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Prepared By:

— Sir Muhammad Usman Ali —

Chemistry Notes Class 9th

According to New Syllabus — For All Punjab Boards



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Chapter 1: States of Matter & Phase Changes

1. Define Chemistry. Write name of any four branches.

The branch of science which deals with the properties, composition and structure of substances.

Branches: 1. Organic chemistry 2. Inorganic chemistry 3. Biochemistry 4. Physical chemistry

2. Define different branches of chemistry.

Branch	Description
Bio Chemistry	It deals with the chemical substances and vital processes occurring in living organisms. In this we study life through chemical processes.
Analytical Chemistry	The branch of chemistry that deals with the analysis of different substances. It involves separation, identification of components present in material things.
Nuclear Chemistry	It deals with the radioactivity, nuclear processes in the nuclei of atom. It also deals with the reactions taking place in the nucleus of an atom. It has many applications in agriculture and medicine.
Physical Chemistry	It deals with the relationship between the composition and physical properties of matter along with the changes in them. It also predicts and changes the rates of reaction.
Environmental Chemistry	The scientific study of the chemical and biochemical events that occur in this planet. It deals with the sources, reactions, effects, and chemical species in the air, soil, and water environment.
Geochemistry	The study of chemical compounds on earth and its sources and minerals is called geochemistry.
Polymer Chemistry	Polymer chemistry focuses on the properties, structure and synthesis of polymers and macromolecules. Polymers are large molecules made by linking together a series of building blocks.
Medicinal Chemistry	It deals with design and synthesis of medicine and drugs which are beneficial for mankind. It includes discovery, delivery, and metabolism of drugs in human body.
Astrochemistry	It deals with the chemical composition and reactions that occur in space, especially in stars, planets, interstellar clouds, comets, and galaxies.

3. Differentiate between Organic and Inorganic Chemistry.

Organic Chemistry	Inorganic Chemistry
Organic Chemistry is the study of covalent compounds of carbon and hydrogen (hydrocarbons) and their derivatives.	Inorganic chemistry deals with the study of all elements and their compounds except those of compounds of carbon and hydrogen (hydrocarbons) and their derivatives.

4. Write scope of organic compounds and inorganic compounds. OR Write uses of organic chemistry and inorganic chemistry.

Scope of Organic chemistry: These are used as food, medicines, cosmetics, dyes etc.

Scope of Inorganic chemistry: These are used as fuel, catalyst, medicines, mineral etc.

5. Write scope of analytical chemistry and biochemistry.

Scope of analytical chemistry: i. It helps to analyze the matter. ii. It helps to analyze food.

Scope of biochemistry: It helps to study biomolecules such as proteins, carbohydrates, fats, vitamins, DNA, RNA etc.

6. Define matter. Give examples. Also write states of matter.

Anything which carries weight and occupies space is called matter. e.g. Chair, book etc.

States of matter: Four states of matter. 1. Solid 2. Liquid 3. Gas 4. Plasma

7. Write properties of primary states of matter.

Solid	Liquid	Gas
Particles are closely packed.	Molecules are closely attached.	Molecules are widely apart.
These are incompressible.	These are not easily compressible.	These are easily compressible.
These have high densities.	These have high densities than gases.	These have very low densities.
These have strongest intermolecular forces.	These have strong intermolecular forces.	These have weakest intermolecular forces.
These have definite shape and volume.	These have definite volume, but not definite shape.	These do not have definite shape and volume.

8. Why are gases easily compressible?

Because gas molecules are widely apart with empty spaces present between molecules, so they are easily compressible.

9. Define Plasma.

The partially ionized gas containing ions, electrons, and photons is called plasma. It exists in fluorescent tubes, lightning and welding arc.

10. What is intermediate state of matter? OR Supercritical fluids, Liquid crystal and Graphene.

Supercritical fluids	Liquid crystal	Graphene
Supercritical fluids are highly compressible states which show both properties of gases and liquids.	Liquid crystal is a state of matter whose properties are between those of conventional liquids and those of crystalline solids. Use: It is used in display devices like computer monitor, clocks, watches and navigation system.	Graphene is a two-dimensional crystal — a single layer of carbon atoms arranged in a hexagonal pattern. It is tough, flexible and has high resistance.

11. Define allotropy and allotropic forms.

The existence of an element in more than one form in the same physical state is called allotropy, and the forms are called allotropic forms. e.g. Allotropic forms of carbon are: diamond, graphite and Buckminster fullerene and allotropic forms of oxygen are oxygen (O₂) and Ozone (O₃).

12. Write physical properties of fullerene.

i. It is soft. ii. It has low melting point. iii. It is poor conductor of electricity.

13. Write physical properties of graphene. OR What is the importance of graphene?

1. It is flexible. 2. It has light weight. 3. It shows high resistance.

14. Why is graphite slippery to touch?

Graphite is slippery to touch because its carbon atoms are arranged in layers, and the layers can slide over each other easily.

15. Differentiate between coal and diamond.

Coal: 1. It is soft. 2. It has amorphous structure. 3. It is used in jewelry (as fuel/power generation industries).

Diamond: 1. It is hard. 2. It has crystalline structure. 3. It is used in power generation (jewelry/cutting tools).

16. Give the reason for brilliance shown by diamonds. Can you improve it?

Diamonds shine with brilliance because they have a high refractive index, which means they bend and reflect light many times inside the stone before it exits. It is due to its unique cut and refraction. To enhance its brightness, cut and polish it precisely.

17. What are Artificial Elements? Give Example.

Those elements which are created by the scientists in the laboratory are called artificial elements. Example: Technetium, Fermium etc.

18. Write the comparison between elements, compounds and mixtures.

Elements	Compounds	Mixture
A substance which consists of only one kind of atoms is known as element.	A pure substance made of two or more elements chemically combined.	The components of a mixture are not chemically bound together.
Elements are only same kind of atoms.	Compounds always have fixed ratio by mass.	Mixture does not have fixed composition.
It cannot be broken into simpler substances by chemical means.	It can be broken into elements by chemical means.	It can be separated by physical method.
Gaseous elements exist in the form of independent molecules.	Compounds have homogeneous composition.	Mixture may be homogeneous or heterogeneous in composition.
Examples: O ₂ , H ₂ , Cl ₂ etc.	Examples: NaCl, H ₂ O, CO ₂ etc.	Examples: Air, Oil, Milk etc.

19. Define solute and solvent.

Solute: The component of solution which is present in smaller quantity.

Solvent: The component of solution present in larger quantity.

Example: In sugar solution, sugar is solute and water is solvent.

20. Define Solution, binary solution, colloidal solution and suspension.

Solution	Binary solution	Colloidal solution	Suspension
Homogeneous mixture of two or more than two substances is called solution. True Solution: A solution in which solute particles are completely homogenized in solvent. Example: Sugar in water	The solution of two components is called binary solution. Example: Sugar in water	A colloidal solution is a heterogeneous mixture where very fine particles are dispersed uniformly in another substance. These particles cannot be seen directly, but they scatter light (Tyndall effect). Example: Milk, blood, starch in water	A suspension is a heterogeneous mixture in which the solid particles are large enough to settle after some time. It can be seen by naked eye. Example: a mixture of chalk in water

21. Define saturated solution. Give example OR how is it prepared?

A solution in which maximum amount of solute has been dissolved at a particular temperature. e.g. adding sugar to water until no more sugar dissolves.

22. Define unsaturated solution. Give example OR how is it prepared?

A solution which contains lesser amount of solute than that which is required to saturate it at a given temperature, is called unsaturated solution. e.g. Take 100g of water in a beaker, add 5g of sugar to it, and stir it. Then sugar will dissolve in water. Then add another 5g of sugar and stir. This will also dissolve.

23. Define solubility. OR Define solubility of a solute. What is effect of temperature on solubility?

The amount of solute which can dissolve in 100g of a solvent at a particular temperature.

Effect of temperature: Generally, solubility increases with the increase of temperature. But it depends on the nature of solute. In some cases, solubility increases with increase in temperature and in some cases it decreases.

24. Give examples of substances whose solubility in water increases with increase in temperature.

Silver nitrate, Sodium nitrate, Potassium nitrate, Potassium chloride

25. Give examples of substances (or compounds) whose solubility in water decreases with increase in temperature.

Lithium carbonate, Calcium chromate, Calcium hydroxide

26. Why does Sugar have Lower Solubility in Water than Sodium Chloride?

Sugar has lower solubility than salt because attraction between sugar molecules and water is weaker while attraction between NaCl molecules and water is stronger.

27. Why do different compounds have different solubilities in water at a particular temperature?

Different compounds have different solubilities due to varying intermolecular forces with water.

28. How can variation of solubility at different temperatures be useful for us?

It is important for: Purification of solids and Dissolving of gases in liquids.

29. Why are Soda Water Bottles stored in Refrigerators?

CO₂ gas is more soluble in water at low temperature. Therefore, soda water bottles are stored in refrigerators to keep carbon dioxide gas dissolved in water for a longer period of time.

30. Why can't NaCl be crystallized from water just like KNO₃?

NaCl has high solubility, making crystallization from water difficult.

LONG QUESTIONS

1. Branches of Chemistry	4. Difference between Elements, Compounds and Mixture
2. States of Matter	5. Effect of temperature on solubility
3. Intermediate state of matter	

EXERCISE**1. Why is there a need to divide Chemistry into many branches? Give three reasons.**

- Chemistry covers a vast range of topics, so dividing it into branches makes it easier to study.
- Different branches focus on specific applications, such as medicine, industry, and environmental science.
- Specialization in branches helps in advancing scientific research and developing new technologies.

2. Reactions may take place due to electrons present outside the nucleus or they may take place inside the nucleus. Which branches of Chemistry cover these two types of reactions?

- Reactions involving electrons outside the nucleus are studied in Physical Chemistry or Organic/Inorganic Chemistry.

- Reactions occurring inside the nucleus are studied in Nuclear Chemistry.

3. What types of problems are solved in analytical chemistry?

- Identification of unknown substances in a sample.
- Determining the composition of chemicals in different materials.
- Determining the purity of pharmaceuticals, food, and industrial products.

4. Both graphite and graphene have hexagonal layered structures. What is the difference?

- Graphite consists of multiple layers of carbon atoms stacked on top of each other with weak forces holding them together.
- Graphene is a single layer of carbon atoms arranged in a hexagonal pattern, making it stronger, more conductive, and more flexible.

5. Why are supercritical fluids important?

- They have unique properties, behaving like both a liquid and a gas, which makes them useful in industrial applications.
- They are used in processes like supercritical fluid extraction (e.g., caffeine removal from coffee).
- They serve as environmentally friendly solvents in chemical and pharmaceutical industries.

6. In which state does matter exist in the Sun?

The Sun's matter exists in the plasma state, a highly ionized form of matter where electrons are separated from nuclei.

7. What is the importance of graphene?

- It is an excellent conductor of electricity and heat.
- It is extremely strong yet lightweight.
- It has applications in electronics, batteries, medical devices, and nanotechnology.

8. Which form of matter do most of the material things in this world belong to?

Most material things in the world belong to the solid state.

Chapter 2: Atomic Structure

1. What is Dalton's Atomic Theory?

It states that all matter is made up of tiny, indivisible particles called atoms.

2. Write the names of fundamental particles of atoms?

Electron, proton and neutron.

3. Write the name of heavy and light elements.

Heavy Elements: Iron

Light Elements: Aluminum, Zinc

4. Name the particles which determine mass of an atom. OR Write the name of particles present in nucleus.

1. Proton 2. Neutron

5. Write the formula to find number of neutrons in an atom.

$$N = A - Z$$

6. Why is the name 'cathode rays' given?

Because they are emitted by the cathode in a discharge tube.

7. Why are elements different from one another?

Because they have different numbers of protons and neutrons.

8. Calculate the number of neutrons, protons and electrons in barium $^{137}_{56}\text{Ba}$.

Number of protons = 56

Number of electrons = 56

Number of neutrons = $137 - 56 = 81$

9. Who discovered Electron, Proton and Neutron?

Electron — J.J. Thomson (1897)

Proton — E. Goldstein (1886)

Neutron — James Chadwick (1933)

10. How are atoms composed?

Atom is composed of electrons, protons and neutrons.

11. Write properties of Cathode particles.

1. These rays have negative charge.
2. These rays are deflected by magnetic field.
3. These rays bend towards the positively charged plate.

12. Write properties of Neutron particles.

1. They are neutral.
2. They are highly penetrating.
3. Their mass is equal to mass of proton.

13. Write properties of proton particles or Positive rays or Canal rays or Anode rays.

1. These rays have positive charge.
2. These rays travel in straight lines.
3. They travel in the opposite direction to cathode rays.

14. Write the difference between shell and subshell.

Shell	Subshell
A shell is the path or level around the nucleus where electrons move.	A subshell is a division of a shell and represents a more specific region where electrons are located.
Shells are denoted by letters K, L, M...	Subshells are denoted by s, p, d, f.

15. Write the number of electrons in K, L, M and N shells.

Number of electrons in a shell can be calculated using the formula $2n^2$, where n is the number of the shell (for K, $n=1$; for L, $n=2$; and so on).

K	L	M	N
2	8	18	32

16. What is Atomic number and Mass number?

Atomic Number / Proton number: The number of protons present in the nucleus of an atom is called atomic number. Symbol: Z

Atomic mass / Mass Number / Nucleon number: The total number of protons and neutrons present in an atom is called mass number. Symbol: A

17. Write the name of subshells and also write no. of electrons in each subshell? OR Write the number of electrons which can be accommodated in sub shells?

s = 2 electrons

p = 6 electrons

d = 10 electrons

f = 14 electrons

18. Why are shells or orbits called energy levels?

Because the electron present in each shell has a fixed energy.

19. Define electronic configuration. Write down the formula. OR Write the maximum capacity of a shell.

The arrangement of electrons around the nucleus is called electronic configuration.

A shell can accommodate a maximum of $2n^2$ electrons.

20. Define isotopes and radioactive isotopes with examples.

Isotopes: Atoms of the same element having the same atomic number but different mass number. e.g. Isotopes of hydrogen are protium, deuterium, tritium.

Radioactive Isotopes: The isotopes which emit energy in the form of radiation are called radioactive isotopes. They emit energy because the nucleus is unstable. e.g. Uranium, radium etc.

21. Why do isotopes of an element show same chemical properties while their physical properties are different?

The isotopes have the same chemical properties because they have the same atomic number, and different physical properties because they have different mass numbers.

22. Write three uses of isotopes. OR Write uses of radioactive isotopes.

1. It is used in medical imaging.
2. They are used to find oil fields.
3. Doctors use them to diagnose disease.

23. What is Carbon dating?

Carbon dating is a method to find the age of once-living things by measuring the leftover carbon-14 in them.

24. Why does a radioactive isotope emit radiation? OR Why are the nuclei of the radioactive element unstable?

Because it has an unstable nucleus, so it emits radiation. The nucleus is unstable because the particles of the nuclei have imbalanced force between them.

25. Why is it said that almost all the mass of an atom is concentrated in the nucleus?

Because protons and neutrons lie in the nucleus, so almost all mass is concentrated in the nucleus.

26. How was it proved that the whole mass of an atom is located at its center?

Rutherford's experiment proved that the whole mass of an atom is located at its center.

27. How can an atom absorb and release energy?

Atoms absorb energy when they jump from a lower energy level to a higher energy level. They release energy when they jump from a higher energy level to a lower energy level.

28. Why is it needed to lower the pressure of the gas inside the discharge tube?

Because current can't pass through the gases at ordinary pressure.

29. Why does the energy of electrons increase as we move from the first shell to the second shell?

Electrons in outer shells are farther from the nucleus, so they have more energy — that's why the energy of electrons increases as we move from the first shell to the second shell.

30. What is Relative Atomic Mass?

The average mass of an atom of that element as compared to $\frac{1}{12}$ th the mass of an atom of carbon-12 isotope. It is represented by the unit atomic mass unit (amu). $1 \text{ amu} = 1.67377 \times 10^{-27} \text{ kg}$.

LONG QUESTIONS

1. Discovery of electron, proton, neutron with their properties	4. Example 2 and Side Box Exercise
2. Bohr's atomic model	3. Isotopes + Application

EXERCISE

1. Why is it said that almost all the mass of an atom is concentrated in its nucleus?

Almost all the mass of an atom is concentrated in its nucleus because the nucleus contains protons and neutrons, which are much heavier than electrons. Electrons have negligible mass compared to protons and neutrons, and since they orbit the nucleus, their contribution to the atom's total mass is minimal.

2. Why are elements different from one another?

Elements are different from one another because they have different numbers of protons in their atomic nuclei. Example: Carbon has 6 protons while Hydrogen has only 1 proton.

3. How many electrons are in $^{210}_{83}\text{Bi}$?

A neutral atom of bismuth-210 has 83 electrons because the number of electrons in a neutral atom is equal to the number of protons, which is 83.

4. Why is tritium ^3_1H a radioactive element?

Tritium (^3_1H) is radioactive because it has an unstable nucleus with an excess of neutrons. It undergoes beta decay, emitting a beta particle (electron) and transforming into helium to achieve a more stable state.

5. How can an atom absorb and evolve energy?

An atom absorbs energy when its electrons jump from a lower energy level to a higher energy level. It evolves (releases) energy when electrons return to lower energy levels.

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Chapter 3: Chemical Bonding

1. Why does an atom form a chemical bond? OR Why does an atom react?

Atoms form chemical bonds to achieve stability by getting a noble gas electronic configuration. Atoms achieve stability when their valence shell is completely filled.

2. Define duplet and octet rule.

Duplet Rule: Attaining 2 electrons in the valence shell is called the duplet rule.

Octet Rule: Attaining 8 electrons in the valence shell is called the octet rule.

Note

H and He are the main elements that follow the duplet rule. While others follow the octet rule.

3. Define chemical bond. Write their types.

The force of attraction which keeps the atoms together is called a chemical bond.

Types: i. Ionic bond ii. Covalent bond iii. Coordinate covalent bond

4. Why does iodine exist as a solid while Cl₂ exists as a gas?

Because iodine has stronger intermolecular forces so it is solid, while Cl₂ has weaker intermolecular forces so it is a gas.

5. What is electronic configuration?

The arrangement of electrons around the nucleus of an atom in shells and subshells is called electronic configuration.

6. Define Ionic Bond or electrovalent bond. Give example.

A bond formed due to complete transfer of electrons from one atom to another is called an ionic bond. e.g. NaCl, KCl, NaF, KBr etc.

7. Define Covalent bond. Give example. Also write their types.

A bond which is formed due to the mutual sharing of electrons between two atoms. e.g. H₂, Cl₂, H₂O etc.

Types: i. Single covalent bond ii. Double covalent bond iii. Triple covalent bond

8. Define Single, Double and Triple covalent bond. Give example.

Single covalent bond: A bond formed by mutual sharing of one electron pair is called a single covalent bond. e.g. Hydrogen (H-H)

Double covalent bond: A bond formed by mutual sharing of two electron pairs is called a double covalent bond. e.g. Oxygen (O=O)

Triple covalent bond: A bond formed by mutual sharing of three electron pairs is called a triple covalent bond. e.g. Nitrogen (N≡N)

9. Define Coordinate covalent bond (Dative covalent bond).

A bond in which the shared pair of electrons is donated by one atom only. The species which donates the electron pair is called the donor and the one which accepts the electron pair is called the acceptor.

e.g. $\text{BF}_3 + \text{NH}_3 \rightarrow [\text{F}_3\text{B} \leftarrow \text{NH}_3]$

10. Differentiate between polar and non-polar covalent bond (pure covalent bond).

Polar Covalent Bond	Non-Polar Covalent Bond
When a covalent bond is formed between two non-identical atoms, the shared pair of electrons bends towards the atom which is more electronegative. e.g. HCl, H ₂ O etc.	When a covalent bond is formed between identical atoms, the shared pair of electrons lies in the middle of the two bonded atoms. e.g. H ₂ , Cl ₂

11. Define metallic bond.

A type of chemical bond which has positively charged ions bound together by mobile electrons.

12. On which factors does the strength of metallic bond depend?

i. Number of free electrons. ii. Number of positive charges on positive ions.

13. Why are alkaline earth metals less reactive than alkali metals?

Because alkaline earth metals lose their valence electrons less easily than alkali metals.

14. Write properties of metals.

i. They show metallic luster. ii. They are good conductors of electricity and heat. iii. They are hard and heavy.

15. Why are metals good conductors of electricity? OR How do metals conduct electricity?

Metals are good conductors of electricity and heat because they have free or mobile electrons.

16. Differentiate between electropositive character and electronegative character.

Electropositive character: The ability of matter to lose electrons is called electropositive character.

Electronegative character: The ability of matter to gain electrons is called electronegative character.

Note

As metals lose electrons more easily, it means greater electropositive character and a direct relation with their reactivity.

17. Why are ionic compounds good conductors of electricity in aqueous or molten form?

Because their ions move freely towards their respective electrodes.

18. Why do ionic compounds not conduct electricity in solid form?

Because in solid form ionic compounds are tightly packed and can't move.

19. Why are covalent compounds not generally soluble in water?

Covalent compounds are not soluble in water because they are non-polar and water is polar.

20. Why do lower molecular mass covalent compounds exist as gases or low boiling liquids?

Due to very weak attractive forces between their molecules, these compounds exist as gases or low boiling liquids.

21. Write three properties of Covalent compounds.

1. They have low melting and boiling points. 2. They are bad conductors of heat and electricity. 3. They are insoluble in water. 4. They are made from 2 or more non-metals.

22. Write three properties of Ionic compounds and metals.

Ionic compounds	Metals
They have high melting and boiling points.	They have high melting and boiling points.
They are good conductors of heat and electricity.	They are good conductors of heat and electricity.
They are crystalline solids.	They are malleable and ductile.
They are soluble in water.	They show metallic luster.
They are neutral compounds.	They are hard and heavy.

23. What is the difference between Ionic Compound and Covalent Compound?

Ionic Compound	Covalent Compound
Compounds in which ionic bonding exists are called ionic compounds.	Compounds in which covalent bonding exists are called covalent compounds.
They have high melting points.	They have low melting points.
They are soluble in polar solvents like water.	They are soluble in non-polar solvents like ether.
They are usually good conductors of electricity in molten state or in solution form.	They are usually bad conductors of electricity.
Example: NaCl, CaCl ₂	Example: H ₂ , Cl ₂

24. What are intermolecular forces?

The forces present between molecules are called intermolecular forces. Intermolecular forces are weaker as compared to chemical bonds. e.g. $\text{H}-\text{Cl}\cdots\text{HCl}\cdots\text{HCl}$

25. Define Dipole-Dipole interaction.

Dipole-dipole interaction is a force between polar molecules. It happens when the positive end of one molecule is attracted to the negative end of another. This attraction helps keep the molecules together. e.g. $\text{H}-\text{Cl}\cdots\text{HCl}\cdots\text{HCl}$

26. Define Hydrogen bonding. Give example.

When hydrogen is covalently bonded to highly electronegative elements like F, O or N, it forms a hydrogen bond. e.g. H₂O

27. What will happen if NaBr is treated with AgNO₃ in water?

A white creamy precipitate of AgBr will be formed when NaBr is treated with AgNO₃.

28. Write uses of coal, graphite and diamond.

Coal	Graphite	Diamond
It is used as fuel to produce electricity.	It is used as a lubricant.	It is used as a cutting/polishing tool.
It is used in industry.	It is used as an electrode.	It is used in jewelry.
It is used in railway engines.	It is used in pencils.	It is used in lenses.

LONG QUESTIONS

1. Ionic Bond	2. Covalent Bond
3. Coordinate covalent bond	5. Metallic Bond
6. Ionic bond/Covalent bond	6. Properties of ionic compound

EXERCISE**1. What type of elements lose their outer electron easily and what type of elements gain electrons easily?**

Metals lose their outer electrons easily, while non-metals gain electrons easily.

2. Why do lower molecular mass covalent compounds exist as gases or low boiling liquids?

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Lower molecular mass covalent compounds exist as gases or low boiling liquids because they have weak intermolecular forces, which require less energy to overcome.

3. Give one example of an element which exists as a crystalline solid and has covalent bonds between its atoms.

Diamond is an example of an element that exists as a crystalline solid and has covalent bonds between its atoms.

4. Which property of metals makes them malleable and ductile?

The property of metals that makes them malleable and ductile is the presence of a sea of delocalized electrons, allowing metal atoms to slide over each other without breaking the metallic bond.

5. Is coordinate covalent bond a strong bond?

Yes, a coordinate covalent bond is a strong bond, comparable in strength to a regular covalent bond.

6. Write down dot and cross formula of HNO_3 .

$\text{H}-\text{O}-\text{N}(=\text{O})(=\text{O})$ — the nitrogen is bonded to three oxygens: one via a single bond to $\text{O}-\text{H}$, one via a double bond, and one via a coordinate (dative) bond, with the nitrogen and oxygen atoms achieving octet configuration.

Chapter 4: Stoichiometry

1. Define Stoichiometry with its use. Which law provides the basis of it?

The branch of chemistry which helps us to calculate the amount of reactants and products by using a balanced chemical equation is called stoichiometry.

Basis of law: It is based on the law of conservation of mass.

Use: Used in industries to determine the amount of raw materials required to produce the desired amount of product.

2. Define reactants and products. Give example.

Reactants: In a chemical reaction, the reacting substances are called reactants.

Products: In a chemical reaction, the produced substances are called products.

Example: $C + O_2 \rightarrow CO_2$ (Reactants: C, O_2 ; Product: CO_2)

3. Define law of conservation of mass.

Matter can neither be created nor destroyed during a chemical reaction. The total mass of all the reactants must be equal to the total mass of products.

4. Define Chemical equation. Give example.

Representing a chemical change in terms of symbols and formulas is called a chemical equation. Example: $C + O_2 \rightarrow CO_2$

5. Define reversible reaction. Give example.

A reaction in which products recombine to produce reactants is called a reversible reaction. e.g. $H_2 + I_2 \rightleftharpoons 2HI$

6. Define Chemical formula. Give example. OR What is meant by chemical formula. Write down chemical formula of barium nitride.

The representation of a chemical compound in terms of symbols and valences is called a chemical formula. e.g. Sodium chloride ($NaCl$), Barium nitride (Ba_3N_2)

7. Differentiate between molecular formula and empirical formula.

Molecular Formula: The molecular formula represents the actual number of atoms present in a molecule of the compound. e.g. Benzene (C_6H_6)

Empirical Formula: The empirical formula represents the whole number ratio of atoms present in that compound. e.g. Benzene (CH)

Chemical Formulae

Calcium phosphate — $Ca_3(PO_4)_2$

Aluminum nitride — AlN

Sodium acetate — CH_3COONa

Ammonium carbonate — $(NH_4)_2CO_3$

Bismuth sulphate — $Bi_2(SO_4)_3$

Barium chloride — $BaCl_2$

8. Find out the molecular formula of a compound whose empirical formula is CH_2O and its molar mass is 180.

Molar mass = 180 g mol^{-1}

Empirical formula mass of $CH_2O = 12 \times 1 + 1 \times 2 + 16 \times 1 = 30$

$n = \text{Molar mass} / \text{Empirical formula mass} = 180/30 = 6$

Molecular formula = $n \times (\text{E.F.}) = 6(CH_2O) = C_6H_{12}O_6$

9. Give two examples of compounds which have the same empirical and molecular formulas.

Compounds	Molecular formula	Empirical formula
Water	H ₂ O	H ₂ O
Ammonia	NH ₃	NH ₃
Methane	CH ₄	CH ₄

10. Write down the names of three compounds which have different empirical and molecular formula.

Compounds	Molecular formula	Empirical formula
Benzene	C ₆ H ₆	CH
Glucose	C ₆ H ₁₂ O ₆	CH ₂ O
Hydrogen peroxide	H ₂ O ₂	HO

Practice numericals

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11. Define molar mass. Give example.

The mass of one mole of a substance is called molar mass. e.g. Molar mass of carbon atom = 12 g

12. Write steps to write the chemical formula of binary ionic compounds.

i. Identify cation and anion. ii. Write cation first then anion with charges. iii. Apply the crisscross method. (Practice Page no. 54)

13. How can you determine chemical formula from its structural formula?

- Write symbols of all the elements.
- Count the number of atoms.
- Write the total number of atoms as subscript.
- Remove the subscript 1.

14. Why does Avogadro's number have immense importance in chemistry?

Avogadro's number relates atoms to measurable amounts easily.

15. Define Chemical change. Give example.

When two or more substances combine in such a way that a new substance is formed which has entirely different properties from its constituents, it is called chemical change. e.g. $C + O_2 \rightarrow CO_2$

16. Different compounds will never have the same molecular formula but may have the same empirical formula. Explain.

Because the empirical formula shows the simplest ratio while the molecular formula shows the actual number of atoms.

LONG QUESTIONS

1. Determine Empirical formula (side box exercise)	2. How to write chemical equation
3. Example 1, 2, 3, 4	

EXERCISE

1. Write the Chemical formula of barium nitride.

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Chemical Formula: Ba₃N₂

2. Find the molecular formula of a compound whose empirical formula is CH₂O and molecular mass is 180 g/mol.

Empirical formula mass of CH₂O = 12 + 2(1) + 16 = 30

$n = 180/30 = 6$

Molecular formula = $n(\text{Empirical formula}) = 6(\text{CH}_2\text{O}) = \text{C}_6\text{H}_{12}\text{O}_6$

3. How many molecules are present in 1.5 g of H₂O?

Molar mass of H₂O = 18 g mol⁻¹

No. of moles = Given mass / Molar mass = $1.5/18 = 0.0833 \text{ mol}$

No. of molecules = Moles $\times$ Avogadro's number = $0.0833 \times 6.022 \times 10^{23} = 5.01 \times 10^{22}$

4. What is the difference between mole and Avogadro's number?

Mole	Avogadro's number
A substance which contains 6.022×10^{23} particles. Mole = Given Mass/Molar Mass. 1 Mole of Carbon = 12 g	A collection of 6.022×10^{23} atoms, ions or molecules is called Avogadro's number. It is represented by N _A . 6.022×10^{23} atoms of C = 1 mole

5. Write the chemical equation for: Copper + Sulphuric acid $\rightarrow$ Copper sulphate + Sulphur dioxide + Water



Chapter 5: Energetics

1. Define energetics or chemical energetics. What do we study in energetics?

The study of energy changes during a chemical reaction is called energetics.

We study the flow of energy in a chemical reaction.

2. Define thermochemical reactions. Write its types.

It is a chemical reaction which takes place with the release or absorption of heat.

Types: 1. Exothermic reaction 2. Endothermic reaction

3. Differentiate between exothermic reactions and endothermic reactions.

Exothermic reactions: A reaction in which heat energy is given out is called an exothermic reaction. e.g. $C + O_2 \rightarrow CO_2 + \text{Heat}$

Endothermic reactions: A reaction in which heat energy is absorbed is called an endothermic reaction. e.g. $H_2 + I_2 \rightarrow 2HI$

4. Write some uses/applications of exothermic reactions.

- Burning of fuels used for cooking, heating homes, running vehicles, and generating electricity.
- During metabolism, the chemical energy present in food is converted to heat to keep us warm.
- Combustion of petrol releases energy which is used to drive vehicles.

5. What is combustion?

Combustion is the process of burning a substance in the presence of oxygen to produce heat and light.

6. Explain why the formation of a bond is always an exothermic process.

Because energy is released to the surroundings when an atom forms a chemical bond.

7. Define chemical energy and heat energy.

Chemical energy: The energy stored in a molecule in which atoms are bonded to each other.

Heat energy: The energy released when a bond is formed and absorbed when it is broken.

8. Define system and surrounding.

System: Any physical or chemical change under study is called a system.

Surrounding: Everything else which does not fall in this system is called surrounding.

Example: If we boil water in a beaker, water is the system while the beaker and burner are the surroundings.

9. Why do energy changes take place during a chemical reaction?

Because breaking and making of chemical bonds takes place during a chemical reaction.

10. How is enthalpy different from heat?

Heat: Heat is the form of energy that flows from a hot body to a cold body because of the temperature difference.

Enthalpy: The total amount of thermal energy stored in a compound. Symbol: H, Unit: KJ/mol

11. When is heat absorbed or released during a chemical reaction?

During a chemical reaction, heat is absorbed when bonds are broken and released when bonds are formed.

12. Can energy be transferred in a form other than heat?

Yes, e.g. chemical reactions can transfer energy in the form of light.

13. Why is heat transferred from a hot to a cold body?

Due to the temperature difference, heat is transferred from a hot body to a cold body.

14. Why is it not possible to calculate the enthalpy of a system?

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Because it is an extensive property, so we can only measure the change in enthalpy.

15. What happens to the enthalpy of a system when energy is absorbed or released?

Enthalpy increases when energy is absorbed; it decreases when energy is evolved.

16. How does a reaction take place?

When two reactant molecules are mixed together, they start colliding with each other. If these collisions have enough energy and proper orientation, the molecules may form a transition state, and after a short time, products are formed.

17. What is activation energy?

The minimum energy required to start a reaction is called activation energy.

18. Define catalyst. Give example.

The substance which can change the rate of reaction, and remains unchanged itself, is called a catalyst. A catalyst speeds up the chemical reaction. e.g. Platinum acts as a catalyst in the production of H_2SO_4 .

19. Differentiate between aerobic and anaerobic respiration.

Aerobic Respiration: The respiration process which takes place in the presence of oxygen.

Anaerobic Respiration: The respiration process which takes place in the absence of oxygen.

20. Define glycolysis and lipids.

Glycolysis: Glycolysis is the process of breaking down glucose into pyruvate, releasing energy in the form of ATP and NADH.

Lipids: Lipids are a group of organic compounds including fats, waxes etc.

21. Explain the role of lipids in our body.

Lipids provide energy, keep us warm and help to absorb vitamins.

22. What is glycogen? Where is it stored? Also write its role in our body.

It is the primary storage form of glucose. It is stored in the liver and muscles.

Role: It is a source of carbohydrate in our body.

LONG QUESTIONS

1. Applications of exothermic reaction	2. How does a reaction take place?
3. Glycolysis	

EXERCISE

1. What is the difference between enthalpy and enthalpy change?

Enthalpy	Enthalpy change
Enthalpy is defined as the total amount of thermal energy stored in a compound. It is denoted by H.	Enthalpy change is the amount of heat absorbed or released by a system during a chemical reaction at constant pressure. It is denoted by ΔH .

2. Why is breaking of a bond an endothermic process?

Bond breaking is an endothermic process because when two atoms combine, energy is stored in them, and when we break the bond, an equal amount of energy is given to break that bond — this energy is also known as dissociation energy.

3. Depict the transition state for the following reaction: $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$

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4. Draw the reaction profiles for two exothermic reactions, one of which moves faster than the other.

(i) $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ and (ii) $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$ — both show an energy vs. reaction-coordinate curve rising from reactants through a transition state at activation energy (E_a), then falling to products below the reactant energy level ($-\Delta H$), with the first reaction having a lower activation energy barrier, allowing it to proceed faster than the second.

5. What is the role of glycogen in our body?

Glycogen is the primary storage form of glucose. It is stored in the liver and muscles. It helps to regulate blood sugar levels in our body.

Chapter 6: Equilibria

1. Differentiate between reversible and irreversible reactions.

Reversible reactions: A reaction in which products recombine to form reactants is called a reversible reaction. e.g. $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

Irreversible reactions: A reaction in which products do not recombine to form reactants is called an irreversible reaction. e.g. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

2. Which type of reactions will never go to completion?

Reversible reactions will never go to completion because the products formed can react again to form the original reactants, leading to a dynamic equilibrium where both reactants and products are present.

3. Why are some reactions irreversible while others are reversible?

Some reactions are irreversible due to stable products and high energy release, while others are reversible because their products can easily change back into reactants.

4. Why are combustion reactions generally irreversible?

Combustion reactions are generally irreversible because they produce stable products like carbon dioxide and water, release a large amount of energy, and involve permanent chemical changes that cannot easily be reversed.

5. What will happen if the temperature of the reaction is increased?

By increasing temperature, the reaction will move in the forward direction.

6. Can you make an irreversible reaction reversible and vice versa?

Yes, by changing conditions like temperature, pressure or concentration you can sometimes make an irreversible reaction reversible and vice versa.

7. What is the effect of the presence of a catalyst in a reversible reaction?

A catalyst is a substance which increases the speed of a chemical reaction. A catalyst can increase the rate of both the forward and reverse reaction of a reversible reaction.

8. What happens when a reversible reaction takes place in the presence of a catalyst? OR How can you decrease the time to attain the position of equilibrium in a reversible reaction?

By adding a catalyst to the reversible reaction, we can decrease the time to attain equilibrium.

9. How do you know if a reaction is reversible or irreversible?

Reversible reaction: reaches equilibrium and proceeds in both directions.

Irreversible reaction: never reaches equilibrium and proceeds in one direction.

10. Write the properties (or characteristics) of Reversible reaction.

i. They never go to completion. ii. They proceed in both ways. iii. These are represented by a double arrow.

11. Write the properties (or characteristics) of Irreversible reaction.

i. The reactions are supposed to be complete. ii. They proceed in one direction.

12. How can a reversible reaction be disturbed? Write down the possible ways.

Reversible reactions can be disturbed by: i. Changing temperature ii. Changing pressure

13. How can you get the maximum yield in a reversible reaction?

Remove products or increase reactant concentration slowly.

14. What is Chemical equilibrium?

When the rates of forward and reverse reactions become equal, it is called the chemical equilibrium state. e.g. $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

15. Difference between static equilibrium and dynamic equilibrium.

Static equilibrium: When a reaction ceases to proceed, it is called static equilibrium. Example: A building remains standing rather than falling down.

Dynamic equilibrium: In dynamic equilibrium, the reaction does not stop, but the rates of forward and reverse reaction become equal at the chemical equilibrium state. e.g. $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

16. How can you disturb the dynamic equilibrium?

We can disturb equilibrium by changing pressure, temperature or concentration.

17. Can physical changes be reversible? Give example.

Yes, physical changes can be reversible. e.g. $\text{Ice} \rightleftharpoons \text{Water}$

18. What are the ways to disturb a reversible chemical system?

- Adding or withdrawing one or more of the reacting species.
- Adding or withdrawing one or more of the product species.
- Effect of catalyst on the reversible reaction.
- Changing the pressure of the reaction.

19. What is the effect of changing temperature on the state of equilibrium?

In exothermic reactions, increasing temperature shifts the equilibrium in the backward direction.

In endothermic reactions, increasing temperature shifts the equilibrium in the forward direction.

20. What is the effect of change of pressure on the state of equilibrium?

Increasing pressure shifts the equilibrium towards the side with fewer gas molecules, i.e., NH_3 .

Decreasing pressure shifts it towards the side with more gas molecules, i.e., N_2 and H_2 .

$\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ — On the left side there are 4 moles of gas and on the right side there are 2 moles of gas.

21. What is the effect of the presence of a catalyst on a reversible reaction?

A catalyst speeds up both the forward and backward reactions equally. So, if a reversible reaction is carried out in the presence of a catalyst, it will decrease the time taken by the reaction to attain the state of equilibrium.

LONG QUESTIONS

1. How is a reversible reaction disturbed?

2. Effect of temperature, pressure and catalyst on state of equilibrium

EXERCISE**1. How is dynamic equilibrium different from static equilibrium?**

Static equilibrium: When a reaction ceases to proceed, it is called static equilibrium. Example: A building remains standing rather than falling down.

Dynamic equilibrium: When the rates of forward and reverse reaction become equal at the chemical equilibrium state, it is called dynamic equilibrium. e.g. $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$

2. How will the following reversible reaction be affected if its temperature is increased? $2\text{H}_2\text{O} \rightleftharpoons 2\text{H}_2 + \text{O}_2$

Increasing the temperature favors the endothermic forward reaction, shifting the equilibrium to the right. As a result, more hydrogen and oxygen are produced, and the amount of water decreases.

3. How can you get the maximum yield in a reversible reaction?

To get the maximum yield in a reversible reaction, adjust conditions (temperature, pressure, concentration) to favor the desired direction, and continuously remove the products to shift the equilibrium toward more product formation.

4. How can you decrease the time to attain the position of equilibrium in a reversible reaction?

You can decrease the time to attain equilibrium in a reversible reaction by adding a catalyst, increasing the temperature, increasing pressure and ensuring proper mixing or stirring to allow reactants to collide more frequently.

5. What is the effect of increasing pressure on the following reaction at equilibrium? $\text{N}_2 + \text{O}_2 \rightleftharpoons 2\text{NO}$

Increasing pressure has no effect on the equilibrium position of this reaction because the number of gas molecules is the same on both sides. Therefore, the equilibrium does not shift.

Chapter 7: Acid, Base Chemistry

1. Differentiate between acid & base. OR Write physical properties of acid & base (or alkalis).

Acid: i. Acid has a sour taste. ii. They turn blue litmus to red.

Bases: i. Bases have a bitter taste. ii. They turn red litmus to blue.

2. Write the name and formulae of mineral acids.

i. Sulphuric acid (H_2SO_4) ii. Hydrochloric acid (HCl) iii. Nitric acid (HNO_3)

3. Write two physical properties of salts. OR Write two important properties of salts.

1. These are neutral compounds. 2. These have high melting and boiling points.

4. Differentiate between strong and weak acids.

Strong acids: Acid which is fully ionized in water is called a strong acid. e.g. H_2SO_4 , HNO_3 etc.

Weak acids: Acid which is partially ionized is called a weak acid. e.g. CH_3COOH

5. Differentiate between strong and weak bases (or alkalis).

Strong base: Base which is fully ionized in water is called a strong base. e.g. NaOH , KOH etc.

Weak base: Base which is partially ionized is called a weak base. e.g. NH_4OH , $\text{Al}(\text{OH})_3$

6. What is the difference between strong acid and concentrated acid?

Strong acid: Acid which is fully ionized in water is called a strong acid. e.g. H_2SO_4 , HNO_3 etc.

Concentrated acid: Acid which contains a large amount of solute in a specific volume of solution.

7. Define Acid and Base according to Arrhenius Concept. OR Define Arrhenius Acid/Base. Give example.

Acid is a substance which dissociates in water to give hydrogen ions. e.g. $\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}^+ + \text{Cl}^-$

Base is a substance which dissociates in water to give hydroxyl ions. e.g. $\text{NaOH} + \text{H}_2\text{O} \rightleftharpoons \text{Na}^+ + \text{OH}^-$

Common examples of alkalis

Sodium hydroxide

Potassium hydroxide

Calcium hydroxide

8. Define Acid and Base according to Bronsted-Lowry concept.

Acid is a substance that can donate a proton.

Base is a substance that can accept a proton.

Example: $\text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{Cl}^-$ (HCl acts as an acid while H_2O acts as a base)

Basic oxides

Calcium oxide

Zinc oxide

Magnesium oxide

9. Define acid rain. Write its harmful effects.

Acid rain: The rainwater which has a pH of 4.2 to 4.4 is called acid rain.

Formation: The gases SO_2 and NO_2 present in air mix with water and convert into acid droplets.

Effects: 1. It makes soil acidic. 2. It makes water bodies too acidic for aquatic animals. 3. It damages buildings, statues, and other structures made of stone and metals.

10. Write the name of gases which cause acid rain.

Acid rain is caused by SO_2 and NO_2 .

11. How are most acids available? In what form and why?

They are available in aqueous solution as they can easily dissolve in water.

12. Differentiate between organic (or natural) acids and inorganic acids (or man-made acids).

Organic acids: Acids obtained from natural sources are called organic acids. e.g. CH_3COOH

Inorganic acids: Acids obtained from minerals are called mineral acids. e.g. HCl , HNO_3 , H_2SO_4

13. All alkalis are bases, but all bases are not alkalis. Why?

An alkali is a base that dissolves in water. Because all bases are not soluble in water, all bases are not alkalis.

14. Define amphoteric substance. Give example.

The substance which behaves as an acid as well as a base is called an amphoteric substance. e.g. Water

15. Define neutralization reaction. Give example.

The reaction between an acid and a base is called a neutralization reaction. e.g. $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$

16. Differentiate between conjugate acid and base.

Conjugate acid: A substance formed by accepting a proton from a base.

Conjugate base: A product formed by donating a proton by an acid.

$\text{CH}_3\text{COOH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}_3\text{O}^+$ (Acid, Base, Conjugate Base, Conjugate Acid)

17. Why does Na_2CO_3 behave like a base in water?

Because it produces a strong base and a weak acid when it is dissolved in water.

18. Why are metal oxides basic in nature?

Because they form salt and water when they react with an acid.

19. Write down the importance of oxalic acid. OR Write uses of oxalic acid.

i. It removes stains from clothes. ii. It cleans bathroom surfaces easily.

LONG QUESTIONS

1. Arrhenius concept of acid, base

2. Bronsted concept of acid, base

3. Properties of acid and base

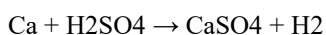
4. Acid rain and its effect

EXERCISE**1. Choose Arrhenius Acids among the following compounds: HF , H_2SO_4 , SO_3 , H_2S , H_2O**

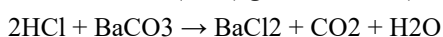
Among the given compounds, the Arrhenius acids are HF , H_2SO_4 , and H_2S because they dissociate in water to produce hydrogen ions (H^+).

2. How does calcium metal react with dilute H_2SO_4 ?

When calcium metal reacts with dilute H_2SO_4 , it produces calcium sulfate (CaSO_4) and releases hydrogen gas (H_2).

**3. Which salt is formed when HCl reacts with BaCO_3 ?**

When HCl reacts with BaCO_3 , the salt formed is barium chloride (BaCl_2). The reaction also produces carbon dioxide (CO_2) gas and water (H_2O).



4. How will you justify that HSO_4^- is a Bronsted-Lowry acid?

HSO_4^- is a Bronsted-Lowry acid as it donates a proton when dissolved in water. Therefore, HSO_4^- is a Bronsted-Lowry acid.

5. Why is HCl not edible although it is present in the stomach and responsible for digestion of food?

HCl is not edible because concentrated acid is corrosive and damages tissues, while in the stomach it's diluted and protected by mucus, safely helping digestion.

Chapter 8: Periodic Table & Periodicity

1. Differentiate between Groups and Periods.

Group: Vertical columns in a periodic table are called groups. There are 18 groups in the periodic table.

Periods: Horizontal rows in a periodic table are called periods. There are 7 periods in the periodic table.

2. Define Modern periodic table.

It is based on the arrangement of elements according to increasing atomic number.

3. Differentiate between Modern Periodic Law and Mendeleev's Periodic Law.

Modern Periodic Law: The properties of elements are periodic functions of their atomic numbers.

Mendeleev's Periodic Law: The properties of elements are periodic functions of their atomic masses.

4. How many elements are present in periods 1–7?

Period No.	Number of elements	Name of the period
1st	2	Short
2nd and 3rd	8	Normal
4th and 5th	18	Long
6th and 7th	32	Very long

Facts

No. of blocks in periodic table: 4 blocks — s, p, d, f blocks
 Why is it difficult to remove an electron from halogens?
 Because their atomic size is small and there is more effective nuclear charge.

5. What are Lanthanides and Actinides?

Lanthanides: The first series starts after lanthanum (La = 57) and is called lanthanides.

Actinides: The second series starts after actinium (Ac = 89) and is called actinides.

6. Define transition elements.

Groups 3 to 12 are called transition elements, and in these elements the inner sub-shells are in the process of completion.

7. Define blocks. And write the name of different blocks.

Elements with similar sub-shell electronic configuration are called blocks. e.g. s, p, d, and f blocks

8. What are s-block, p-block, d-block, and f-block elements?

Block	Description
s-block	Elements of groups 1 and 2 are called s-block elements because in them the s-subshell of the outermost shell is being filled.
p-block	Elements present in groups 13 to 18 are called p-block elements because the p-subshell is filled in these elements.
d-block	The d-block elements lie between the s and p block.
f-block	The f-block elements, in the form of two rows, lie at the bottom of the periodic table.

9. Define metallic character and electron affinity. Give examples.

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Metallic character: It is the tendency of an element to lose an electron and form a cation.

Electron affinity: Energy released when an electron is added to the valence shell of a gaseous atom.

Example: The electron affinity of fluorine is -328 KJmol^{-1}

Memory aid

I.E.E (Inc. in Period) — I for Ionization Energy, E for Electron Affinity, E for Electronegativity

C.A.M (Inc. in Group) — C for Chemical reactivity, A for Atomic radius, M for Metallic character

10. Define Electronegativity. Give example.

The ability of an atom to attract the shared pair of electrons towards itself in a bond. Example: Electronegativity of fluorine is 4

11. Ionization Energy. Write its unit, also give example.

Energy required to remove the most loosely bound electron from the valence shell of a gaseous atom. Unit: kilojoule per mole.

Example: The first ionization energy of sodium atom is 496 KJmol^{-1} .

12. How does ionization energy vary with atomic size?

When atomic radius decreases, ionization energy increases.

13. Write the trend (or variation) of Electronegativity or Ionization energy or Electron affinity in groups and periods.

Trend in Period: It increases from left to right in a period, because nuclear charge increases, atomic radius decreases and attraction between nucleus and electrons increases.

Trend in Group: It decreases from top to bottom in a group, because new electron shells are added, atomic radius increases and shielding effect reduces.

14. Write the trend of Chemical reactivity or atomic radius or metallic character in groups and periods.

Trend in Period: It decreases from left to right in a period, because nuclear charge increases and attraction between nucleus and electrons increases.

Trend in Group: It increases from top to bottom in a group, because new electron shells are added and atomic radius increases with reduced shielding effect.

15. Define Atomic Radius. Give example. Also write its unit.

The half of the distance between the nuclei of two bonded atoms. e.g. Radius of carbon atom is 77 pm. Unit: Picometer (pm)

16. What is the relationship between group number and charge of ion?

The ion charge of an element is related to its group number: Groups 1–3 form +1 to +3 ions, Groups 15–17 form –3 to –1 ions, and Group 14 rarely forms simple ions.

17. How would you find the atomic radius of an atom?

By using x-ray diffraction or spectroscopy methods, we can measure atomic radius.

18. Why are elements of group 1 electropositive? OR Why are alkali metals electropositive?

Because these elements have a strong tendency to lose their valence electrons to form cations.

19. Why are elements of group 17 electronegative?

Because these elements can gain one electron easily to form an anion.

20. Define Periodicity of properties or periodic function.

It is the repetition of properties after a regular interval. e.g. chemical properties.

21. Why does periodicity in the properties of elements occur? OR Why do elements show periodicity of properties?

It occurs due to the recurrence of similar electronic configuration in the outermost shells.

22. Why do elements present in the same group show the same chemical properties? OR Why do all members of a group show same chemical properties?

Because they have the same number of electrons in their valence shells.

23. In a group, the elements have the same number of electrons in the outermost shell. Why is it so?

Because they have the same electronic configuration in their outermost shell. This gives them similar chemical properties.

24. Why can atomic size not be defined exactly?

Because the electronic cloud of an atom has no exact limit.

25. Why does atomic radius decrease along the period or from left to right in the period?

Because nuclear charge increases and electrons are drawn closer to each other.

26. Why does atomic size decrease along the period?

Because effective nuclear charge increases.

27. Why does the size of a period increase (along the group) as we move down the periodic table?

As we move down the periodic table, new shells are added, which can accommodate more electrons.

28. Which element has the least value of ionization energy and which element has the highest value?

Cesium has the least ionization energy value; Fluorine has the highest ionization energy trend among the noted elements (fluorine also has the highest electronegativity).

29. Suppose a new element is discovered. Where would you like to accommodate this element in the periodic table?

To accommodate a new element in the periodic table, we would need to know its atomic number, electronic configuration, and properties.

30. What is the first element of the periodic table? Will it lose an electron or gain it?

The first element is Hydrogen (H). It can both lose and gain an electron — it tends to lose an electron to form H^+ or gain an electron to form H^- . However, typically, it loses an electron.

31. Atomic radii of boron and aluminum are 88 pm and 125 pm respectively. Which element is expected to lose electrons easily?

An element with a greater atomic radius will lose electrons more easily. So, aluminum will lose electrons more easily because it has a larger atomic radius (125 pm).

32. Why is aluminum/silicon less reactive than magnesium/sodium?

Aluminum/silicon are less reactive due to a more stable electronic configuration of electrons in the same shell.

LONG QUESTIONS

1. Similarities in chemical properties of elements in same group

2. Atomic Radius, Ionization energy, Electron affinity, Electronegativity

3. Metal character and reactivity

EXERCISE**1. Why was atomic number chosen to arrange the elements in the periodic table?**

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Because arranging elements by atomic number ensures that elements with similar properties are grouped together.

2. What is the significance of the word 'periodic'?

The word periodic signifies that the properties of elements repeat at regular intervals when arranged by increasing atomic number. This repeating pattern helps in predicting the behavior of elements in the periodic table.

3. Why does the size of an atom increase as we move down the periodic table?

The size of elements increases as we move down a group. This is because with each step down a group, a new electron shell is added, increasing the atomic radius.

4. In a group, the elements have the same number of electrons in the outermost shell. Why is it so?

In a group, elements have the same number of electrons in the outermost shell because they have a similar valence electron configuration, which gives them similar chemical properties. This is due to their positions in the same vertical column of the periodic table.

5. Do you expect calcium to be more reactive than sodium? Give the reason for your answer.

No, calcium is less reactive than sodium because sodium loses its electron more easily.

6. Which element has the maximum atomic radius and which element has the minimum atomic radius in the third period?

In the 3rd period, sodium has the maximum atomic radius, and chlorine (Cl) has the minimum atomic radius.

7. Why are the most electronegative elements present in the sixth and seventh groups?

Because elements of groups 16 and 17 have a strong tendency to attract electrons to complete their octets.

8. The first ionization energy value of magnesium is less than the second one. Give reason.

The first ionization energy of magnesium is less than the second one because removing the first electron requires less energy, whereas removing the second electron requires more energy.

9. Is it possible for two metals or two non-metals to form an ionic bond?

No. Ionic bonds are formed between a metal and a non-metal. Two metals or two non-metals cannot easily transfer electrons this way.

10. Which element has the least value of ionization energy and which has the highest value of electronegativity?

Least ionization: Cesium has the least ionization energy value (i.e. 377 KJ/mol)

Highest Electronegativity: Fluorine has the highest electronegativity value (i.e. 4)

Chapter 9: Group Properties & Elements

1. Differentiate between metals and non-metals.

Metals: The elements which can lose electrons to form positive ions. e.g. Iron, Gold etc.

Non-Metals: The elements which can gain electrons to form negative ions. e.g. Oxygen, Carbon etc.

2. Write properties of metals and non-metals. OR Compare the properties of metals and non-metals.

Metals	Non-metals
They are malleable and ductile.	They are brittle.
They are lustrous.	They are dull.
They have high melting and boiling points.	They have low melting and boiling points.
They are good conductors of heat and electricity.	They are bad conductors of heat and electricity.

3. Why are metals generally tough and strong whereas non-metals are neither tough nor strong?

Metals are tough and strong due to metallic bonding, while non-metals are weak due to weaker intermolecular forces.

4. Why are elements of group 1 called alkali metals?

Because they have 1 electron in their valence shell.

5. How do alkali metals react with oxygen?

Alkali metals react with oxygen to produce oxides. e.g. $4K + O_2 \rightarrow 2K_2O$

6. Why do group 1 metals become softer and their melting point decreases down the group?

Because increase in atomic size reduces melting point and hardness.

7. Why does the density of group 1 metals increase down the group?

Because the increase in mass of the elements is greater than the increase in volume.

8. Why are the elements of group 17 called halogens?

Because they produce salt products when they react with metals.

9. Define transition elements. Give examples.

Elements from group 3 to 12 are called transition elements. e.g. Nickel, Cobalt etc.

10. Write down the properties of transition elements.

i. These are metals. ii. They have high densities. iii. They have high melting and boiling points. iv. They are malleable and ductile.

11. Define noble gases. Give examples.

The elements of group 18 are called noble gases. e.g. Helium, Neon etc.

12. Why do noble gases show very little reactivity? OR Why are noble gases non-reactive?

Because their valence shells are complete, so they show very little reactivity.

13. Which noble gas should have the lowest boiling point and why?

Helium. Because its atomic size is smallest, so intermolecular forces become weakest.

3 elements which exist as liquids

Mercury
Bromine

Gallium

14. Why do metals have high melting and boiling points?

Because they have strong metallic bonds, so they have high melting and boiling points.

15. Why can't non-metals be polished?

Because they exist in powder or gaseous form.

16. Define catalytic converter. Give examples. OR Explain the role of a catalytic converter in an automobile.

A device used in the exhaust of vehicles, converts harmful gases into gases which do not pollute the air. e.g. Platinum, Palladium etc.

17. V₂O₅ catalyst is preferred over platinum in the oxidation of sulphur dioxide. Give reason.

Because V₂O₅, compared to platinum, is less expensive, more efficient and remains stable during the oxidation process.

18. Why are almost all metals solid while non-metals exist as gases and solids?

Metals have strong metallic bonds while non-metals have weaker bonds.

19. Compare the reactions of alkali metals with chlorine.

It reacts strongly with chlorine to form metal chlorides. e.g. $2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}$

20. Compare the reactivity of chlorine and bromine as reducing/oxidizing agents.

Chlorine is a stronger oxidizing agent than bromine — it gains electrons more easily.

21. Which element is the most reactive and which is the least reactive among halogens? Give two reasons.

Most reactive: Fluorine — highly electronegative and smallest atomic size.

Least reactive: Iodine — less electronegative and largest atomic size.

22. What is meant by malleability and ductility?

Malleability: Metals can be hammered into thin sheets — it is called malleability.

Ductility: Metals can be drawn into wires — it is called ductility.

23. Why are transition elements different from normal elements?

Transition elements have special properties due to inner shell electrons.

LONG QUESTIONS

1. Properties of group 1, 17, transition elements, noble gases, physical properties of metals and non-metals

2. Metal/Non-metal

EXERCISE**1. Why does it become easier to cut an alkali metal when we move from top to bottom in a group?**

As we move down Group 1 (alkali metals), the atomic size increases and the metallic bonding weakens due to the increasing distance of the valence electron from the nucleus. This results in a softer metal. Therefore, alkali metals become softer and easier to cut from top (e.g., lithium) to bottom (e.g., cesium) in the group.

2. Predict the reactivity of potassium towards halogens.

Potassium is a highly reactive alkali metal. It reacts vigorously with halogens (like chlorine, bromine, or iodine) to form ionic salts (e.g., KCl, KBr, KI). The reaction is highly exothermic.

3. In the following reaction, chlorine acts as an oxidizing agent. Which is the reducing agent? $\text{Cl}_2 + 2\text{NaBr} \rightarrow 2\text{NaCl} + \text{Br}_2$

Chlorine gains electrons and is reduced, so it acts as the oxidizing agent. Bromide ions (Br^-) lose electrons and are oxidized to bromine (Br_2), making NaBr the reducing agent.

4. Why does iodine exist in the solid state at room temperature?

Iodine has strong Van der Waals forces due to its large molecular size. These forces hold the molecules tightly together. As a result, iodine remains solid at room temperature.

5. How does Ni catalyze the reaction involving hydrogenation of oil?

Nickel (Ni) provides a surface for the reaction, helping hydrogen atoms add to carbon–carbon double bonds in unsaturated oils, converting them into saturated fats.

Chapter 10: Environmental Chemistry

1. Define environmental chemistry. Also write its purpose/main objective of environmental education.

The branch of chemistry which deals with the study of chemicals and other pollutants in the environment.

Purpose: Environmental education teaches people to protect and care for nature.

2. How is the excessive use of insecticides and pesticides harmful for birds?

Their excessive use can harm birds' health and ecosystems.

Some gases that pollute the atmosphere

Coal

Petroleum

Natural gas

3. Define atmosphere. Write names of its major constituents.

Earth is covered with a blanket of air called the atmosphere.

Constituents: Nitrogen, Oxygen, Argon

4. Define air pollutants. Also write their main sources.

Any substance in air which has a bad effect on human health and quality of life.

Sources: i. Burning of fossil fuels ii. Volcanic eruption iii. Population growth iv. Urbanization

5. How is acid rain formed?

Acid rain is formed when oxides of nitrogen and sulphur mix with moisture present in the atmosphere.

6. How does acid rain affect forests?

It damages tree leaves, makes soil more acidic and harms plant growth.

7. How can farmers prevent acid rain damage?

i. Monitoring soil condition. ii. Adding CaO to neutralize acid.

8. What is the effect of acid rain on human health?

Acid rain can indirectly affect human health by contaminating drinking water and food sources.

9. Write down the harmful effects of ozone. OR How can ozone affect human health?

It can cause chest pain, coughing, throat irritation etc.

10. Why do only some pollutant gases present in the atmosphere cause greenhouse effect while others do not?

Greenhouse gases (like CO₂) cause the greenhouse effect because these gases trap infrared radiation, while others (like O₂, N₂) don't.

11. Write the name of air pollutants, also write their harmful effects.

Pollutants	Harmful effect
CO ₂	i. It increases global warming. ii. Extreme weather changes. iii. Rise of sea level.
CO	i. It causes suffocation and death. ii. It stops red blood cells from carrying oxygen.
Hydrocarbons	i. It causes global warming. ii. It causes lung diseases and coughing.

12. Define smog. Also write its cause/main reason.

The mixing of smoke in atmosphere and fog is called smog.

Main Reason: i. Crop burning ii. Vehicle emission iii. Industrial pollution

Quick facts

Gas released when a carbonated drink is opened: CO₂
Gas produced by rice plants / given out on rotting garbage:
Methane (CH₄)

13. Write down the harmful effects of smog. OR Write the name of diseases caused by smog.

It causes allergies, asthma and lung infection in humans.

14. Write the name of main factors for environmental pollution.

i. Transportation ii. Urbanization iii. Industrialization

15. Define Greenhouse effect. Also define Global warming.

Greenhouse effect: It is the progressive warming up of the earth's surface due to the blanketing effect of man-made CO₂, methane etc. in the atmosphere.

Global warming: Global warming is the gradual increase in Earth's average temperature due to excess greenhouse gases. It is mainly caused by human activities like burning fossil fuels and deforestation.

16. Write the sources of greenhouse gases. OR Name any three major sources responsible for greenhouse effect.

i. Methane ii. CO₂ iii. Burning of fossil fuels

17. Write the adverse/harmful effects of global warming.

i. Rise in sea level ii. Increase in rainfall iii. Hot summers and winters

18. Write the name of main factors of polluted air in big cities.

i. Burning fossil fuels ii. Automobiles iii. Aeroplanes iv. Heavy machines

19. Define air quality index (AQI). Write its value when air quality is good.

A rating system which shows how bad the atmosphere is around us.

AQI under 50 is expected to be good quality.

Facts

Gas which increases the risk of lung cancer: Radon
Some renewable energy sources: Wind energy, Solar energy, Biomass

20. How to avoid harmful effects of air pollution? OR Write the steps that should be taken when air quality is bad.

i. Reduce your outdoor time. ii. Wear a face mask. iii. Install a carbon monoxide detector.

21. In what way does sulfur present in fossil fuels become dangerous?

When fossil fuels are burned, sulfur is released as SO₂, which causes acid rain.

22. How can you reduce the emission of CO present in the gases emitted by burning of fuel in automobile engines?

i. Use catalytic converters. ii. Use CNG or electrical vehicles. iii. Maintain vehicles regularly.

23. Define flue gas.

The mixture of waste gases produced when fossil fuels are burnt in power plants.

24. Percentage of CO₂ in air is 0.04%. Then how does it become harmful for ecosystem?

A small increase in CO₂ causes significant global warming.

LONG QUESTIONS

1. Pollutants and harmful effects	2. Acid rain with effects
3. Global warming with sources	3. Effects of climate change
4. Strategies to reduce environmental issues	5. Avoid harmful effects of air pollution

EXERCISE**1. What is the main objective of environmental education?**

The main aim is to help people understand the environment and learn how to take care of it. It also teaches us how to solve problems that harm nature.

2. How is particulate matter released into the atmosphere?

Particulate matter is released into the atmosphere through burning fossil fuels, dust from construction, wildfires, and industrial activities. These tiny particles can come from smoke, dust, and chemicals in the air.

3. Which gas is more poisonous, CO₂ or CO?

Carbon monoxide (CO) is more poisonous than carbon dioxide (CO₂). CO binds with hemoglobin in the blood more effectively than oxygen does, reducing the blood's ability to carry oxygen, which can lead to serious health effects or even death.

4. How does acid rain affect forests?

Acid rain damages forests by washing away important nutrients from the soil and harming leaves and bark. This weakens trees, making them more likely to get sick or die.

5. In what way does sulphur present in fossil fuels become dangerous?

When fossil fuels containing sulphur are burned, they release sulphur dioxide (SO₂) gas. This gas causes acid rain, which harms plants, animals, and buildings, and can also cause breathing problems in humans.

6. Name any three major sources responsible for the greenhouse effect.

i. Methane ii. CO₂ iii. Burning of fossil fuels

7. How is wind energy useful for us?

Wind energy is a clean, renewable source of power that helps reduce pollution and greenhouse gas emissions. It provides sustainable electricity without harming the environment.

Chapter 11: Hydrocarbons

1. Differentiate between organic and inorganic compounds.

Organic compounds: The compounds which contain carbon and hydrogen atoms.

Inorganic compounds: The compounds which do not contain carbon and hydrogen atoms.

Facts

Simplest hydrocarbon (parent alkane): Methane

Some organic compounds: Protein, Enzyme, Lipids

Some inorganic compounds: Carbonates, Carbides, Cyanides

2. How are organic compounds useful for us? OR Describe the importance of organic compounds in daily life.

Because they are used as fuels, medicines and plastics etc.

3. Write the name of five organic compounds which exist naturally. OR Give some examples of organic compounds.

i. Methane ii. Glucose iii. Ethanol iv. Acetic acid v. Citric acid

4. Why are organic compounds not soluble in water?

Because water is polar and organic compounds are non-polar.

5. How do you compare the melting and boiling points of inorganic and organic compounds?

Melting and boiling points of organic compounds are lower than inorganic compounds due to weaker intermolecular forces.

6. Define Hydrocarbons. Also write its types.

The organic compounds which contain carbon and hydrogen atoms only are called hydrocarbons. e.g. Methane, Ethane etc.

Types: i. Saturated hydrocarbons ii. Unsaturated hydrocarbons

Some organic biomolecules

Protein

Vitamins

Lipids

7. How is naphtha fraction decomposed to give lower hydrocarbons?

By cracking, large hydrocarbons are broken down into smaller ones using heat or a catalyst.

8. Define structural formula. Also write structural formula of methane.

The formula which shows the arrangement of molecules is called structural formula.

9. Define catenation.

The ability of a carbon atom to link with other carbon atoms to form rings and chains.

10. Define alkanes/saturated hydrocarbons. Give examples. Write its general formula.

The hydrocarbons with only carbon-carbon and carbon-hydrogen single bonds. e.g. Methane

General formula: C_nH_{2n+2}

Popular medicines which are organic compounds

Aspirin

Penicillin

Vitamin C

Paracetamol

11. Write uses of methane.

i. It is used as fuel. ii. It is used to make carbon black.

12. Why are alkanes called saturated hydrocarbons?

Alkanes are saturated because they only contain single covalent bonds.

13. Why are alkanes called 'paraffins' or the least reactive?

Because they are saturated and less reactive, so they are called paraffins.

14. Why do alkanes show little reactivity towards other reagents?

Because they are non-polar and have strong C-C and C-H bonds, so they are less reactive.

15. How do the melting and boiling points of alkanes change when we move from lower members to higher members?

The melting and boiling points increase as the molecular weight increases.

16. Reactions of alkanes with chlorine take place in the presence of sunlight. What is the role of sunlight in this reaction?

Sunlight plays an important role in the reaction of alkanes with chlorine by providing the required energy to start the reaction.

17. Define photochemical reactions.

The reaction which takes place in the presence of light.

18. Write the name of products obtained when natural gas is oxidized under controlled conditions.

i. Water ii. Heat energy iii. CO₂

19. 'Neem' is a common tree grown throughout our country. Comment on the medicinal benefits of the 'Neem' tree.

It is used to treat skin conditions, boost immunity etc.

20. How are combustion reactions of alkanes useful to us?

It provides energy for cooking, heating and running motor cars.

21. Why is carbon so important as an element that the whole branch of chemistry is based on it?

Because carbon forms many molecules, essential for life and chemistry.

22. Why does a mixture of natural gas and air explode?

The mixture explodes due to rapid combustion when an ignition source, like a spark or flame, is introduced.

23. How do you compare the boiling points of n-butane with that of iso-butane?

n-Butane has a higher boiling point (i.e. -0.5°C) than isobutane (i.e. 11.7°C).

24. What is the reason for explosion of natural gas when a gas valve is kept open in the kitchen? How to avoid it?

Gas leaks mix with air and ignite, causing explosion.

Prevention: i. Turn off the gas valve after use. ii. Ensure proper ventilation.

LONG QUESTIONS

1. Preparation of alkane	
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EXERCISE

Visit Our Website: Ahsa.Pk for more notes, mcqs, guess papers, schemes, and more

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1. Difference between organic and inorganic compounds.

Organic compound: The compound which contains carbon and hydrogen atoms only.

Inorganic compound: The compound which does not contain carbon and hydrogen atoms.

2. Why are organic compounds found in large numbers?

Organic compounds are found in large numbers because carbon atoms link together creating different molecules easily. Carbon makes 4 bonds because it has 4 electrons in its valence shell.

3. Name the products which are obtained when natural gas is oxidized under controlled conditions.

Carbon monoxide (CO), Carbon dioxide (CO₂), Water (H₂O), Hydrogen (H₂)

4. How is naphtha fraction decomposed to give lower hydrocarbons?

Naphtha consists of a mixture of liquid hydrocarbons. It is heated at around 500°C in the presence of a catalyst called zeolite to give hydrocarbons which have five to ten carbon atoms.

5. Write down molecular formula, structural formula and condensed structural formula for iso-butane.

Molecular formula: C₄H₁₀

Structural formula: CH₃–C(CH₃)H–CH₃ (a central carbon bonded to three CH₃ groups and one H)

Condensed formula: (CH₃)₂CHCH₃

6. How are organic compounds useful for us?

Organic compounds are used in fuels, medicines, plastics, textiles and food, supporting essential needs and industries.

7. Give IUPAC name of the given branched-chain compound (CH₃–CH(CH₃)–CH₂–CH₃).

The IUPAC name is 2-methylbutane

8. How do the melting and boiling points of alkanes change when we move from lower members to higher members?

Melting and boiling points of alkanes increase when we move from lower members to higher members due to stronger intermolecular forces and increased molecular weight.

Chapter 12: Empirical Data Collection & Analysis

1. Define science.

It is the systematic study of this world through observation and experimentation.

2. Write the names of five physical quantities used in chemistry.

i. Length ii. Mass iii. Time iv. Temperature v. Amount of substance

3. Write the name of some base units used in chemistry.

i. Metre ii. Kilogram iii. Second iv. Kelvin v. Mole

4. Define derived units. Give examples.

The quantities derived from base units are called derived units. e.g. m^2 , m^3 etc.

5. Why is the Celsius scale used instead of the Kelvin scale? OR Why is Celsius scale most often used to measure temperature in chemistry rather than Kelvin scale?

Because: i. Celsius scale is easier to use. ii. It is easy to convert temperature on Celsius scale into Kelvin scale.

6. Why is the unit cubic centimeter (cm^3) used in chemistry rather than (m^3)? OR Why do we prefer to use smaller units of mass and volume in chemistry?

Because: i. It is precise. ii. It is easy to measure.

7. Define error. Also write its types.

The difference between the measured value and actual value.

Types: i. Systematic error ii. Random error

8. Why does error occur?

An error occurs due to: i. Limitation of measuring instruments. ii. Skill of the student taking the measurements.

9. Differentiate between systematic error and random error. OR Define systematic error and random error.

Systematic error: A consistent and repeatable error caused by faulty equipment or biased methods.

Random error: The error committed by students during measurements is called random error.

10. How can you avoid systematic error in your measurement?

i. Follow rules. ii. Train properly. iii. Use proper methods.

11. Differentiate between accuracy and precision.

Accuracy: It measures how close results are to the true or known value.

Precision: It is the closeness of two or more measurements to each other.

12. Write the relationship between accuracy and precision.

Precision is independent of accuracy. A student may be accurate but not precise and vice versa.

13. Write 2 benefits scientists get by using SI units.

i. Makes calculations easy. ii. Ensures global consistency.

14. How does taking measurements in SI units ensure safety and reliability?

SI units ensure accuracy and safety consistently every time.

15. Explain the three units derived from the basic SI units.

1. Newton from force 2. Pascal from pressure 3. Joule from energy

16. Compare the units in the SI system with those in the MKS system.

The SI system includes all MKS units (meter, kilogram, second). It also adds more units for quantities like temperature (kelvin) and electric current (ampere).

17. What are our indigenous units of measurement for mass, volume and length?

i. Mass: Seer ii. Volume: Litre iii. Length: Yard

18. Can a student be both inaccurate and imprecise in his measurement?

Yes, when methods and tools are used incorrectly.

19. What difficulties would we expect to encounter if we use different units of measurement in daily life?

i. Confusion ii. Incorrect calculations iii. Errors

LONG QUESTIONS

1. SI units in chemistry

2. Tools and techniques to manage accuracy and precision

EXERCISE**1. What is consistency of results?**

Consistency of results means getting the same or very similar results when an experiment or measurement is repeated under the same conditions.

2. Why are SI units user friendly?

SI units are user friendly because they are standardized, easy to use, and based on the decimal system, which makes calculations simple.

3. Define systematic error and random error.

Systematic error: A consistent and repeatable error caused by faulty equipment or biased methods.

Random error: The error committed by students during measurements is called random error.

4. What is the reason behind a random error?

Random errors happen because of uncertain mistakes or changes.

5. Does systematic error affect accuracy?

Yes, systematic errors affect accuracy because they cause measurements to be consistently off from the true value.

6. Which other systems of measurement are used apart from SI units?

Other measurement systems include the Imperial system (used in the UK and US) and the CGS system (centimeter-gram-second).

7. Define metre.

A metre is the SI unit of length, defined as the distance light travels in a vacuum in $1/299,792,458$ seconds.

8. Mention two benefits scientists get by using SI units.

- i. SI units provide a universal standard, allowing easy communication and comparison of results worldwide.
- ii. They simplify calculations because they are based on the decimal system.

Chapter 13: Laboratory & Practical Skills

1. Define chemistry laboratory.

A place where a student is trained to observe the physical and chemical characteristics of substances by following definite procedures.

2. Why are signs and symbols posted on lab and chemical bottles? OR Why are hazard signs used?

To warn of danger and provide safety information, signs and symbols are posted.

3. Put forward at least two suggestions to improve safety in the lab.

Wear protective gear. Follow safety rules carefully.

4. What types of damages can reactive chemicals cause?

i. Fire ii. Explosion iii. Toxic gas release

5. How will you handle an emergency situation caused by fire due to short circuiting?

i. Turn off the main supply. ii. Call emergency services. iii. Activate fire alarm.

6. Why are flammable liquids not stored in the refrigerator?

Because they can cause an explosion in the refrigerator, so they should not be refrigerated.

7. Why can some sudden shock cause some chemicals to explode?

Some sudden shock causes chemicals to react quickly and explode.

8. Write some names of flammable, explosive, corrosive and toxic chemicals.

Category	Examples
Flammable chemicals	Petrol, Acetone, Methanol
Explosive Chemicals	Nitric acid, H ₂ SO ₄ , NaOH
Corrosive Chemicals	HCl, H ₂ SO ₄ , NaOH
Toxic Chemicals	Cl ₂ , CO, NH ₃

9. Define corrosive chemicals. Write the name of some corrosive chemicals.

The chemicals which attack living tissues when they come in contact with them. e.g. HCl, H₂SO₄, NaOH etc.

10. Define toxic chemicals. Give examples.

A poisonous material capable of causing health problems is called a toxic chemical. e.g. Chlorine gas, carbon monoxide etc.

11. How do corrosive chemicals affect us?

These chemicals affect skin, eyes and the respiratory system.

12. Explain hazards or dangers due to explosive and toxic chemicals.

Explosions can cause injuries, damage property, and generate toxic fumes.

Toxic chemicals can cause problems through inhalation, skin contact, or ingestion.

13. What type of safety precautions are adopted to avoid damage due to explosive chemicals?

Use small quantities. Store in a cool and dry place. Avoid friction and heat near explosive chemicals.

14. Under what conditions (or circumstances) are explosive chemicals likely to explode?

These chemicals explode when subjected to heat, shock or friction.

15. What type of reactions should be carried out in a fume cupboard?

Reactions like volatile, toxic hazardous chemicals should be carried out in a fume cupboard.

16. Describe the effect of spreading toxic gas in the lab.

Spreading toxic gas in a lab can cause effects ranging from mild irritation to severe injury or death. e.g. Distress, damage to tissues etc. and even death.

17. Write the name of chemicals which can cause problems for the safety of workers.

Paints, Solvents, Compressed gas cylinders

18. What precautionary measures should be adopted in a chemistry laboratory? OR Write down five common safety instructions used to avoid all types of hazards.

Don't bring any food items.

- Never eat or drink in the laboratory.
- Never taste or smell any chemical.
- Dispose of waste properly.

Familiarize yourself with how to use fire extinguishers etc.

LONG QUESTIONS

1. Chemical hazards in the laboratory with precautions

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EXERCISE**1. Name some corrosive chemicals.**

HCl, H₂SO₄, NaOH

2. What type of safety precautions are adopted to avoid damage due to explosive chemicals?

Ensure enough ventilation. Handle containers with care. Wear a lab coat, chemical-resistant gloves, and eye protection.

3. What type of damages can reactive chemicals cause?

Reactive chemicals can cause explosions, fires, and toxic releases, and severe burns due to violent reactions. They also harm the environment by polluting air, water, and soil.

4. Indicate two such safety instructions which are required to avoid radiation.

- Wear protective gear.
- Follow safety rules to avoid radiation.

5. Which chemicals can cause suffocation?

- Cl₂
- H₂S
- CO

6. Why are signs and symbols posted on lab and chemical bottles?

Signs and symbols are posted on lab and chemical bottles to warn of potential hazards and to ensure safe handling, storage, and disposal of substances. They help prevent accidents by providing clear, standardized safety information.

7. How should fire caused by chemicals be handled?

- Avoid using water.
- Use suitable fire extinguishers.

8. Why are emergency drills important to face emergency situations?

Drills prepare us to react quickly and stay safe in emergency situations.

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