Osmosis is a specific type of diffusion. It can be demonstrated using U – tube shown in *Figure 16.2* below. The left arm of the tube contains pure water and the right arm − aqueous solution of sugar. The water molecules move through the membrane from left to right (from pure solvent to solution) reach the solution and try to produce the same concentration on both sides. The rate of the movement of water molecules (from left to right) is greater than the rate in the opposite direction. As a result, the level of the solution in the right tube rises. The water molecules continue flowing through in either direction with different rates until the levels in the two arms stop changing which means that equilibrium has been reached.

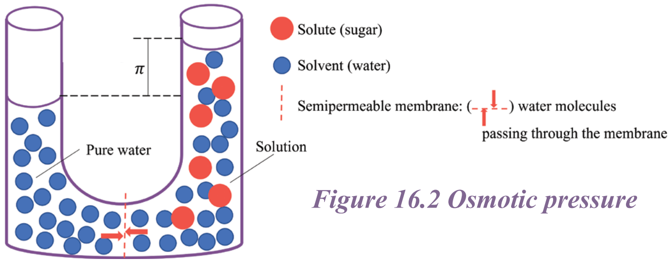


Figure 16.2 Osmotic pressure

- **Osmotic pressure**

Osmotic pressure is the pressure that needs to be applied to the solution side to stop the molecules of a solvent move through a semipermeable membrane. It is equal to the pressure that is applied to the right arm of the tube. The osmotic pressure is proportional to the concentration of dissolved solute particles in each volume of a solution: $\pi = c.R.T = \frac{n}{V}.R.T$

π – osmotic pressure; c – molar concentration of solution, mol/L; R – gas constant ($0.08206\ L.atm/mol.K$); T – the absolute temperature in degrees Kelvin.

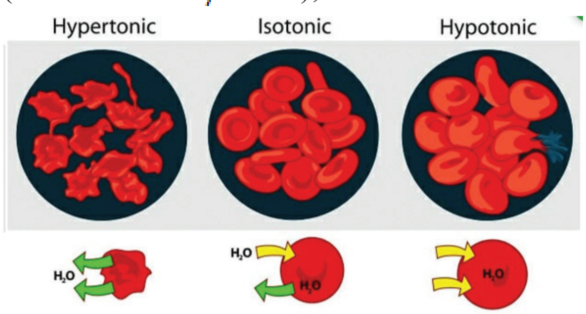


Figure 16.3
https://biologydictionary.net/osmotic-pressure/

- **Applications of osmotic pressure**

Osmotic pressure is of essential importance as it affects cells in living organisms. This effect is illustrated in *Figure 16.3* to the left. Assume that each blood cell is in a solution of sugar or salt water. When the red blood cells are placed in a **hypertonic solution**, the water moves out of the cells faster than it comes in. As a result, the cells shrink. The osmotic pressure of the solution outside the cell is higher than the osmotic pressure inside the cell. When the blood cells are placed in an **isotonic solution**, the osmotic pressure inside and outside the cells is the same. The cells are in their normal size, and it is the normal condition for them in the plasma. If the cells are placed in a **hypotonic solution**, the osmotic pressure inside the cell is higher than the osmotic pressure outside the cell. The water enters the cells faster than it leaves. The cells are in a solution with a lower solute concentration than their own. This causes the cells to swell and burst.

Test your knowledge:

1. Complete the following sentence: Diffusion is the______ movement of solute particles from a region of ______ concentration to a region of ______ concentration until they are ______ spread out.
2. Many people like wearing perfumes and usually choose the fragrance according to their preferences. When a bottle of perfume is being opened, the molecules of the scent escape from the container and spread outward in every direction. Explain the process in relation to diffusion.
3. What occurs during the osmosis? Choose the letter of the best answer.
   a. Pure solvent and the solute both diffuse at the same time through a membrane.
   b. Pure solute diffuses through a membrane, but solvents do not.
   c. Pure solvents diffuse through a membrane, but solutes do not.
   d. Gases diffuse through a membrane, but liquids do not.
4. Calculate the osmotic pressure of the solution containing 0.3 mol/L sugar solution at 25°C?

A) Multiple choice questions. Choose the correct answer in each of the following questions:

1. **Colligative properties include:**
   a. vapour pressure, type of a solution, freezing point depression, boiling point
   b. boiling point elevation, freezing point, vapour pressure, osmotic pressure
   c. boiling point, freezing point depression, vapour pressure, osmotic pressure
   d. boiling point elevation, freezing point depression, vapour pressure, osmotic pressure

2. **Colligative properties depend on:**
   a. the concentration of dissolved particles
   b. the nature of the solute
   c. the nature of the solvent
   d. the chemical formula of the solute

3. **What happens to the freezing point and the boiling point of a solution when a solute is added?**
   a. The freezing point increases but the boiling point decreases.
   b. The freezing point decreases and the boiling point decreases.
   c. The freezing point increases and the boiling point increases.
   d. The freezing point decreases but the boiling point increases.

4. **When a solvent such as water is added to a solution, the concentration of the solution:**
   a. decreases
   b. does not change
   c. increases
   d. freezing unchanged

5. **When a nonvolatile solute is added to water, the solution increases its:**
   a. vapour pressure
   b. temperature
   c. boiling point
   d. freezing point

6. **Which of the following aqueous solutions has the lowest freezing point:**
   a. $0.75\ mol/L$ glucose
   b. $0.05\ mol/L$ formaldehyde
   c. $1.5\ mol/L$ ethyl alcohol
   d. $1.0\ mol/L$ benzene

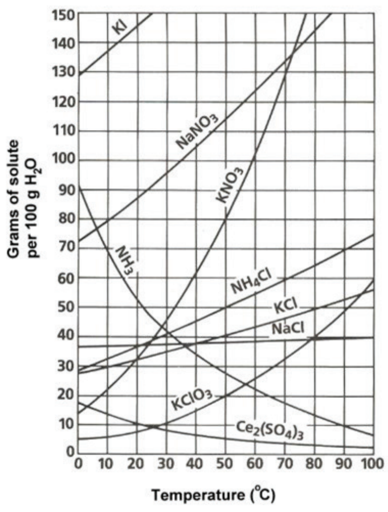
7. **An unsaturated solution contains:**
   a. less solute than the saturated solution
   b. more solute than the saturated solution
   c. the same amount of solute as the saturated solution
   d. less solvent than the saturated solutions

8. **Which of the following statements is TRUE about vapour pressure?**
   a. The vapour pressure of the solution is higher than that of a pure solvent at the same temperature
   b. The vapour pressure of the pure solvent is equal to the vapour pressure of the solution at the same temperature.
   c. The vapour pressure of the solution is less than that of the pure solvent at the same temperature.
   d. The vapour pressure of the pure solvent is lower than the vapour pressure of the solution at the same temperature.

B) What is the molarity of a solution in which the osmotic pressure is $2\ atm$ at $27°C$? Assume the solute is a nonelectrolyte.

C) Calculate the osmotic pressure (in atmospheres) at $22°C$ of $1.00\ L$ solution containing $36.3\ g$ of glucose ($M(C_6H_{12}O_6)) = 180.16\ g/mol$.

D) Sodium chloride is often used to clear icy roads and pavements. Explain how this process works in terms of colligative properties.



E) Use the provided solubility graph to determine the answers in the following questions:

a. If $80\ grams$ of NH_4Cl are dissolved in water at $60°C$ would the resulting solution be unsaturated, saturated, or supersaturated?

b. If $135\ g$ of KNO_3 are added to $100\ g$ of water at $50°C$, how many grams do not dissolve?

c. What mass of NH_3 is needed to form a saturated solution if the NH_3 was dissolved in $150\ g$ of water at $40°C$?

d. Determine which of the following solutions is more concentrated at $30°C$: a saturated solution of $Ce_2(SO_4)_3$ or a saturated solution of $KClO_3$.

F) The table below shows the vapour pressure of three different liquids in the range $20°C - 25°C$. The data have been taken from Engineering ToolBox. Retrieved from https://www.engineeringtoolbox.com/. The external pressure at sea level is $101.3\ kPa$. Rank the liquids in order of:

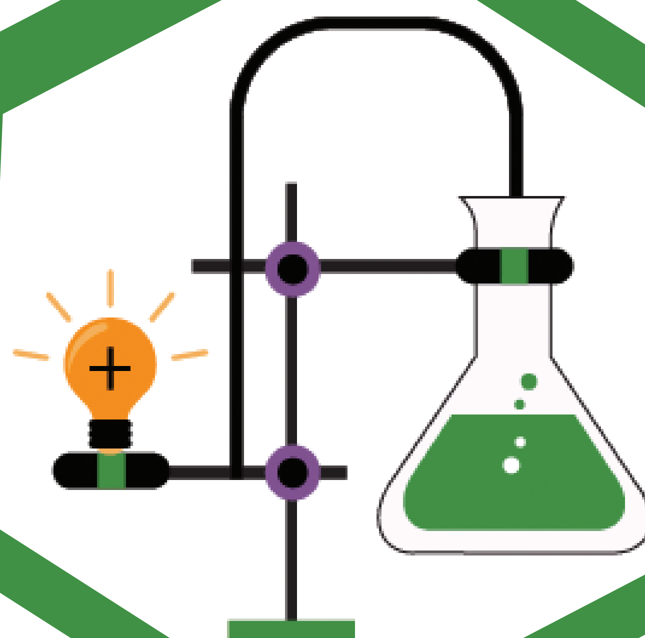
Substance	ethyl alcohol	hexane	acetic acid
Vapour pressure, kPa	12.4	17,6	2,1

a. increasing boiling point
b. decreasing freezing point
c. decreasing osmotic pressure

G) Which aqueous solution has the lowest freezing point?

a. $0.30\ mol/L$ glucose
b. $1.2\ mol/L$ methanol
c. $1.5\ mol/L$ ethanol
d. $0.5\ mol/L$ propanol

SOLUTIONS OF ELECTROLYTES



18 ELECTROLYTIC DISSOCIATION

1. Electrolyte and nonelectrolyte solutions

An electrolyte is a substance that ionises in an aqueous solution. The ions (charged particles) the electrolyte forms move freely in the solution, conducting electricity. Electrolytes are soluble compounds with an ionic structure, such as alkalis and salts or polar covalent bonds such as acids.

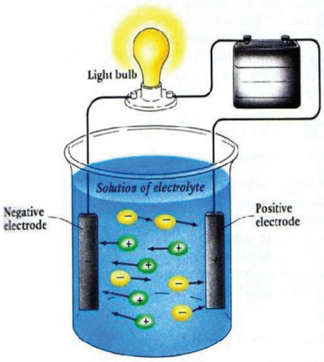


Figure 18.1
https://www.slideshare.net/

Figure 18.1 shows an apparatus for testing the electrical conductivity of a solution. It consists of two electrodes immersed into an aqueous solution of an electrolyte and connected to the positive and negative terminal of a battery. One of the terminals of the battery is connected between the battery and the electric bulb on each side. The lamp lights up, which indicates that an electric current passes through it. A nonelectrolyte is a substance that does not produce ions when it is dissolved in an aqueous solution. Its molecules remain intact (whole). Therefore, it cannot conduct electricity. Nonelectrolytes are most organic compounds (fats,sucrose, alcohols, etc.) and compounds of nonpolar molecules, such as hydrogen, oxygen, etc.

2. Electrolytic dissociation (ionisation)

2.1. Definition: The properties of solutions of electrolytes were explained in 1884 by the Swedish scientist Svante Arrhenius. Arrhenius's theory of electrolytic dissociation states that the molecules of an electrolyte separate (dissociate) in water into positive ions, termed cations and negative ions, termed anions. **The separation of an electrolyte into ions during dissolution in water is called electrolytic dissociation.**

For example, when sodium chloride is dissolved in water, it produces sodium cations and chloride anions according to the equation: $NaCl \xrightarrow{H_2O} Na^+ + Cl^-$

2.2. Mechanism of electrolytic dissociation

✓ **Dissociation of substances with an ionic structure**

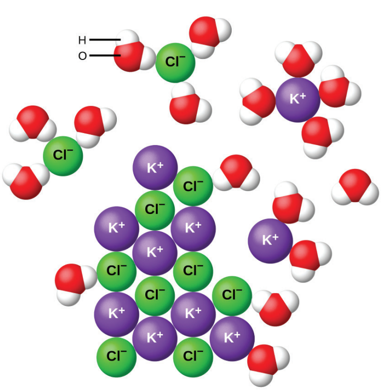


Figure 18.2
https://socratic.org/questions/

When an ionic substance, such as KCl is placed in water (*Figure 18.2*), the water molecules are attracted to the ions on the surface of the lattice. The negative poles of water molecules (oxygen atoms) attract the positively charged potassium cations (K^+), whereas the negatively charged chloride anions (Cl^-) are attracted to the positive poles of the water molecules. As a result, the water molecules surround each K^+ or Cl^- ion, weaken the attractions of the ions to one another and the bonds between $K^+$and Cl^-ions in the crystal break.

The hydrated potassium and chloride ions disjoin and move freely in the solution. The dissociation or ionisation of KCl can be expressed by the equation: $KCl \xrightarrow{H_2O} K^+ + Cl^-$

Ionic substances, such as alkalis and bases conduct an electric current in the molten (melted) state as well. Their ions are not freely moving considering that they are held together in the crystal lattice by strong electrostatic forces of attraction. However, if the solids are heated to their melting points, they transform into liquids which undergo electrolytic dissociation. For example, the molten sodium chloride conducts electric current the same way it does when dissolved in water.

Nevertheless, not all solid substances melt on heating. Some burn or decompose.

✓ **Dissociation of substances with a molecular structure**

Acids are compounds with a molecular structure as their molecules are held together by shared polar covalent bonds. When an acid dissolves in water, the polar bonds in its molecule become ionic due to the interaction between the molecules of acid and water. The acid molecules separate into ions. For example, the hydrochloric acid, which is usually prepared by dissolving hydrogen chloride, dissociates into hydrogen cation (H^+) and chloride anion **(Cl^-).** The hydrogen cation, which is only a proton interacts strongly with nonbonding electron pairs of the surrounding water molecules to form a hydronium ion or the aqueous cation (H_3O^+) as visualised in *Figure 18.3*. The electrolytic dissociation of HCl can be expressed by the equation: $HCl \rightarrow H^+ + Cl^-$

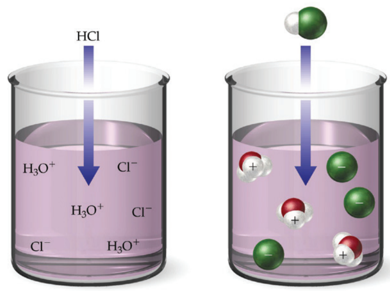


Figure 18.3
https://socratic.org/questions/

The hydrogen cation can (H^+) be written as the hydronium ion (H_3O^+) when an acid is dissolved in water: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$

3. Colligative properties of electrolyte solutions

Both electrolytes and nonelectrolytes have colligative properties. The difference is that electrolytes have a much greater effect on colligative properties than nonelectrolytes. Electrolytes dissociate into two or more ions when are dissolved in a solvent, while nonelectrolytes do not separate into ions, they exist in a molecular form. For example, the boiling point elevation of calcium chloride ($CaCl_2$) is higher than the boiling point elevation of sucrose ($C_{12}H_{22}O_{11}$) of the same concentration because $CaCl_2$ dissociates into three ions per formula unit.

$1\ mol\ CaCl_2 \xrightarrow{H_2O} Ca^{2+} + 2Cl^-$ ($1\ mol\ Ca^{2+}$ and $2\ moles\ Cl^-$); $1\ mol\ C_{12}H_{22}O_{11} \xrightarrow{H_2O} 1\ mol\ molecules$

Test your knowledge:

1. What substances are called electrolytes and nonelectrolytes?
2. Which of the following substances conduct an electric current?

a. potassium iodide solution	d. hydrogen chloride solution
b. molten sodium sulphate	e. molten sucrose
c. acetic acid	f. solid sodium chlorid

3. What is electrolytic dissociation? Explain the role of water in it.
4. Express the dissociation in water of 1 mol of the following compounds:
   $NaOH, HNO_3, HBr, KHCO_3, HCN, K_2SO_4, FeCl_3$
5. Which aqueous solution has a higher boiling point and a lower freezing point: 0.3 mol/L calcium phosphate, $Ca_3(PO_4)_2$ or 0.3 mol/L ethyl alcohol, C_2H_5OH?
6. Which of the following aqueous solutions has a higher osmotic pressure: 0.2 mol/L potassium bromide, KBr or 0.2 mol/L formaldehyde, $HCHO$?

19 STRONG AND WEAK ELECTROLYTES

1. Degree of electrolytic dissociation.

Studies on electrolytes have shown that electrolytes differ in their ability to conduct an electric current as they dissociate to a different extent. Some electrolytes dissociate almost completely into ions, while others - to a little extent. The quantitative measure of dissociation of a given electrolyte in a solution is defined as its degree of dissociation. It is indicated by the Greek letter alpha, α. Degree of dissociation is the ratio of the number of dissociated molecules to the total number of dissolved molecules in a definite solution.

$$\alpha\% = \frac{\textit{number of dissociated molecules}}{\textit{total number of dissolved molecules}} \cdot 100$$

α is expressed either in fractions of unity, which vary from zero (dissociation is absent) to 1 (complete dissociation) or in percent (from 0% to 100%) as it is expressed by the equation above.

The degree of dissociation depends on the:

- **Concentration of electrolyte**. The degree of dissociation of an electrolyte always increases with decreasing concentration of the electrolyte. If a solvent is added (water) to a solution, the separation distance between the ions becomes greater and their recombination into molecules become less probable.
- **Temperature**. With a rise in temperature, α increases because the electrolytic dissociation is an endothermic process.
- **Nature of electrolyte (solute)**
- **Nature of solvent**. The main function of a solvent is to influence the process of breaking up the electrolyte molecules into ions. Each solvent has a different capacity of separating ions. The universal solvent is water, although it does not dissolve nonpolar molecules.

Based on the degree of dissociation, electrolytes are classified as strong, weak and nonelectrolytes.

2. Strong electrolytes and weak electrolytes.

2.1. Strong electrolytes readily dissociate into ions in a solution, so they ionise completely. Therefore, they are excellent conductors of an electric current. The degree of dissociation of strong electrolytes is high and varies from 30% to 100% ($30\% < \alpha \le 100\%$). The process of dissociation of a strong electrolyte is **irreversible**. The general equation for the dissociation of an electrolyte *AB* is as follows: $AB \rightarrow A^+ + B^-$, where A^+ is a positively charged ion (cation), whereas B^- is a negatively charged ion (anion). Strong electrolytes include strong acids (HCl, $HClO_4$, H_2SO_4), strong bases ($NaOH$, KOH, $Ca(OH)_2$, etc.) and salts ($NaCl$, K_2SO_4, $Mg(NO_3)_2$, etc.).

2.2. Weak electrolytes separate into ions with difficulty, so they dissociate slightly. They create a small electric current as they are poor conductors. Weak electrolytes have a much lesser degree of dissociation than strong electrolytes - $\alpha < 30\%$. The dissociation of a weak electrolyte is a **reversible** chemical process. Ions present in a solution simultaneously recombine into neutral molecules. Thus, the system attains a state of dynamic equilibrium between dissociated and undissociated molecules. For example, acetic acid dissociates partially in an aqueous solution into a hydrogen cation (H^+) or a proton and an acetate anion (CH_3COO^-). The two oppositely charged ions combine into a molecule of acetic acid at the same time. The equation of the reaction is written as follows: $CH_3COOH \rightleftarrows CH_3COO^- + H^+$. Weak electrolytes include weak acids (HF, H_2S, H_2SO_3, H_3PO_4, H_2CO_3, etc.), weak bases (NH_4OH, $Cu(OH)_2$, $Zn(OH)_2$, etc.), organic bases and organic acids.

3. Monoatomic and polyatomic ions

➢ **A monatomic ion is formed from a single atom.** Its charge is equal to the difference between the number of protons and neutrons. If the number of protons is bigger than that of electrons, the charge is positive. If there is an excess of electrons, the charge is negative. For example, the molecule of $NaCl$ dissociates into two monoatomic ions. The sodium ion has 11 protons (its atomic number is 11) in the nucleus and 10 electrons in its electron shell. Hence, sodium ion has a positive charge (Na^+). The chlorine ion contains protons (its atomic number is 17) and 18 electrons which explains its negative charge (Cl^-). One more example is potassium sulphide (K_2S) which dissociates into two monoatomic potassium cations (K^+) and one sulphide anion (S^{2-}). $K \rightarrow 19$ protons, 18 electrons ($19 - 18 = +1$); $S \rightarrow 6$ protons and 8 electrons ($6 - 8 = -2$). The overall charge of each ionic compound is zero.

Common Polyatomic Ions

Ion	Name	Ion	Name
NH_4^+	Ammonium	CO_3^{2-}	Carbonate
NO_2^-	Nitrite	HCO_3^-	Hydrogen carbonate Or Bicarbonate
NO_3^-	Nitrate	ClO^-	Hypochlorite
SO_3^{2-}	Sulfite	ClO_2^-	Chlorite
SO_4^{2-}	Sulfate	ClO_3^-	Chlorate
HSO_4^-	Hydrogen sulfate Or Bisulfate	ClO_4^-	Perchlorate
OH^-	Hydroxide	$C_2H_3O_2^-$	Acetate
CN^-	Cyanide	MnO_4^-	Permanganate
PO_4^{3-}	Phosphate	$Cr_2O_7^{2-}$	Dichromate
HPO_4^{2-}	Hydrogen phosphate	CrO_4^{2-}	Chromate
$H_2PO_4^{2-}$	Dihydrogen phosphate	O_2^{2-}	Peroxide

➢ **A polyatomic ion contains a group of covalently bonded atoms with a positive or negative overall net charge.** A list of some common polyatomic ions is given in the table below for students' convenience. It is important to know the charges of polyatomic ions as most electrolytes dissociate into one or more polyatomic ions.

For example, calcium phosphate ($Ca_3(PO_4)_2$) breaks apart in water into calcium cations and phosphate anions. The equation of the dissociation of $Ca_3(PO_4)_2$ is as follows: $Ca_3(PO_4)_2 \rightleftarrows 3Ca^{2+} + 2PO_4^{3-}$ In the balanced equation, the number of atoms in 1 mole of $Ca_3(PO_4)_2$ on the left side of the equation is equal to the number of their corresponding ions on the right. One mole of calcium phosphate contains three calcium atoms and two phosphate atoms. so the coefficient in front of the Ca^{2+} is 3 and 2 in front of the PO_4^{3-}.

https://www.expii.com/t/polyatomic-ions-nomenclature-compounds

Test your knowledge:

1. What does the expression "*The degree of dissociation of hydrosulphuric acid* (H_2S) *in a solution with concentration of* 0.1 mol/L, *at* 18°C *is* 0.1% ($\alpha = 0.1\%$)" mean?

2. Identify the following compounds as either electrolytes, or nonelectrolytes by putting a check in the box.

Compound	$HF_{(aq)}$	C_2H_5OH	$C_{12}H_{22}O_{11}$	$CuSO_4$	H_2CO_3	CH_3OH	$Ba(OH)_2$
Electrolytes							
Nonelectrolytes							

3. Express the dissociation in water of 1 mol of the following compounds:

$$NaOH, HNO_3, HBr, KHCO_3, HCN, Sr(OH)_2, FeCl_3, NH_4Cl$$

4. A solution of a weak electrolyte is given. How can its dissociation degree be increased?

5. Write the molecular formula of each of the following electrolytes if you know the ions in the aqueous solution.

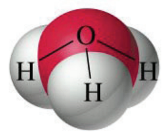
a. ____ $\rightarrow 2Rb^+ + SO_4^{2-}$ d. ____ $\rightarrow Cr^{3+} + 3Cl^-$

b. ____ $\rightarrow Ca^{2+} + 2NO_3^-$ e. ____ $\rightarrow 2NH_4^+ + SO_4^{2-}$

c. ____ $\rightarrow 3H^+ + PO_4^{3-}$ f. ____ $\rightarrow 2K^+ + Cr_2O_7^{2-}$

20 ARRHENIUS ACID-BASE DEFINITION. THE *pH* SCALE

1. Acids. According to Arrhenius theory of electrolytic dissociation, acids are substances that dissociate into hydrogen cations (H^+) or protons and anions (A^-) when are they are dissolved in water. The general equation for the acid dissociation is as follows: $H_nA \xrightarrow{H_2O} n\,H^+ + A^-$



Hydronium ion
H_3O^+
Figure 20.1
https://deskarati.com/wp-content/uploads/2014/02/hydronium.jpg

As the hydrogen cation (H^+) is produced, it cannot exist alone in a water solution. Instead, it immediately bonds with water molecules to form a hydronium ion (H_3O^+). The high positive charge density of the hydrogen ion (proton) makes it attract the negative pole of water molecules and form covalent bonds with one of them.

So, a more accurate equation for the dissociation of any acid is: $H_nA \xrightarrow{H_2O} n\,H_3O^+ + A^-$. In practice, the (H_3O^+) is traditionally referred to as (H^+). **Acids can be monoprotic and polyprotic (diprotic and triprotic)** based on the number of hydrogen ions which an acid produces when is ionised in water. **Monoprotic acids** can produce only one hydrogen ion when ionised in water. I mol of a monoprotic acid ionises in $1\,mol$ hydrogen cation and $1\,mol$ an anion. Monoprotic acids are HCl, HBr, HNO_3, CH_3COOH, C_6H_5COOH, etc. **Polyprotic acids** produce two or three hydrogen ions when dissolved in water. **Diprotic acids** produce $2\,mol\,H^+$ (protons) when ionised in water. Diprotic acids are H_2SO_4, H_2CO_3, H_2S, $H_2C_2O_4$, etc. **Triprotic acids**, such as H_3PO_4 and $C_6H_8O_7$, produce $3\,mol\,H^+$ when ionised in water. The ionisation of polyprotic acids proceed in stages, as it shown below.

$1\,mol$ sulphuric acid produces $2\,mol\,H^+$

Stage 1: $H_2SO_4 \overset{H_2O}{\Longleftrightarrow} H^+ + HSO_4^- \quad \alpha_1$

Stage 2: $HSO_4^- \overset{H_2O}{\Longleftrightarrow} H^+ + SO_4^{2-} \quad \alpha_2$

Total: $H_2SO_4 \overset{H_2O}{\Longleftrightarrow} 2H^+ + SO_4^{2-} \quad \alpha_1 > \alpha_2$

$1\,mol$ phosphoric acid produces $3\,mol\,H^+$

Stage 1: $H_3PO_4 \overset{H_2O}{\Longleftrightarrow} H^+ + H_2PO_4^- \quad \alpha_1$

Stage 2: $H_2PO_4^- \overset{H_2O}{\Longleftrightarrow} H^+ + HPO_4^{2-} \quad \alpha_2$

Stage 3: $HPO_4^{2-} \overset{H_2O}{\Longleftrightarrow} H^+ + PO_4^{3-} \quad \alpha_3$

Total: $H_3PO_4 \overset{H_2O}{\Longleftrightarrow} 3H^+ + PO_4^{3-} \quad \alpha_1>\alpha_2>\alpha_3$

The degree of dissociation of sulphuric acid in the first stage (α_1) is greater than that in the second stage (α_2). Phosphoric acid is triprotic and its degree of dissociation in the first stage is the greatest. Hence, the degree of dissociation in the first stage is the greatest and decreases for the following stage.

Acids possess some common properties because of the presence of hydrogen ions. Acids have a sour taste (lemon, vinegar), have a pH less than 7, turn blue litmus paper red, conduct electricity when dissolved in water, react with metals, basic oxides, and basic hydroxides (bases).

Bases. Arrhenius bases are substance that produce metal cations (M^{n+}) and hydroxide anions (OH^-) when they are dissolved in water. Therefore, they are soluble in water and are known as alkalis. The general equation for the dissociation (ionisation) of bases is as follows: $M(OH)_n \xrightarrow{H_2O} M^{n+} + n\,OH^-$. Depending on the number of OH^- ions, bases can be monovalent and polyvalent (di- and trivalent). The adjective *valent* is used to describe the number of OH^- groups produced from a base when it is dissolved in water. **Monovalent bases** produce one OH^-. Monovalent bases are alkaline metal hydroxides ($LiOH, NaOH, KOH, RbOH, CsOH$). For example, $1\,mol\,KOH$ dissociates into $1\,mol\,K^+$ and $1\,mol\,OH^-$ in water: $KOH \xrightarrow{H_2O} K^+ + OH^-$

20 ARRHENIUS ACID-BASE DEFINITION. THE pH SCALE

Divalent bases produce $2\,mol$ hydroxide anions when dissolved in water. Divalent bases are some alkaline earth metal hydroxides ($Ca(OH)_2$,$Sr(OH)_2$,$Ba(OH)_2$). **Trivalent bases produce** $3\,mol\ OH^-$ when dissolved in water. For example: $Bi(OH)_3$. Their dissociation of polyvalent bases proceeds in stages: For example, $1\,mol\ Ca(OH)_2$ dissociates in $1\,mol\ Ca^{2+}$ and $2\,mol\ OH^-$.

Stage 1: $Ca(OH)_2 \xrightarrow{H_2O} Ca(OH)^+ + OH^-$

Stage 2: $Ca(OH)^+ \xrightarrow{H_2O} Ca^{2+} + OH^-$:

Total: $Ca(OH)_2 \xrightarrow{H_2O} Ca^{2+} + 2OH^-$

The presence of hydroxide ions determines the common properties of bases. They have the pH more than 7, change the litmus paper blue, conduct electricity when dissolved in water, react with acids and acidic oxides.

2. The pH scale. Very sensitive instruments detect an extraordinarily low conductivity of pure water, which explains why water is a weak electrolyte. However, two water molecules can interact with each other to produce a hydronium ion and a hydroxide ion by proton donation (H^+) from one molecule to another. The process is called **self-ionisation** or **autoionisation** of water: The equation of the reaction is: $H_2O + H_2O \rightleftarrows H_3O^+ + OH^-$ or $H_2O \rightleftarrows H^+ + OH^-$, $K_w = c(H_3O^+).c(OH^-) = 1.10^{-14}$ at $25°C$, where K_w is the dissociation constant or equilibrium constant. In pure water, the concentrations of hydronium ion (H_3O^+) and hydroxide ion (OH^-) are equal, $c(H_3O^+) = c(OH^-) = 1.10^{-7}$, so the water is neutral. In an **acidic solution**, $c(H_3O^+) > c(OH^-)$, whereas in a **basic solution**, $c(H_3O^+) < c(OH^-)$. If a solution is **neutral**, $c(H_3O^+) = c(OH^-)$. Because the quantities of H_3O^+ and OH^- are very small, we use a scale, pH (meaning "power/potential of hydrogen") which is the negative logarithm of the concentration of H_3O^+: $pH = -log_c(H_3O^+)$. pH is a scale for indicating the concentration of hydrogen ions. It is used to measure how acidic or basic a solution is. It *Figure 20.2* shows a universal indicator pH colour chart.

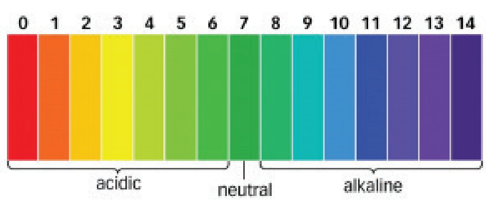


Figure 20.2
https://litmus-paper.com/complete-guide-testing-improving-body-ph/ph-scale/

Test your knowledge:

1. Complete the table below.

Properties	Acids	Bases
pH range		
Conductivity		
Litmus paper		
Common chemical reactions		
Ions (when they dissolved in water)		

2. Classify the following acids as mono-, di- or triprotic: HCN, H_2SO_3, $HClO_4$, HI, H_3PO_4, CH_3COOH, $H_2Cr_2O_7$, H_3BO_3. Express the dissociation of H_2CO_3 stepwise.

3. Express the dissociation of $Sr(OH)_2$ stepwise. Determine if the base is mono-, di- or trivalent.

4. An aqueous solution contains 0.0025 $mol/L\ HCl$. Calculate the $c(H_3O^+)$, $c(OH^-)$ and pH of the solution at 25°C.

(Hint: Use the equations: $c(H_3O+).\ c(OH^-) = 1.\ 10^{-14}$, $c(OH^-) = 1.\ 10^{-14}/c(H_3O^+)$) Answer: ($2.\ 5x10^{-2}\ mol/L$; 2.6; $4.10x10^{-13}$)

21 SALTS

Salts are crystalline compounds which dissociate into ions in aqueous solutions, from the view-point of the theory of electrolytic dissociation. Salts are obtained by either complete or partial neutralisation between an acid and a base.

Salts can be classified in a variety of ways. Based on the quantity of reacting acids and bases, there are three types of salts: normal salts, hydrogen salts and basic salts.

1. Normal salts

Normal salts are formed by the **complete neutralisation** between an acid and a base. A normal salt dissociates in water into a metal cation and an anion. For example, sodium sulphate is dissociated into two sodium cations and one sulphate anion: $Na_2SO_4 \rightarrow 2Na^+ + SO_4^{2-}$

Figure 21.1 . KNO_3
https://en.wikipedia.org/wiki/

All the hydrogen ions of an acid are replaced by the metal ions of a base when they react in an aqueous solution. The hydrogen cation of the nitric acid (HNO_3), for example, is replaced by the potassium cation of the potassium hydroxide (KOH). Theproduct is potassium nitrate which is a soluble salt:

$$KOH + HNO_3 \rightarrow KNO_3 + H_2O$$

One more example of a neutralisation reaction is described by the equation:

$$Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$$

Normal salts are Na_2CO_3, Na_2SO_4, KCl, CH_3COONa, $MgCO_3$, etc.

2. Hydrogen salts

Hydrogen salts are obtained by the **partial (incomplete) neutralisation** of polyprotic acids. A hydrogen salt dissociates in water into two different cations (a metal cation and a hydrogen cation) and an anion of an acid. For example, sodium hydroxide reacts with carbonic acid (dissolved in water carbon dioxide) and yield sodium hydrogen carbonate ($NaHCO_3$), also known as sodium bicarbonate or baking soda. The equation of the reaction is as follows: $NaOH + H_2CO_3 \rightarrow NaHCO_3 + H_2O$

$NaHCO_3$ dissociates in stages:

Stage 1.	$NaHCO_3 \rightarrow Na^+ + HCO_3^-$	α_1
Stage 2.	$HCO_3^- \rightarrow H^+ + CO_3$	α_2
Total:	$NaHCO_3 \rightarrow Na^+ + H^+ + CO_3^{2-}$	$\alpha_1 > \alpha_2$

The name of the hydrogen salts starts with the name of the cation (of the base) followed by the word "hydrogen" before the name of the anion (of the acid). For example, $Mg(HSO_4)_2$ is called magnesium hydrogen sulphate.

3. Basic salts

Basic salts are formed by the **partial (incomplete) neutralisation** of polyvalent bases. A basic salt dissociates in water into a cation and two different anions (a hydroxide anion) and an anion of an acid. For example, magnesium hydroxide chloride is obtained by the partial neutralisation of hydrochloric acid. The equation is as follows: $Mg(OH)_2 + HCl \rightarrow Mg(OH)Cl + H_2O$

$Mg(OH)Cl$ dissociated in stages: Stage 1. $Mg(OH)Cl \rightarrow Mg(OH)^+ + Cl^-$ α_1

Stage 2. $Mg(OH)^+ \rightarrow Mg^{2+} + OH^-$ α_2

Total: $Mg(OH)Cl \rightarrow Mg^{2+} + OH^- + Cl^-$ $\alpha_1 > \alpha_2$

Figure 21.2
https://www.dreamstime.com/inorganic-salts-test-tubes-copper-sulfate-blue-nickel-chloride-green-potassium-dichromate-orange-cobalt-sulfate-brown-sodium-image185361234

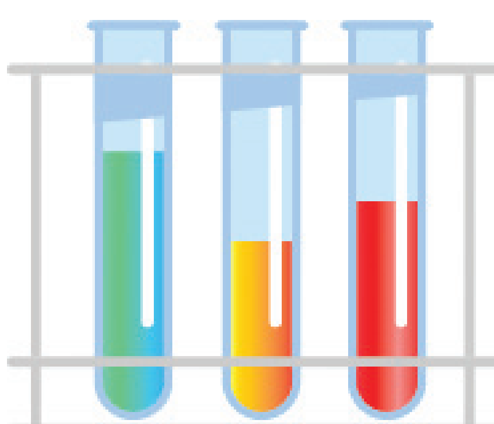
The name of the basic salts starts with the name of the cation (of the base), followed by the name "hydroxide" before the name of the anion (of the acid). For example, $Pb(OH)NO_3$ is called lead hydroxide nitrate.

Figure 21.2 Several inorganic salts in test tubes: copper sulphate blue, nickel chloride green, potassium dichromate orange, cobalt sulphate brown, sodium permanganate dark blue.

Salts are usually electrolytes when they are dissolved in water. Their solutions can be neutral, acidic, or basic. This depends upon the salt used. Aqueous solutions of salts react with acids, bases, and other salts.

Test your knowledge:

1. Classify each of the following salts as a normal, hydrogen or basic salt: $Cd(OH)Cl$, K_2HPO_4, $NaCl$, $Fe(OH)_2NO_3$, LiH_2PO_4, NH_4NO_3, $Al_2(SO_4)_3$, $Ba(OH)Cl$, Na_2S, $CaCO_3$.
2. Express the dissociation of $KHSO_4$ stepwise. What type of salt is it according to the classification described in the lesson?
3. Write the molecular equations of the following reactions:
   a. sodium hydroxide + hydrochloric acid → sodium chloride + water
   b. ammonium hydroxide + carbonic acid → ammonium hydrogen carbonate + water
   c. copper (II) hydroxide + sulphuric acid → copper sulphate + water
   d. calcium hydroxide + nitric acid → calcium hydroxide nitrate + water
4. How can you determine if an aqueous solution of a given salt is acidic, basic, or neutral?

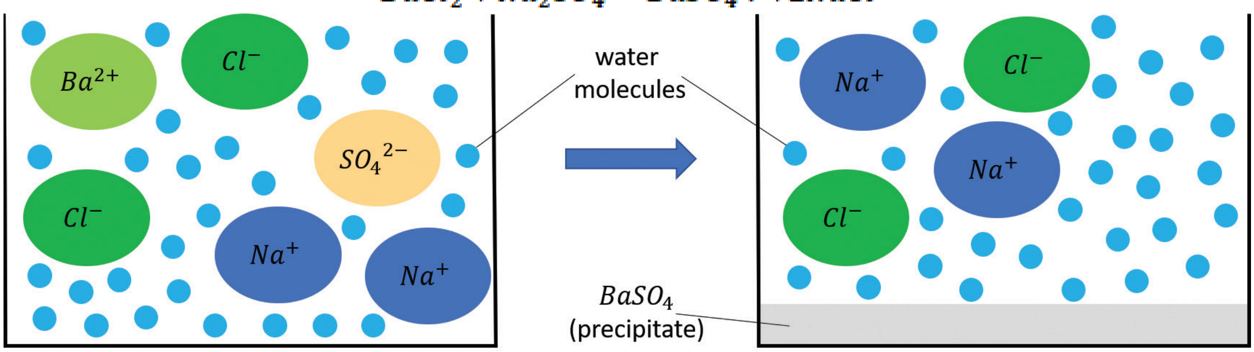


Double replacement reactions usually occur in aqueous solutions. They can be defined as reactions with an exchange of ions between two different compounds. The general form of a double replacement (also called double displacement) reaction is represented by the equation: $AB + CD \rightarrow AD + CB$, where A and B are positively charged ions, while B and D are negatively charged ions. **Double replacement reactions can be classified as precipitation, gas-forming, and acid-base reactions**.

1. Precipitation reactions. A precipitation reaction occurs when two soluble compounds react with each other to produce an insoluble compound called **precipitate**.

A typical precipitation reaction, for example, occurs when an aqueous solution of $BaCl_2$ is mixed with one containing Na_2SO_4. Two reaction products are formed as the result of the reaction – a white precipitate of barium sulphate and a sodium chloride. The chemical equation of the reaction is:

$$BaCl_2 + Na_2SO_4 \rightarrow BaSO_4\downarrow + 2NaCl$$



The figure on the left above shows the ions present in the solution before the reaction starts. Barium chloride ($BaCl_2$) and sodium sulphate (Na_2SO_4) are ionic compounds and when dissolved in water, they separate into ions which move freely throughout the solution. The figure on the right above illustrates the system after the reaction occurs. The solution contains Na^+ and Cl^- ions and a white insoluble solid composed of Ba^{2+} and SO_4^{2-}. The solid settled at the bottom of the container is a precipitate of $BaSO_4$. In chemical equations, the precipitate is expressed by the download (↓) arrow symbol.

It is important to know how substances react with each other in an aqueous solution. Some of them dissolve in water, others produce a precipitate or solid and a small number of compounds react with water. The information about solubility of common substances is summarised into a table known as **the solubility table**. The solubility table or chart (see at the end of the book) is a list of ions of acids, bases, and salts.

The precipitation reactions are written as the **complete ionic equations** which express the separation of soluble reactants and products into their respective cations and anions. In our case, the reaction between $BaCl_2$ and Na_2SO_4 is described by the following complete ionic equation:

$$Ba^{2+} + 2Cl^- + 2Na^+ + SO_4^{2-} \rightarrow BaSO_4\downarrow + Na^+ + Cl^-$$

$BaSO_4$ does not dissolve in water, so it does not separate into ions. Na^+ and Cl^- ions do not change throughout the reaction and they are called **spectator ions**. They can be eliminated from the equation and thus to simplify it. **The net equation** is expressed as follows: $Ba^{2+} + SO_4{}^{2-} \rightarrow BaSO_4 \downarrow$. **The net ionic equation** includes only those ions that are involved in the precipitation reaction.

2. Gas-Forming reactions

A gas-forming reaction is a chemical reaction in which one of the products is a gas, such as oxygen, ammonium, carbon dioxide, hydrogen sulphide, etc. For example, the reaction between sulphuric acid (H_2SO_4) and sodium sulphide (Na_2S) produces sodium sulphate (Na_2SO_4) and hydrogen sulphide gas (H_2S). The evolution(formation) of the gas is indicated by upload (↑) arrow symbol. The chemical equation for the reaction is: $Na_2S + H_2SO_4 \rightarrow Na_2SO_4 + H_2S \uparrow$

The complete ionic equation for the same reaction is as follows:

$$2Na^+ + S^{2-} + 2H^+ + SO_4{}^{2-} \rightarrow 2Na^+ + SO_4{}^{2-} + H_2S \uparrow$$

Na^+ and $SO_4{}^{2-}$ are spectators ions because they do not contribute to the gas-forming reaction.

The net equation includes only H^+ and S^{2-} ions as they are involved in the reaction. It is as follows: $2H^+ + S^{2-} \rightarrow H_2S \uparrow$

3. Acid-Base reactions

An acid-base reaction is a chemical reaction which occurs between an acid and a base in an aqueous solution. The products of the reaction are salt and water. The reaction is called neutralisation if a strong acid is neutralised by a strong base ($pH = 7$). An example of an acid-based reaction is the reaction between calcium hydroxide $Ca(OH)_2$ and nitric acid HNO_3. The resulting products of their interaction are calcium nitrate $Ca(NO_3)_2$ and water H_2O. **The chemical equation is as follows**: $Ca(OH)_2 + 2HNO_3 \rightarrow Ca(NO_3)_2 + 2H_2O$

The complete ionic equation is: $Ca^{2+} + 2OH^- + 2H^+ + 2NO_3{}^- \rightarrow Ca^{2+} + 2NO_3{}^- + 2H_2O$

The net ionic equation is as follows: $2H^+ + 2OH^- \rightleftarrows H_2O$

The net equation includes only the involved in the reaction H^+ and OH^- ions. Ca^{2+} and NO^{3-} ions do not participate in the reaction, so they are spectator ions.

Test your knowledge:

1. Classify each of the following reactions as precipitation, gas-forming, or acid-base. Write the complete ionic and net ionic equations and define the spectator ions.

a. $KOH + HCN \rightarrow KCN + H_2O$
b. $ZnS + 2HCl \rightarrow ZnCl_2 + H_2S \uparrow$
c. $CuSO_4 + 2NaOH \rightarrow Cu(OH)_2 \downarrow + Na_2SO_4$
d. $NaOH + HNO_3 \rightarrow NaNO_3 + H_2O$
e. $CdSO_4 + K_2S \rightarrow CdS \downarrow + K_2SO_4$
f. $Na_2CO_3 + 2HCl \rightarrow 2NaCl + CO_2 \uparrow + H_2O$

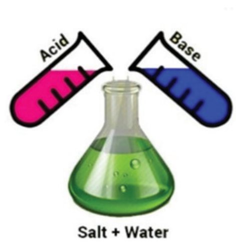
2. Complete the double replacement reactions and then reduce them to the net ionic equations.

a. $Pb(NO_3)_2 + K_2SO_4 \rightarrow$
b. $NaOH + HNO_2 \rightarrow$
c. $CaCl_2 + Na_3PO_4 \rightarrow$

3. Using the table for solubility at the end of the book, predict if the following compounds will be soluble or insoluble in water.

a. Ag_2S
b. PbI_2
c. $NiCl_2$
d. $Fe(OH)_2$
e. $Ba(NO_3)_2$

23 HYDROLYSIS OF SALTS



Hydrolysis is a double replacement reaction between salt and water. The origin of the word is from the Greek word "*hydro*" meaning water and "*lysis*" meaning loosening. Hydrolysis is the reverse of neutralisation. The process is expressed by the equation: $\boldsymbol{acid} + \boldsymbol{base} \underset{hydrolysis}{\overset{}{\rightleftarrows}} \boldsymbol{salt} + \boldsymbol{water}$. There are four possible types of forming salts.

1. Salts that form neutral solutions, $pH = 7$

Salts obtained from the reaction between a strong acid and a strong base do not hydrolyse. This is because the ions of salts do not react with water and therefore form a neutral solution, $pH = 7$. For example, sodium chloride ($NaCl$) is formed from the neutralisation of HCl by $NaOH$. The salt solution is neutral indicating that neither ion is capable of hydrolysing.

Strong acids: HCl, HBr, HI, HNO_3, $HClO_3$, $HClO_4$, H_2SO_4

Strong bases: $LiOH$, $NaOH$, KOH, $RbOH$, $CsOH$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$

2. Salts that form acidic solutions, $pH < 7$

Salts formed by a strong acid and a weak base hydrolyse because the salt cations react with water to produce a weak base. For example, ammonium chloride (NH_4Cl) is a salt formed by the weak base ammonia (NH_3) and the strong hydrochloric acid (HCl). Ammonium ions (NH_4^+) react with water to form ammonium hydroxide (NH_4OH), which is unstable and gradually leaves the solution. Thus, the concentration of hydrogen ions in the solution increases and it becomes acidic. The equation of the reversible reaction is as follows:

$$\underbrace{NH_4^+ + Cl^-} + \underbrace{H^+ + OH^-}_{\substack{\upharpoonleft\downharpoonright \\ H_2O}} \rightleftarrows \underset{\substack{\upharpoonleft\downharpoonright \\ NH_3 + H_2O}}{NH_4OH} + \underbrace{H^+ + Cl^-} \quad pH < 7$$

3. Salts that form basic solutions, $pH > 7$

Salts formed by a weak acid and a strong base hydrolyse because the salt anions react with water to produce a weak acid. For example, sodium carbonate (Na_2CO_3) is formed by the weak carbonic acid (H_2CO_3) and strong sodium hydroxide ($NaOH$). Sodium ions do not attach to hydroxide ions of water molecules because $NaOH$ is a strong electrolyte. Strong attractions appear between the carbonate ion (CO_3^{2-}) and H^+ of water molecules to form H_2CO_3. Thus, the concentration of hydroxide ions (OH^-) in the solution increases and it becomes basic. The equation of the reversible reaction is as follows:

$$\underbrace{2Na^+ + CO_3^{2-}} + \underbrace{2H^+ + 2OH^-}_{\substack{\upharpoonleft\downharpoonright \\ H_2O}} \rightleftarrows \underbrace{2Na^+ + 2OH^-}_{pH > 7} + \underbrace{H_2CO_3}_{\substack{\upharpoonleft\downharpoonright \\ CO_2 + H_2O}}$$

4. Salts of weak acids and weak bases

Salts obtained by neutralisation of weak acids and weak bases form solutions depending on the relative strength of the acid and the base. If the $pH < 7$, $\alpha(acid) > \alpha(base)$, the solution of the salt is acidic. If the $pH > 7, \alpha(acid) < \alpha(base)$, the solution of the salt is basic and the solution is neutral when $pH = 7$ then $\alpha(acid) = \alpha(base)$. For example, the ammonium acetate (CH_3COONH_4) is formed by ammonium hydroxide (NH_4OH) and acetic acid (CH_3COOH). The complete ionic equation is as follows:

$$\underbrace{CH_3COO^- + NH_4^+}_{\substack{\uparrow\downarrow\\ H_2O}} + \underbrace{H^+ + OH^-} \rightleftarrows \underset{\substack{\uparrow\downarrow\\ NH_3 + H_2O}}{NH_4OH} + CH_3COOH$$

The degree of dissociation (α) of CH_3COOH is equal to the degree of dissociation (α) of NH_4OH with the same molar concentration of the two solutions. Therefore, the solution of CH_3COONH_4 is neutral. Some salts of weak acids and weak bases have a weakly basic character, such as $(NH_4)_2CO_3$, others are weakly acidic.

Test your knowledge:

1. Determine which of the following salts are acidic, basic, or neutral. Express the complete ionic equations for those which hydrolyse.

   $BaCO_3$, NH_4CN, $(NH_4)_2CO_3$, Na_2SO_4, CH_3COONa, NaF, KCN.

2. Choose the correct answer:

   A. $NaBr$ can be formed by a reaction between:

      a. a weak acid and a strong base
      b. a weak acid and a weak base
      c. a strong acid and a strong base
      d. a weak acid and a strong base

   B. The pH of an aqueous solution of K_2CO_3 is:

      a. less than 7
      b. equal to 7
      c. greater than 7
      d. none of them

3. Predict the pH for the aqueous solutions of the following salts: $Ba(CH_3COO)_2$, $NaHCO_3$, $ZnCl_2$, K_2SO_4, $Fe(NO_3)_2$.

4. When we eat, the gastric juice in our stomach dissolves food particles and initiate the food digestion. Why is baking soda used to lower the activity of gastric juice?

24 OXIDATION-REDUCTION REACTIONS. OXIDATION STATE

Oxidation-reduction reactions ("redox" for short) involve a transfer of electrons between atoms, molecules, or ions. You need to know what an oxidation state is for better understanding the redox reactions.

1. Oxidation state (OS)

Oxidation state or oxidation number is the total number of electrons lost, gained, or shared in a chemical reaction. It is the positive or negative charge, either actual or imaginary, acquired by each atom in the process of bond formation with the atoms of other elements in a particular compound. Consider the element nitrogen which forms five different oxides with oxygen. In N_2O, nitrogen is in the $+1$ −oxidation state; in NO, nitrogen is in the $+2\ OS$; in N_2O_3, nitrogen is in the $+3\ OS$; in NO_2, nitrogen is in the $+4\ OS$ and in N_2O_5 nitrogen is in the $+5\ OS$. In all oxides oxygen is in the $-2\ OS$.

➢ In compounds with polar covalent bonds, such as the nitrogen oxides, the oxidation state of the elements is the same as the imaginary charges determined by assuming the compounds ionic.

➢ In ionic compounds, the oxidation state is equal to the ionic charge. For example, in calcium chloride ($CaCl_2$) the chemical bond is ionic and the calcium ion a has $2+$ charge (Ca^{2+}), whereas the two chloride ions bear a -1 charge (Cl^-). In $CaCl_2$, calcium is in the $+2\ OS$ and oxygen is in the $-1\ OS$. The respective number of the oxidation states is written above the chemical symbol of the element. Consider the following compounds with indicated OS:

$$\overset{+1}{Na}\overset{-1}{Cl},\ \overset{+2}{Ca}\overset{-1}{Cl_2},\ \overset{+1}{H_2}\overset{-2}{S},\ \overset{+1}{K}\overset{-1}{H},\ \overset{+2}{N}\overset{-2}{O},\ \overset{+1}{N_2}\overset{-2}{O}$$

Rules for assigning oxidation state (OS):

✓ The oxidation state of an element in free form is zero. For example, an uncombined sodium (Na) has $OS = 0$, the molecule of N_2 has $OS = 0$.

✓ Combined oxygen (O) has as a rule oxidation state -2. There are few exceptions: in peroxides, such as H_2O_2, oxygen has $OS = -1$ and in oxygen difluoride, oxygen has $OS = +2$.

✓ Combined fluorine (F) has $OS = -1$.

✓ Combined hydrogen (H) has $OS = +1$ in all compounds, except in the metal hydrides, where it is -1. For example, in NaH, hydrogen has $OS = +1$, in CaH_2, hydrogen has $OS = +2$.

✓ Combined alkali, alkaline earth metals, and group $3A$ metals always have positive oxidation numbers: calcium has $OS = +2$, lithium has $OS = +1$, aluminum has $OS = +3$.

✓ The highest oxidation state of most chemical elements is equal to the number of the group in the Periodic table in which they appear. For example, nitrogen – $5A$ group, $OS = +5$

✓ The sum of oxidation states of all atoms in a neutral compound must be zero.

If the oxidation state of an element in the compound is unknown, we denote it by $\boldsymbol{x}$. Then we write an equation, sticking to the rule: *The algebraic sum of the oxidation states of all atoms in a neutral compound is equal to zero*. For example, in $KMnO_4$, K has $OS = +1$, O has $OS = -2$, M has $OS = x$; $+1 + x + 4.(-2) = 0$, $x - 7 = 0$, $x = +7$. Hence, Mn has $OS = +7$.

2. Oxidising and reducing agents

A reducing agent or **reductant** loses electrons as it gives electrons to another substance and increases its oxidation state. The process is known as **oxidation**.

An oxidising **agent** or **oxidant** gains electrons because it takes electrons from another substance and decreases its oxidation state. The process is known as **reduction**.

Oxidation and reduction take place simultaneously, each of them being a half reaction. In the redox reaction, the atoms of one substance **lose** electrons whereas the atoms in another substance **gain** them. *Figure 24.1.* illustrates the redox reactions:

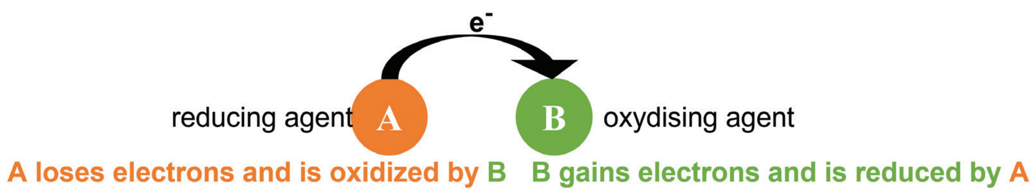


Figure 24.1

A typical example of a redox reaction is the reaction between sodium and chlorine which results in sodium chloride: $2Na + Cl_2 \rightarrow 2NaCl$

Each sodium atom loses one electron to become a sodium ion. Sodium is oxidised because it increases its oxidation state. In contrast, each chlorine atom gains an electron to become a chloride ion. Chlorine is reduced because it decreases its oxidation state. The two half-reactions are as follow:

$$\overset{0}{Na} - 1e^- \rightarrow \overset{+1}{Na}, \overset{0}{Cl} + 1e^- \rightarrow \overset{-1}{Cl}, 2\overset{0}{Na} - 2.1e^- \rightarrow 2\overset{+1}{Na}, \overset{0}{Cl_2} + 2.1e^- \rightarrow 2\overset{-1}{Cl}$$

To express the equations of half-reactions more precisely, we consider that the two sodium atoms react with one diatomic chlorine molecule, so we write additional coefficients in front of the sodium atom and ion and the chlorine ion.

Test your knowledge:

1. Assign the oxidation state of all elements in the following compounds:
   a. Na_2O Na = ___, O = ___
   b. $AlBr_3$ Al = ___, Br = ___
   c. HNO_2 H = ___, N = ___, O = ___
   d. $Ca_3(PO_4)$ Ca = ___, P = ___, O = ___
   e. H_2SO_4 H = ___, S = ___, O = ___
   f. $K_2S_2O_3$ K = ___, S = ___, O = ___
2. Determine the oxidation state of phosphorus in each of the following compounds:
   a. P_2O_5 ___
   b. Ca_3P_2 ___
   c. Na_3PO_4 ___
   d. H_3PO_4 ___
   e. Ag_3PO_4 ___
3. The formula of the compound, in which magnesium has $OS = +2$ and fluorine has $OS = -1$, is:
   a. Mg_2F b. MgF_2 c. MgF d. Mg_2F_2
4. Identify the oxydising and reducing agent in the following reactions:
   a. $2Li + S \rightarrow Li_2S$
   b. $4Fe + 3O_2 \rightarrow 2Fe_2O_3$
   c. $3Mg + N_2 \rightarrow Mg_3N_2$
   d. $Si + 2F_2 \rightarrow SiF_4$
5. Are all alkali metals oxidised or reduced? Explain the reasons.

25 OXIDATION-REDUCTION REACTIONS. THE HALF-REACTION METHOD

Some equations for the redox reactions are complicated and difficult to balance by careful examination. The half-reaction method for balancing redox equations is used to provide a systematic approach. In this method, the overall redox reaction is separated into two half-reactions: one for an oxidation and another for a reduction. The half-reactions are balanced one by one and finally the overall redox reaction is balanced. There are several steps that should be followed to balance a redox equation:

Step 1. Write the unbalanced chemical equation and determine oxidation states of all components.

$$\overset{0}{Cu} + \overset{+1}{H}\overset{+5}{N}\overset{-2}{O_3} \rightarrow \overset{+2}{Cu}(\overset{+5}{N}\overset{-2}{O_3})_2 + \overset{+4}{N}\overset{-2}{O_2} + \overset{+1}{H_2}\overset{-2}{O}$$

In the above example, copper is oxidised by the nitric acid to produce a copper nitrate salt with nitrogen oxide and water.

Step 2. Indicate which element loses electrons and increases its oxidation state. Write the half-equation for oxidation.

In the given example, copper increases its oxidation state, so it is a reducing agent (RA). The half-equation is as follows:

$$\overset{0}{Cu} - 2e^- \rightarrow \overset{+2}{Cu}$$

Step 3. Indicate which element gains electrons and decreases its oxidation state. Write the half-equation for reduction.

Nitrogen decreases its oxidation state, so it an oxidising agent (OA).

$$\overset{+5}{N} + 1e^- \rightarrow \overset{+4}{N}$$

Step 4. Indicate the total number of electrons by balancing the two half-reactions.

$$\left.\begin{array}{l} \overset{0}{Cu} - 2e^- \rightarrow \overset{+2}{Cu} \\ \overset{+5}{N} + 1e^- \rightarrow \overset{+4}{N} \end{array}\right| \, 2 \, \left| \begin{array}{l} \times 1 \\ \times 2 \end{array}\right.$$

To balance the half-reactions, we find the least multiple of the lost and gained electrons and put the coefficients in front of reducing and oxidising agents. In this example, we put the coefficient 2 in front of the NO_2.

$$Cu + HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

Step 5. Determine the additional coefficients in front of the oxidising and reducing agents to equalise the number of atoms on both sides of the redox equation.

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + H_2O$$

Step 6. Balance the oxygen and hydrogen atoms on both sides of the equation and then check to see if all atoms are balanced.

$$Cu + 4HNO_3 \rightarrow Cu(NO_3)_2 + 2NO_2 + 2H_2O$$

25 OXIDATION-REDUCTION REACTIONS. THE HALF-REACTION METHOD

Test your knowledge:

1. Determine the reducing and oxidising agent in each of the following redox reactions:

a. $CO_2 + H_2 \rightarrow CO + H_2O$ RA ___ OA ___
b. $SO_2 + HNO_2 \rightarrow H_2SO_4 + NO$ RA ___ OA ___
c. $Fe + O_2 \rightarrow Fe_2O_3$ RA ___ OA ___
d. $SO_2 + O_2 \rightarrow SO_3$ RA ___ OA ___
e. $H_2S + Cl_2 \rightarrow S + HCl$ RA ___ OA ___
f. $KI + K_2Cr_2O_7 + HCl \rightarrow CrCl_3 + KCl + H_2O + I_2$ RA ___ OA ___

2. Balance the following redox equations using the half-reaction method.

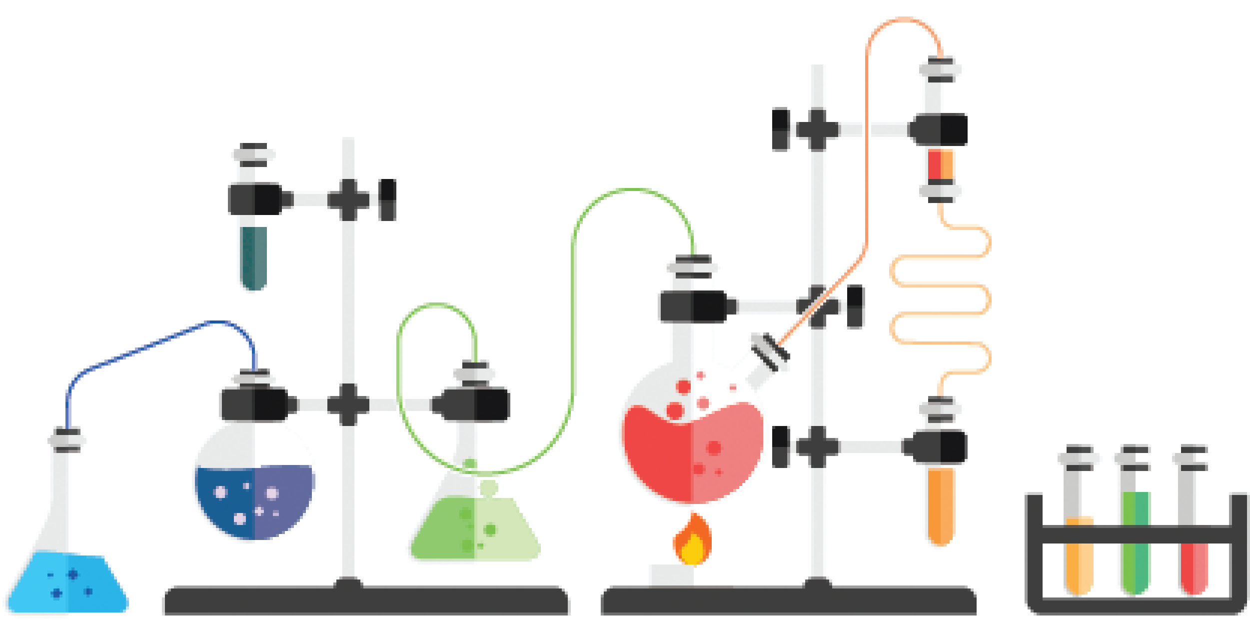
- $Fe_2O_3 + CO \rightarrow Fe + CO_2$
- $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$

3. Decide whether each statement is true or false.

a. Reductant loses electrons and thus decreases its oxidation state.
b. Reductant gains electrons and thus increases its oxidation state.
c. Reductant loses electrons and thus increases its oxidation state.
d. Reductant gains electrons and thus decreases its oxidation state.

4. What is necessary for a redox reaction to occur?

5. Is it possible for a redox reaction to take place without oxygen?



26 THE REACTIVITY SERIES OF METALS

1. Single displacement reactions

Displacement reactions, also known as replacement reactions, allow us to predict the products of a chemical reaction. Displacement reactions are single displacement and double displacement reactions. As you know, the precipitation reaction is a type of a double displacement reaction. Single displacement reactions occur when one element is displaced (substituted) from another in a chemical compound, yielding a new compound and a new element.

For example, $\boldsymbol{Zn + 2HCl \rightarrow ZnCl_2 + H_2}$, where a zinc atom ($Zn$) displaces a hydrogen atom (H) in the molecule of hydrochloric acid (HCl), producing zinc chloride ($ZnCl_2$) and hydrogen (H_2). Zn loses two electrons (reductant), whereas H_2 gains two electrons (oxidant) and thus the reaction is redox.

2. The reactivity series of metals

The products of single displacement reactions are determined by the chart called the **reactivity series of metals**. The metals and their ions are arranged in order of reactivity from the most reactive to the least reactive.

Li/Li^+, K/K^+, Ca/Ca^{2+}, Na/Na^+, Mg/Mg^{2+}, Al/Al^{3+}, Zn/Zn^{2+}, Fe/Fe^{2+}, Ni/Ni^{2+}, Sn/Sn^{2+}, Pb/Pb^{2+}, H/H^+, Cu/Cu^{2+}, Hg/Hg^{2+}, Ag/Ag^+, Pt/Pt^{2+}, Au/Au^{3+}

$\longrightarrow$

Reactivity decreases

Although hydrogen is a nonmetal, it is included in the reactivity series of metals because like metals, it also loses electrons to become a positive ion.

The reactivity series of metals is an effective tool to predict a displacement reaction in which two different metals are involved. Only a metal with higher activity can displace a less reactive metal from its compound. For example, zinc (Zn) is more reactive than copper (Cu) and can displace it from its compounds. In the reaction, Zn acts as a reductant (it changes its OS from 0 to $+2$) and Cu acts as an oxidant (it changes its OS from $+2$ to 0). The equation is as follows:

$$Zn + \underbrace{Cu^{2+} + SO_4^{2-}} \rightarrow \underbrace{Zn^{2+} + SO_4^{2-}} + Cu$$

However, Zn is not capable of displacing the metals on the left side of it. Al, Mg, Na, Ca, K and Li are more reactive than Zn as they are among the most active metals.

To conclude, more reactive metals are stronger reducing agents than less reactive metals.

3. The reactivity series of nonmetals.

In a similar way, nonmetals are arranged in terms of their reactivity. A more reactive nonmetal which is a stronger oxidising agent, displaces a less active nonmetal from its compound during a displacement reaction.

$\boldsymbol{F, Cl, O, Br, I, S, P}$

Reactivity decreases

$\longrightarrow$

For example: the more reactive chlorine can replace the ions of less reactive iodine. The complete ionic equation of the reaction is as follows:

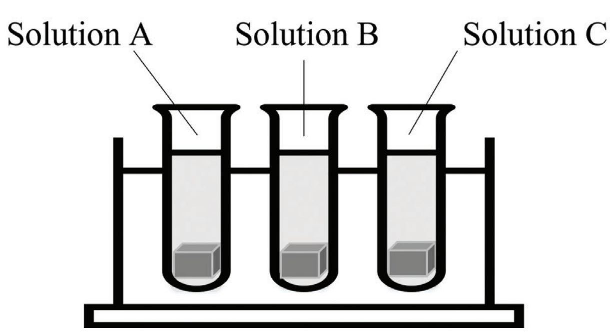
$$\overset{0}{Cl_2} + \underbrace{Ca^{2+} + 2I^-} \rightleftarrows \underbrace{Ca^{2+} + 2Cl^-} + \overset{0}{I_2}$$

Test your knowledge:

1. What determines the chemical reactivity of metals?
2. Consider the following arrangements of metals according to the reactivity series of metals:
   $Na\ Mg\ Zn\ Sn\ Cu$
   a. Which of these metals is the most reactive?
   b. Which metal cannot displace the other metals?
3. The table below shows a few metals and their correspondent nitrates. Put a thick in the box where a reaction is possible.

	Cu	Mg	Zn	Pb
$Cu(NO_3)_2$				
$Mg(NO_3)_2$				
$Zn(NO_3)_2$				
$Pb(NO_3)$				

4. A piece of tin (Sn) is placed into three different metal salt solutions as shown in the figure below. In which solution will a reaction occur?



Solution A - K_2SO_4; Solution B - $CuSO_4$; Solution C - $FeSO_4$

5. Which of the following reactions is not possible?
   a. $Cu + ZnSO_4$
   b. $Fe + Pb(NO_3)_2$
   c. $Ca + HCl$
   d. $Al + HNO_3$

27 ELECTROLYSIS

1. Definition and applications

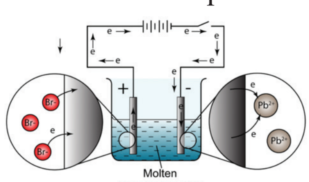
Electrolysis is a process by which electrolytes are decomposed (broken down) into their components when an electric current is passed through them. Electrolytes can be dissolved in water or molten (melted).

Electrolysis has many industrial applications, such as extraction and purification of metals (copper, silver, gold), production of chemicals (chlorine, caustic soda), electrolytic reduction of metals from their compounds (pure aluminum obtained by reducing aluminum oxide) and so on.

The process is carried out in an electrolytic cell. The cell is an apparatus that consists of two electrodes dipped into a dissolved in water or molten electrolyte. When an electric current is applied to the two immersed electrodes, the positive ions move to the negative electrode – **cathode**, while the negative ions move to the positive electrode – **anode**. The electrodes are metal or graphite rods.

2. Electrolysis of molten ionic compounds

When substances are in a solid state, they are not capable to conduct electricity. However, when they are heated up to their melting points, the crystal lattice ions acquire enough energy that allows them to move freely around, and hence, the electrolysis occurs. For example, during electrolysis, metallic lead, and bromine gas both are obtained from the melted lead (II) bromide ($PbBr_2$). The process is represented in *Figure 27.1*. The positive lead ions (Pb^{2+}) are attracted to the negative electrode – cathode where they gain electrons and become lead atoms. Reduction occurs at the cathode. The negative bromide ions (Br^-) are attracted to the positive electrode – anode where they lose electrons and become bromide atoms which couple up to form Br_2 molecules. Oxidation takes place at the anode. The ionic equations for the half-reactions are as follows:



Cathode (-):

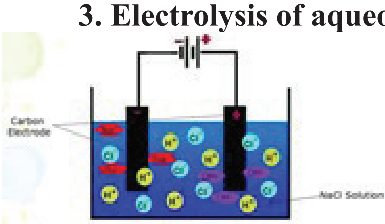
$$Pb^{2+} + 2e^- \to \overset{0}{Pb}$$

Anode (+):

$$2Br^- - 2.1e^- \to 2\overset{0}{Br} \to \overset{0}{Br_2}$$

Figure 27.1 https://igcsechemrevision.wordpress.com/section-1/i-electrolysis/

3. Electrolysis of aqueous solutions



An example of electrolysis in an aqueous solution is the electrolysis of a concentrated solution of $NaCl$, called brine. The process is shown in *Figure 27.2.* and can be described by the following equations: $NaCl \to Na^+ + Cl^-$; $H_2O \rightleftarrows H^+ + OH^-$

Figure 27.2 https://www.slideshare.net/azieda86/63-a-electrolysis-of-an-aqueous-solution

When an electric current is applied, the positive ions present in the solution - Na^+ and H^+ are attracted to the cathode, whereas the negative ions - Cl^- and OH^- are attracted to the anode. The hydrogen ions (H^+) are released at the cathode because they are less reactive than sodium ions (Na^+) (Na^+ are higher up in the reactivity series of metals). H^+ ions gain electrons and become H atoms, which pair up to form H_2 molecules. The process is reduction. At the anode, chloride ions

(Cl^-) are released as they are weaker oxidising agents than hydroxide ions (OH^-). Consider the reactivity series of nonmetals and their ions: $S^{2-}, I^-, Br^-, Cl^-, OH^-, SO_4^{2-}, SO_3^{2-}, NO_3^-, F^-$

The ions are arranged in order of decreasing ability of oxidation.

The chloride ions lose electrons and become Cl atoms which couple up to form Cl_2 molecules. Oxidation takes place at the anode. Redox reaction occurs at the two electrodes. Sodium ions (Na^+) and hydroxide ions (OH^-) stay in the solution and form sodium hydroxide ($NaOH$).

The ionic equations for the half-reactions are as follows:

Cathode (-): $H^+ + 2.1e^- \rightarrow 2\overset{0}{H} \rightarrow \overset{0}{H_2}$

Anode (+): $2Cl^- - 2.1e^- \rightarrow 2\overset{0}{Cl} \rightarrow \overset{0}{Cl_2}$

The electrolysis of brine is an industrial process used to produce important chemicals, such hydrogen, chlorine, and sodium hydroxide.

Test your knowledge:

1. Fill in the gaps to complete the summary about electrolysis.

Word bank: reduction; ions; reactions; molten; cathode; electrolytes; negatively; electric current; electrolyte; takes place.

Electrolysis is a process in which (1) _____ undergo decomposition into (2) _____. They conduct an (3) _____ when dissolved or (4) _____. When an electric current is passed through an (5) _____, the positively charged ions move to the (6) _____ and (7) _____ occurs. The (8) _____ charged ions are attracted to the anode and the oxidation (9) _____. Redox (10) _____ occur at the electrodes.

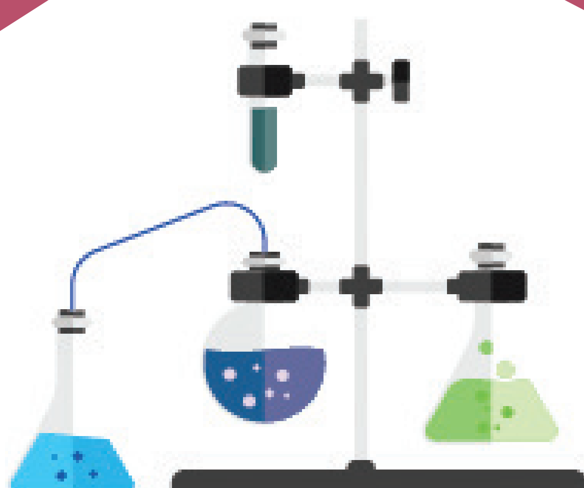
2. The bauxite is an aluminium ore that is purified to produce aluminium oxide (Al_2O_3) from which aluminium can be extracted by hydrolysis. Al_2O_3 does not dissolve in water, but it is melted instead. Identify the ions present in the molten electrolyte and identify the electrodes they move towards:

Compound	Al_2O_3	
Ions present		
Electrode	Cathode (-)	Anode (+)

Express the half-reactions that occur at the cathode and anode.

3. 3. What ions are present in an aqueous solution of copper (*II*) sulphate, $CuSO_4$?
   a. H^+, SO_4^{2-}, OH^-
   b. Cu^{2+}, SO_4^{2-}, OH^-
   c. Cu^{2+}, H^+, SO_4^{2-}
   d. $Cu^{2+}, H^+, SO_4^{2-}, OH^-$
4. What products are formed from the electrolysis of an aqueous solution of zinc chloride, $ZnCl_2$?
   a. Zinc is produced at the cathode and oxygen at the anode.
   b. Hydrogen is produced at the cathode and chlorine at the anode.
   c. Zinc is produced at the cathode and chlorine at the anode.
   d. Zinc is produced at the cathode and hydrogen at the anode.

CLASSIFICATION



28 INORGANIC AND ORGANIC COMPOUNDS

1. Inorganic compounds

Inorganic compounds are substances in which two or more chemical elements are chemically bonded together in definite proportions. They can be divided into oxides, hydroxides, acids, and salts.

Oxides are classified as basic (Na_2O, CaO), acidic (SO_2, CO_2), amphoteric (Al_2O_3), or neutral (CO, NO, N_2O).

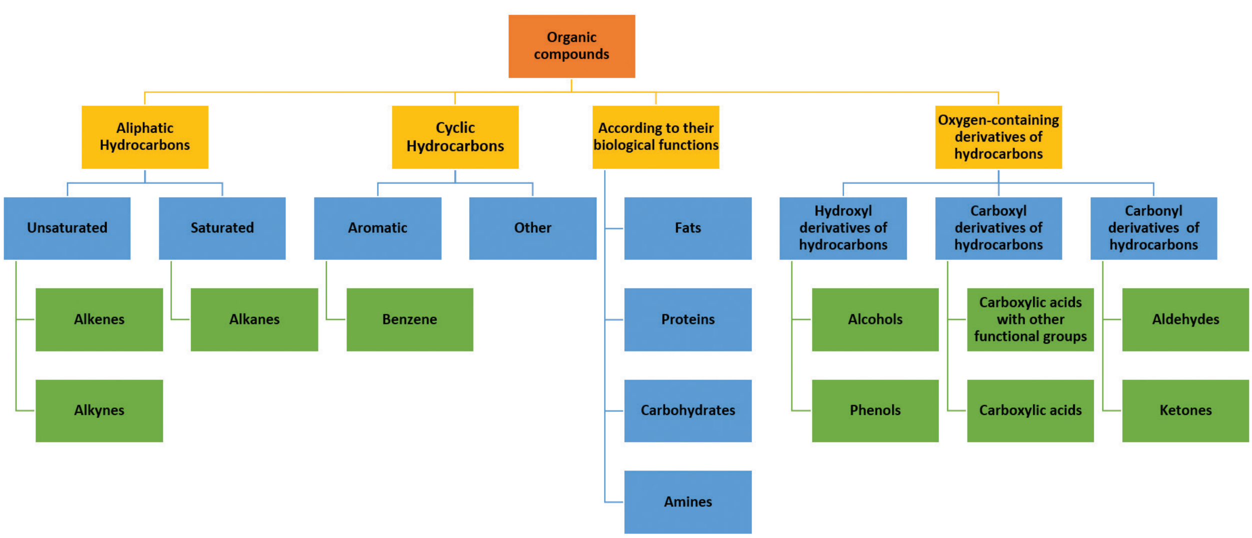
Hydroxides are classified as basic and acidic. Alkali hydroxides ($LiOH$, $NaOH$, KOH, $CsOH$) are soluble in water and form basic solutions. Alkaline earth hydroxides are slightly soluble in water ($Ca(OH)_2$, $Mg(OH)_2$). Amphoteric hydroxides have both amphoteric and basic properties, such as $Al(OH)_3$ and dissolve in basic and acidic solutions.

Acids are hydrogen-containing substances which donate H^+ to another substance (HCl, HNO_3, H_2SO_4).

Salts are ionic substances with only a few exceptions. They can be basic salts ($Ba(OH)Cl$), acid salts ($NaHCO_3$) and neutral salts ($NaCl$).

2. Organic compounds

They are compounds which contain carbon-hydrogen bonds and create 90% of all compounds. They can be classified according to their quality composition.



I. Aliphatic hydrocarbons contain open carbon chains in their molecules and are:

- Saturated hydrocarbons, alkanes, contain simple $C-C$ bonds, (H_3C-CH_3).
- Unsaturated hydrocarbons, alkenes and alkynes contain respectively $C=C$ bonds ($H_2C=CH_2$) and $C\equiv C$ ($HC\equiv CH$).
- Aromatic hydrocarbons contain a ring of six or more carbon atoms For example, C_6H_6 − benzene, $C_6H_5CH_3$ − toluene, $C_{10}H_8$ − naphthalene (it contains two benzene rings).

II. Oxygen-containing derivatives of hydrocarbons. They can be classified as:

➢ **Hydroxyl derivatives** – contain a hydroxyl group, $-OH$, such as alcohols (CH_3OH), phenols contain OH group which is bonded to an aromatic ring (C_6H_5OH).

➢ **Carbonyl derivatives**, such as the aldehydes and ketones contain a carbonyl group ($-\underset{|}{C}=O$) in their molecules.

➢ **Carboxylic acids** – contain a carboxyl group ($-\underset{\underset{OH}{|}}{C}=O$) which is bonded to the carbon atom.

Test your knowledge:

1. **Which of the following compounds is an oxide?**
   a. KOH; b. KCl; c. KNO_3; d. K_2O
2. **Sodium sulphate is:**
   a. an acid; b. an acid salt; c. a neutral salt; d. a basic salt.
3. **Which of the following statements in NOT true?**
   a. Al_2O_3 is a basic oxide.
   b. Al_2O_3 is an amphoteric oxide.
   c. Al_2O_3 is an acidic oxide.
   d. Al_2O_3 is a neutral oxide.
4. **Indicate the formula of the hydrocarbon that contain one $C = C$ double bond.**
   a. C_2H_4; b. C_2H_6; c. CH_4; d. C_2H_2.
5. **Which formula is of a ketone?**
   a. CH_3COOH; b. CH_3COCH_3; c. $HCOOH$; d. CH_3CHO.
6. **Which of the following substances has no toxicity?**
   a. acetone; b. glycol; c. glycerin; d. ethyl alcohol.
7. **Define the organic substances that are essential to human functioning. Provide concrete examples.**

APPLIED CHEMISTRY

29 METALS AND ALLOYS

1. Metals

Metals are elementary substances with atomic structure, in a solid state, except for mercury (Hg) which is a liquid. They conduct heat and electricity relatively well. Metals are separated into main A groups metals and secondary B groups metals. The chemical elements iron, copper and zinc belong to the secondary B groups. They form alloys widely used in everyday life.

➢ Iron (Fe). Iron is a greyish metal of Group VIII B of the Periodic Table, very ductile and malleable. Fe rusts under humid conditions as it covers with a reddish-brown iron oxide (rust) which decomposes it into other chemicals. Iron reacts with acids, and nonmetals because of its $OS = 2; +3$. At room temperature and dry atmosphere, it does not react with oxygen but at temperature up to $200\ ^\circ C$ the surface of the iron is covered with a thin coating of Fe_2O_3 which stops it from reacting. In a dry atmosphere iron wool burns by heating. The equation for the reaction is as follows: $4Fe + 3O_2 \rightarrow 2Fe_2O_3$. The process is exothermic.

➢ Copper (Cu). Copper is a reddish, very ductile metal of Group IB of the Periodic Table. Copper is an excellent conductor of electricity and heat. It has the $OS = +1; +2$ and reacts with oxygen and nonmetals. The metal burns in oxygen with a green flame to form copper oxide. The equation for the reaction is: $2Cu + O_2 \rightarrow 2CuO$

➢ Zinc (Zn). Zn is a silvery-greyish metal of Group IIB of the Periodic Table. Zn is brittle and crystalline at room temperature, but malleable and ductile at temperature between $110°C$ and $150°C$. Zn is a fair conductor of electricity. It has the $OS = +2$ and burns in oxygen at high temperature. The equation for the reaction is: $2Zn + O_2 \rightarrow 2ZnO$. Zinc reacts with halogens, sulphur, and other nonmetals to form compounds. It dissolves in a dilute H_2SO_4 to produce a solution of Zn^{2+}, $SO_4{}^{2-}$ and a hydrogen gas ($H_2 \uparrow$).

2. Alloys

Alloys are mixtures of two or more metallic elements in a solid solution, or a metallic element combined with non-metallic element. Alloys are usually produced by mixing their ingredients in a molten state. After that, they are cooled in special containers to solidify. The table below provides information about the most common and familiar alloys:

Properties of Alloys

Alloy	Composition	Properties	Uses
Bronze	• 90% copper • 10% tin	• Hard and strong • Doesn't corrode easily • Has shiny surface	• To build statues and monuments. • In the making of medals, swords and artistic materials.
Brass	• 70% copper • 30% zinc	• Harder than copper	• In the making of musical instruments and kitchenware.
Steel	• 99% iron • 1% carbon	• Hard and strong	• In the construction of building and bridges. • In the building of the body of cars and railway tracks.
Stainless steel	• 74% iron • 8% carbon • 18% chromium	• Shiny • Strong • Doesn't rust	• To make cutlery and surgical instruments.
Duralumin	• 93% aluminum • 3% copper • 3% magnesium • 1% manganese	• Light • Strong	• To make the body of aeroplanes and bullet trains.
Pewter	• 96% tin • 3% copper • 1% antimony	• Luster • Shiny • Strong	• In the making of souvenirs.

Figure 29.1. https://www.techglads.com/cse/sem1/properties-alloys/

Test your knowledge:

Choose one of the alloys given in the above table or select another one from the Internet resources. Do a project about the properties and applications of the chosen alloy and its ingredients, using different resources, such as pictures, drawings, diagrams, and anything that may be helpful. Your project may include a report, a power point presentation or a poster.

Here are the requirements:

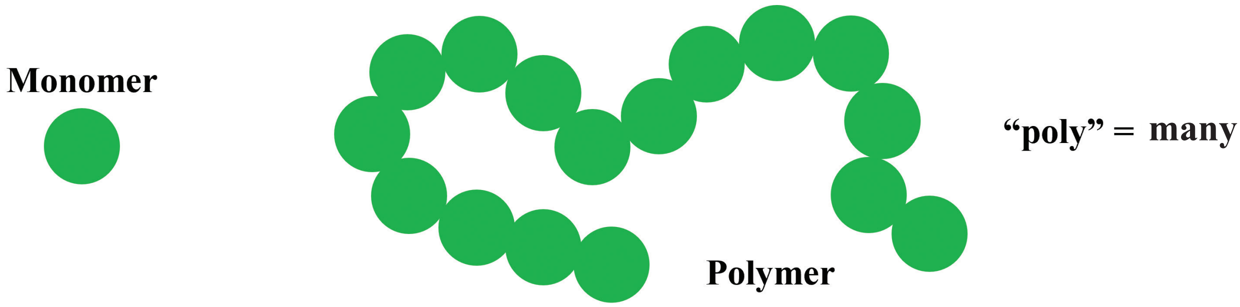
◊ A maximum of three students per project.
◊ Choose a team leader.
◊ Distribute the work equally among the members.
◊ Set realistic project deadlines.
◊ Attach materials permanently to your power point presentation and improve your report.
◊ Design your presentation to be interesting, useful, and original.
◊ Decide exactly what each member will present, especially in combination with their knowledge of the selected alloy.
◊ Present your project in front of the class.

30 POLYMERS, PLASTICS AND CHEMICAL FIBERS

1. Polymers

Polymers are macromolecular compounds that are built of repeating atomic groups called monomers. They contain molecules with extremely large mass and size. For example, polyethylene $(C_2H_4)_n$ is produced from the polymerization of ethylene monomer:

$$nCH_2 = CH_2 \rightarrow (-CH_2-CH_2-)_n$$



2. Plastics

Plastics are materials based on natural and synthetic polymers. They are poor conductors of heat and do not conduct electricity, and therefore they are applied as insulators. Plastics are excellent materials for industry and everyday use because they neither react with acids, bases, and salts nor decay easily. One of the most important plastics is polyethylene.

3. Chemical fibers

Fibers are polymers with a molecular crystalline structure. Fibers can be classified into two main categories:

- ✓ Natural fibers, such as cotton, wool, silk as they are produced by natural raw materials.
- ✓ Chemical fibers, which are man-made fibers. They are artificial and synthetic fibers.
    - Artificial fibers are based mainly on cellulose. They are viscose fibers and acetate fibers.
    - Synthetic fibers are very strong, and they are used as the substitute of wool.

Suggested Presentation Topics

- Polymers. Production of synthetic polymers.
- Plastics. Composition. Properties and applications.
- Health problems raised by plastics during their production and use.
- Chemical fibers. Types and properties.
- Fibers. Applications.
- Health benefits of natural fibers.

Presentation Project

Choose one of the topics given above and do a project using different resources, such as pictures, drawings, diagrams, and anything that may be helpful. Your project may include a report, a power point presentation, or a poster.

Here are the requirements:

◊ A maximum of three students per project.

◊ Choose a team leader.

◊ Distribute the work equally among the members.

◊ Set realistic project deadlines.

◊ Attach materials permanently to your power point presentation and improve your report.

◊ Design your presentation to be interesting, useful, and original.

◊ Decide exactly what each member will present, especially in combination with their knowledge of the selected topic.

◊ Present your project in front of the class.

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Periodic Table of the Elements

Key

11	Atomic number
Na	Element symbol
Sodium	Element name
22,99	Average atomic mass *

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8	9 8B	10	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** Hydrogen 1,01																	2 **He** Helium 4,00
2	3 **Li** Lithium 6,94	4 **Be** Beryllium 9,01											5 **B** Boron 10,81	6 **C** Carbon 12,01	7 **N** Nitrogen 14,01	8 **O** Oxygen 16,00	9 **F** Fluorine 19,00	10 **Ne** Neon 20,18
3	11 **Na** Sodium 22,99	12 **Mg** Magnesium 24,31											13 **Al** Alumi nium 26,98	14 **Si** Silicon 28,09	15 **P** Phosphorus 30,97	16 **S** Sulphur 32,07	17 **Cl** Chlorine 35,45	18 **Ar** Argon 39,95
4	19 **K** Potassium 39,10	20 **Ca** Calcium 40,08	21 **Sc** Scandium 44,96	22 **Ti** Titanium 47,87	23 **V** Vanadium 50,94	24 **Cr** Chromium 52,00	25 **Mn** Manganese 54,94	26 **Fe** Iron 55,85	27 **Co** Cobalt 58,93	28 **Ni** Nickel 58,69	29 **Cu** Copper 63,55	30 **Zn** Zinc 65,39	31 **Ga** Gallium 69,72	32 **Ge** Germanium 72,61	33 **As** Arsenic 74,92	34 **Se** Selenium 78,96	35 **Br** Bromine 79,90	36 **Kr** Krypton 83,80
5	37 **Rb** Rubidium 85,47	38 **Sr** Strontium 87,62	39 **Y** Yttrium 88,91	40 **Zr** Zirconium 91,22	41 **Nb** Niobium 92,91	42 **Mo** Molybdenum 95,94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101,07	45 **Rh** Rhodium 102,91	46 **Pd** Palladium 106,42	47 **Ag** Silver 107,87	48 **Cd** Cadmium 112,41	49 **In** Indium 114,82	50 **Sn** Tin 118,71	51 **Sb** Antimony 121,76	52 **Te** Tellurium 127,60	53 **I** Iodine 126,90	54 **Xe** Xenon 131,29
6	55 **Cs** Cesium 132,91	56 **Ba** Barium 137,33	57 **La** Lanthanum 138,91	72 **Hf** Hafnium 178,49	73 **Ta** Tantalum 180,95	74 **W** Tungsten 183,84	75 **Re** Rhenium 186,21	76 **Os** Osmium 190,23	77 **Ir** Iridium 192,22	78 **Pt** Platinum 195,08	79 **Au** Gold 196,97	80 **Hg** Mercury 200,59	81 **Tl** Thallium 204,38	82 **Pb** Lead 207,2	83 **Bi** Bismuth 208,98	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Ruth erfordium (265)	105 **Db** Dubnium (268)	106 **Sg** Seaborgium (271)	107 **Bh** Bohrium (270)	108 **Hs** Hassium (277)	109 **Mt** Meitnerium (276)	110 **Ds** Darmstadtium (281)	111 **Rg** Roentgenium (280)	112 **Cn** Copernicium (285)	113 **Nh** Nihonium (286)	114 **Fl** Flerovium (289)	115 **Mc** Moscovium (289)	116 **Lv** Livermorium (293)	117 **Ts** Tennessine (294)	118 **Og** Oganesson (294)

58 **Ce** Cerium 140,12	59 **Pr** Praseodymium 140,91	60 **Nd** Neodymium 144,24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150,36	63 **Eu** Europium 151,96	64 **Gd** Gadolinium 157,25	65 **Tb** Terbium 158,93	66 **Dy** Dysprosium 162,50	67 **Ho** Holmium 164,93	68 **Er** Erbium 167,26	69 **Tm** Thulium 168,93	70 **Yb** Ytterbium 173,04	71 **Lu** Lutetium 174,97
90 **Th** Thorium 232,04	91 **Pa** Protactinium 231,04	92 **U** Uranium 238,03	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

* If this number is in parentheses, then it refers to the atomic mass of the most stable isotope.

Solubility Table

Anions \ Cations	H^+	NH_4^+	K^+	Na^+	Ag^+	Ba^{2+}	Ca^{2+}	Mg^{2+}	Zn^{2+}	Cu^{2+}	Pb^{2+}	Fe^{2+}	Fe^{3+}	Al^{3+}
OH^-	■	↑			—		⇣	↓	↓	↓	↓	↓	↓	↓
F^-						⇣	↓	↓	⇣	↓	↓	↓	↓	⇣
Cl^-					↓						⇣			
Br^-					↓						↓			
I^-					↓					—	↓			
S^{2-}	↑				↓	—	—	—	↓	↓	↓	↓	↓	↓
SO_3^{2-}	↑				⇣	⇣	⇣	⇣	⇣	—	↓	⇣	—	—
SO_4^{2-}					⇣	↓	⇣				↓			
NO_3^-														
PO_4^{3-}					↓	↓	↓	↓	↓	↓	↓	↓	—	↓
CO_3^{2-}	↑				↓	↓	↓	↓	↓	↓	↓	↓	—	—
SiO_3^{2-}	↓				↓	↓	↓	↓	↓	↓	↓	↓	↓	↓
CH_3COO^-	■													

soluble	precipitate	gas	slightly soluble	reacts with water	weak electrolyte
	↓	↑	⇣	—	■