Chemistry Class 9 Notes - Complete Chapter-wise Notes for Punjab 9th Class (2026-27)

9th Class - Punjab Board - Session 2026-27

Original study notes by Zahid Notes for revision. These are summaries, not the official textbook.

Introduction

These chemistry class 9 notes cover the full 9th class Chemistry book chapter by chapter, in simple English, for students of Punjab in the 2026-27 session. Each chapter gives a short summary of the main ideas, the definitions students most need to remember, important short and long questions with concise answers, and practice MCQs with an answer key. The notes are written as original study summaries: they explain the ideas in the student's own study language and do not reproduce the textbook.

A good way to use these chemistry class 9 notes is to read the summary of a chapter first, then cover the answers and try the short questions, and finally test yourself with the MCQs. Chemistry becomes easier when you connect each new idea to something you can picture - for example, a melting ice cube for a change of state, or table salt dissolving in water for a solution.

A note on the book edition: The chapter sequence used here follows the newer Single National Curriculum (SNC) edition of the Punjab 9th class Chemistry book, which third-party book listings describe as having 13 chapters and as being in use in Punjab since the 2025-26 session. The older Punjab edition had 8 chapters with different titles. The official PECTAA notification for the 2026-27 session has NOT been independently verified for this draft, so students should always check the chapter titles against their own issued book.

Chapter 1: States of Matter and Phase Changes

Summary

Everything around us is matter, and matter is commonly found in three states: solid, liquid and gas. In a solid, particles are packed closely and vibrate in fixed positions, which is why a solid keeps its shape and volume. In a liquid, particles are close but can slide past each other, so a liquid keeps its volume but takes the shape of its container. In a gas, particles are far apart and move freely, so a gas fills its container completely.

Changing the temperature or pressure can change the state of a substance. Melting turns a solid into a liquid, freezing turns a liquid into a solid, evaporation turns a liquid into a gas, condensation turns a gas into a liquid, and sublimation turns a solid directly into a gas. Diffusion is the spreading of particles from a region of higher concentration to a region of lower concentration, and it happens fastest in gases because their particles move most freely. Pressure in a gas comes from particles striking the walls of their container, and heating a gas makes its particles move faster.

Key definitions

- **Matter:** Anything that has mass and occupies space.
- **Diffusion:** The spontaneous movement of particles from higher to lower concentration until they are evenly spread.
- **Melting point:** The temperature at which a solid changes into a liquid.
- **Boiling point:** The temperature at which a liquid changes into a gas throughout the liquid, not only at the surface.
- **Evaporation:** The change of a liquid into a gas at the surface, at any temperature below the boiling point.
- **Sublimation:** The direct change of a solid into a gas without passing through the liquid state.

Important short questions

Q1. Why does a gas fill its container completely while a liquid does not?

A: Gas particles are far apart and move freely in all directions, so they spread through the whole container. Liquid particles stay close together, so a liquid keeps its own volume.

Q2. What is the difference between evaporation and boiling?

A: Evaporation happens only at the surface of a liquid and at any temperature, while boiling happens throughout the liquid and only at the boiling point.

Q3. Why does diffusion happen faster in gases than in liquids?

A: Gas particles have more space between them and move faster, so they mix more quickly than liquid particles.

Q4. Give one everyday example of sublimation.

A: Dry ice (solid carbon dioxide) changes directly into gas, and mothballs slowly disappear into the air, without becoming liquid.

Important long questions

Q1. Compare the three states of matter in terms of particle arrangement, particle movement, shape and volume.

A: In solids, particles are tightly packed in fixed positions and only vibrate, so solids have a fixed shape and volume. In liquids, particles are close but can move past one another, so liquids have a fixed volume but no fixed shape. In gases, particles are far apart and move freely and rapidly, so gases have neither a fixed shape nor a fixed volume and fill their container.

Q2. Explain, using particle ideas, what happens when ice is heated until it becomes steam.

A: Heating gives particles more energy. In ice, particles vibrate faster until the forces holding them in fixed positions weaken and the solid melts into water. Further heating makes liquid particles move faster until they escape from the surface and then from the whole liquid, forming steam. Cooling reverses the process: steam condenses to water, and water freezes to ice.

MCQs

1. Which state of matter has a fixed shape and a fixed volume? a) Liquid b) Gas c) Solid d) Plasma - **Answer: c**

2. The process by which a liquid changes into a gas at its surface is called: a) Condensation b) Evaporation c) Freezing d) Sublimation - **Answer: b**

3. Diffusion is fastest in: a) Solids b) Liquids c) Gases d) All at the same rate - **Answer: c**

4. When a gas is heated in a closed container, its pressure increases because: a) Particles become larger b) Particles strike the walls more often and harder c) Particles stop moving d) The container shrinks - **Answer: b**

5. Which of the following is an example of sublimation? a) Ice melting b) Water boiling c) Dry ice turning into gas d) Steam condensing - **Answer: c**

Answer key (Ch. 1): 1-c, 2-b, 3-c, 4-b, 5-c.

Chapter 2: Atomic Structure

Summary

An atom is the smallest unit of an element that can take part in a chemical reaction. Atoms are built from three smaller particles: protons, which carry a positive charge; electrons, which carry a negative charge; and neutrons, which carry no charge. Protons and neutrons sit together in the tiny central nucleus, while electrons move in energy levels, or shells, around the nucleus.

The atomic number of an element is the number of protons in its atom, and it decides which element the atom is. The mass number is the total number of protons and neutrons. Atoms of the same element with different numbers of neutrons are called isotopes; they have the same atomic number but different mass numbers. Electrons fill shells in a regular order, and the arrangement of electrons in shells is called the electronic configuration. Early models of the atom, such as Rutherford's nuclear model and Bohr's model of fixed energy levels, helped scientists explain how atoms are built and why they give out light of particular colours.

Key definitions

- **Atom:** The smallest particle of an element that can exist and take part in a chemical reaction.
- **Atomic number:** The number of protons in the nucleus of an atom.
- **Mass number:** The total number of protons and neutrons in the nucleus of an atom.
- **Isotopes:** Atoms of the same element with the same atomic number but different mass numbers (different numbers of neutrons).
- **Electronic configuration:** The arrangement of electrons in the energy levels (shells) around the nucleus.
- **Shell (energy level):** A region around the nucleus where electrons with a particular amount of energy are found.

Important short questions

Q1. What is the difference between atomic number and mass number?

A: Atomic number is the number of protons only, while mass number is the number of protons plus neutrons.

Q2. Why is an atom electrically neutral overall?

A: An atom has equal numbers of positively charged protons and negatively charged electrons, so their charges cancel out.

Q3. What are isotopes? Give one example.

A: Isotopes are atoms of the same element with the same number of protons but different numbers of neutrons, such as carbon-12 and carbon-14.

Q4. Write the electronic configuration of sodium (atomic number 11).

A: Sodium has 11 electrons arranged as 2 in the first shell, 8 in the second shell and 1 in the third shell (2, 8, 1).

Important long questions

Q1. Describe the structure of an atom, naming the three subatomic particles, their charges and where each is found.

A: An atom has a central nucleus containing positively charged protons and uncharged neutrons. Negatively charged electrons move around the nucleus in shells. The nucleus is very small compared with the whole atom but holds almost all of its mass, because protons and neutrons are much heavier than electrons.

Q2. Explain the main ideas of Bohr's model of the atom and how it improved on Rutherford's model.

A: Rutherford showed that the atom has a small, dense, positively charged nucleus, but his model could not explain why electrons do not fall into the nucleus. Bohr proposed that electrons move only in certain allowed energy levels and do not radiate energy while staying in one level. An electron gives out or absorbs a fixed amount of energy only when it jumps between levels, which explains the line spectra of elements.

MCQs

1. The particle with a negative charge is the: a) Proton b) Neutron c) Electron d) Nucleus - **Answer: c**

2. The atomic number of an element is equal to the number of: a) Neutrons b) Protons c) Protons + neutrons d) Electrons + neutrons - **Answer: b**

3. Isotopes of an element differ in their number of: a) Protons b) Electrons c) Neutrons d) Shells - **Answer: c**

4. The mass number of an atom with 6 protons and 6 neutrons is: a) 6 b) 12 c) 18 d) 0 - **Answer: b**

5. In Bohr's model, electrons are found in: a) Random positions b) Fixed energy levels c) The nucleus d) Only the outermost shell - **Answer: b**

Answer key (Ch. 2): 1-c, 2-b, 3-c, 4-b, 5-b.

Chapter 3: Chemical Bonding

Summary

Atoms join together because a full outer shell is a stable arrangement, and most atoms do not have one on their own. Chemical bonding is how atoms achieve stability by gaining, losing or sharing electrons. When a metal atom loses electrons, it becomes a positively charged ion called a cation; when a non-metal atom gains electrons, it becomes a negatively charged ion called an anion. The strong attraction between opposite charges holds an ionic compound together, as in common salt.

In a covalent bond, two atoms share pairs of electrons so that both reach a stable outer shell. A single bond shares one pair, a double bond shares two pairs, and a triple bond shares three pairs. Some substances have both kinds of bonding, and the type of bonding explains many properties: ionic compounds are usually hard, have high melting points and conduct electricity when molten or dissolved, while simple covalent substances are usually soft, have low melting points and do not conduct electricity. Metallic bonding, in which a "sea" of free electrons holds metal ions together, explains why metals conduct electricity and heat and can be hammered into sheets.

Key definitions

- **Chemical bond:** The force that holds atoms together in a molecule or compound.
- **Ionic bond:** The strong electrostatic attraction between oppositely charged ions, formed by the transfer of electrons.
- **Covalent bond:** A bond formed when two atoms share one or more pairs of electrons.
- **Cation:** A positively charged ion, formed when an atom loses electrons.
- **Anion:** A negatively charged ion, formed when an atom gains electrons.
- **Metallic bond:** The attraction between metal ions and the shared, free-moving electrons around them.

Important short questions

Q1. Why do atoms form chemical bonds?

A: Atoms bond to reach a stable electronic arrangement, usually a full outer shell, which is lower in energy and more stable than the separate atoms.

Q2. What is the difference between a cation and an anion?

A: A cation is a positively charged ion formed by losing electrons, while an anion is a negatively charged ion formed by gaining electrons.

Q3. Why do ionic compounds conduct electricity when dissolved in water but not as solids?

A: In a solid, ions are held in fixed positions and cannot move. When molten or dissolved, the ions are free to move and carry charge.

Q4. What is a double covalent bond? Give an example.

A: A double bond is formed when two atoms share two pairs of electrons, as in an oxygen molecule (O=O).

Important long questions

Q1. Explain how sodium chloride is formed, using electron transfer.

A: A sodium atom has one electron in its outer shell and loses it to become a positively charged sodium ion with a stable outer shell. A chlorine atom has seven outer electrons and gains one to become a negatively charged chloride ion with a full outer shell. The strong attraction between the opposite charges forms the ionic bond in sodium chloride.

Q2. Compare ionic and covalent compounds in terms of melting point, hardness and electrical conductivity, and explain the differences.

A: Ionic compounds have high melting points and are hard because strong forces between ions must be overcome, and they conduct electricity when molten or dissolved because their ions can move. Simple covalent substances have low melting points and are soft because only weak forces between molecules must be overcome, and they do not conduct electricity because they have no free ions or electrons.

MCQs

1. An ionic bond is formed by the: a) Sharing of electrons b) Transfer of electrons c) Sharing of protons d) Transfer of neutrons - **Answer: b**

2. A positively charged ion is called a: a) Anion b) Cation c) Molecule d) Radical - **Answer: b**

3. In a water molecule, the atoms are held together by: a) Ionic bonds b) Metallic bonds c) Covalent bonds d) No bonds - **Answer: c**

4. Which substance is most likely to conduct electricity when dissolved in water? a) Sugar b) Wax c) Salt d) Oil - **Answer: c**

5. Metals conduct electricity because of: a) Fixed ions b) Free-moving electrons c) Covalent bonds d) Neutrons - **Answer: b**

Answer key (Ch. 3): 1-b, 2-b, 3-c, 4-c, 5-b.

Chapter 4: Stoichiometry

Summary

Stoichiometry is the study of the quantities of reactants and products in a chemical reaction, using balanced equations. A balanced chemical equation shows the same number of atoms of each element on both sides, because matter is neither created nor destroyed in a reaction. The relative atomic mass of an element compares the mass of its atoms with a standard, and the relative molecular mass is the sum of the relative atomic masses in a molecule.

The mole is the chemist's counting unit: one mole of any substance contains the same number of particles, known as Avogadro's number. The molar mass of a substance is the mass of one mole of it, and it lets us convert between mass and number of moles. An empirical formula shows the simplest whole-number ratio of atoms in a compound, while a molecular formula shows the actual number of atoms in one molecule. From a balanced equation, we can work out how much product forms from a given mass of reactant, and which reactant runs out first is called the limiting reactant.

Key definitions

- **Stoichiometry:** The calculation of the quantities of reactants and products in a chemical reaction from a balanced equation.
- **Mole:** The amount of substance containing as many particles as there are atoms in a standard sample of carbon-12 (Avogadro's number of particles).
- **Avogadro's number:** The number of particles in one mole of a substance (approximately 6.02×10^{23}).
- **Molar mass:** The mass of one mole of a substance, usually expressed in grams per mole.
- **Empirical formula:** The simplest whole-number ratio of atoms of each element in a compound.
- **Molecular formula:** The actual number of atoms of each element in one molecule of a compound.
- **Limiting reactant:** The reactant that is used up first in a reaction and so limits the amount of product formed.

Important short questions

Q1. Why must a chemical equation be balanced?

A: Because atoms are neither created nor destroyed in a reaction, so the number of atoms of each element must be the same on both sides of the equation.

Q2. What is the difference between an empirical formula and a molecular formula?

A: The empirical formula gives the simplest ratio of atoms, while the molecular formula gives the actual numbers of atoms in a molecule. For example, the molecular formula of glucose is $C_6H_{12}O_6$, while its empirical formula is CH_2O.

Q3. Define molar mass. What is the molar mass of water (H_2O)?

A: Molar mass is the mass of one mole of a substance. For water: H = 1 × 2 = 2, O = 16, so the molar mass is 18 g/mol.

Q4. What is a limiting reactant?

A: It is the reactant that runs out first in a chemical reaction, stopping the reaction and limiting how much product can form.

Important long questions

Q1. Calculate the number of moles in 20 g of calcium carbonate ($CaCO_3$). (Relative atomic masses: Ca = 40, C = 12, O = 16.)

A: Molar mass of $CaCO_3$ = 40 + 12 + (16 × 3) = 100 g/mol. Number of moles = mass ÷ molar mass = 20 ÷ 100 = 0.2 mol.

Q2. Explain the steps for finding the empirical formula of a compound from percentage composition data.

A: First, divide the percentage (or mass) of each element by its relative atomic mass to get the moles of each element. Next, divide all the mole values by the smallest one to get a ratio. Finally, if needed, multiply to get whole numbers, which give the empirical formula.

MCQs

1. One mole of any substance contains: a) 6.02×10^{23} particles b) 6.02×10^{22} particles c) 12 particles d) 100 particles - **Answer: a**

2. The molar mass of carbon dioxide (CO_2) is: (C = 12, O = 16) a) 28 g/mol b) 32 g/mol c) 44 g/mol d) 60 g/mol - **Answer: c**

3. The empirical formula of glucose ($C_6H_{12}O_6$) is: a) $C_6H_{12}O_6$ b) CH_2O c) $C_2H_4O_2$ d) CHO - **Answer: b**

4. In a balanced equation, the numbers in front of formulas are called: a) Subscripts b) Coefficients c) Indices d) Charges - **Answer: b**

5. The reactant that is completely used up first is the: a) Excess reactant b) Limiting reactant c) Catalyst d) Product - **Answer: b**

Answer key (Ch. 4): 1-a, 2-c, 3-b, 4-b, 5-b.

Chapter 5: Energetics

Summary

Energetics is the study of energy changes in chemical reactions. Every chemical bond stores chemical energy, and breaking bonds takes in energy while forming new bonds gives out energy. In an exothermic reaction, more energy is given out when new bonds form than is taken in to break old bonds, so the surroundings warm up - burning fuels and respiration are examples. In an endothermic reaction, more energy is taken in than is given out, so the surroundings cool down - examples include photosynthesis and the dissolving of some salts in water.

The energy change of a reaction can be shown on an energy-level diagram, which compares the energy of the reactants with the energy of the products. The difference between them is the enthalpy change. A catalyst speeds up a reaction without being used up, by providing a pathway with a lower activation energy - the minimum energy colliding particles need before they can react. Activation energy explains why some reactions need a spark or heating to start, even though they give out energy overall.

Key definitions

- **Exothermic reaction:** A reaction that gives out energy to the surroundings, usually as heat.
- **Endothermic reaction:** A reaction that takes in energy from the surroundings.
- **Enthalpy change:** The overall heat energy change of a reaction at constant pressure.
- **Activation energy:** The minimum energy that colliding particles must have for a reaction to happen.
- **Catalyst:** A substance that speeds up a reaction without being used up, by lowering the activation energy.
- **Bond energy:** The energy needed to break one mole of a particular type of bond.

Important short questions

Q1. What is the difference between an exothermic and an endothermic reaction?

A: An exothermic reaction gives out energy and warms the surroundings, while an endothermic reaction takes in energy and cools the surroundings.

Q2. Why does breaking bonds require energy, while forming bonds releases energy?

A: Bonds hold atoms together, so pulling them apart needs an energy input; when new, more stable bonds form, energy is released.

Q3. How does a catalyst speed up a reaction?

A: A catalyst provides an alternative reaction pathway with lower activation energy, so more particles have enough energy to react. The catalyst itself is not used up.

Q4. Give two examples of exothermic reactions in everyday life.

A: Burning wood or gas for cooking, and respiration in the body, both of which release heat energy.

Important long questions

Q1. Using bond breaking and bond forming, explain why a reaction can be exothermic overall.

A: Energy is taken in to break the bonds of the reactants and given out when the bonds of the products form. If the energy released in forming the new bonds is greater than the energy taken in to break the old bonds, the extra energy is released to the surroundings and the reaction is exothermic.

Q2. Draw and explain an energy-level diagram for an exothermic reaction, labelling reactants, products, enthalpy change and activation energy.

A: On the diagram, the reactants start at a higher energy level than the products, because energy has been released overall. The curve rises first to an energy barrier, which is the activation energy that must be supplied to start the reaction. The vertical drop from the reactants' level to the products' level is the enthalpy change, shown as a negative value because energy is given out.

MCQs

1. Which of the following is an exothermic process? a) Melting ice b) Photosynthesis c) Burning natural gas d) Dissolving ammonium nitrate in water - **Answer: c**

2. In an endothermic reaction, the temperature of the surroundings: a) Increases b) Decreases c) Stays the same d) First rises then falls - **Answer: b**

3. Activation energy is the: a) Total energy of the products b) Minimum energy needed for colliding particles to react c) Energy released by a catalyst d) Energy stored in a fuel - **Answer: b**

4. A catalyst works by: a) Increasing the activation energy b) Being used up in the reaction c) Lowering the activation energy d) Cooling the reactants - **Answer: c**

5. When bonds form, energy is: a) Absorbed b) Released c) Destroyed d) Unchanged - **Answer: b**

Answer key (Ch. 5): 1-c, 2-b, 3-b, 4-c, 5-b.

Chapter 6: Equilibria

Summary

Many chemical reactions can go in both directions: the forward reaction turns reactants into products, while the reverse reaction turns products back into reactants. Such a reaction is called reversible. In a closed system, a reversible reaction eventually reaches a state of dynamic equilibrium, where the forward and reverse reactions continue at equal rates, so the amounts of reactants and products stay constant even though the reactions have not stopped.

The position of equilibrium can be shifted by changing conditions. According to Le Chatelier's principle, if a change is made to a system at equilibrium, the system responds in a way that opposes the change. Raising the temperature favours the endothermic direction, increasing pressure favours the side with fewer gas molecules, and adding more of a reactant pushes the equilibrium towards the products. Catalysts speed up both directions equally, so they help equilibrium be reached faster but do not change its position. Equilibrium ideas are used in industry, for example in making ammonia, where temperature, pressure and a catalyst are chosen to balance speed against yield.

Key definitions

- **Reversible reaction:** A reaction that can proceed in both the forward and reverse directions.
- **Dynamic equilibrium:** A state in a closed system where forward and reverse reactions occur at the same rate, so concentrations stay constant.
- **Closed system:** A system in which no substances can enter or leave.
- **Le Chatelier's principle:** If a condition of a system at equilibrium is changed, the system shifts to oppose that change.
- **Yield:** The amount of product obtained from a reaction, often compared with the maximum possible amount.

Important short questions

Q1. What is meant by dynamic equilibrium?

A: It is a state where the forward and reverse reactions continue at equal rates, so the amounts of reactants and products remain constant, even though both reactions are still happening.

Q2. State Le Chatelier's principle.

A: If a change is applied to a system at equilibrium, the system shifts in the direction that opposes the change.

Q3. What happens to the position of equilibrium if the pressure of a gaseous equilibrium is increased?

A: The equilibrium shifts towards the side with fewer gas molecules, because that reduces the pressure.

Q4. Does a catalyst change the position of equilibrium? Explain.

A: No. A catalyst speeds up the forward and reverse reactions equally, so equilibrium is reached faster but its position does not change.

Important long questions

Q1. Explain, using Le Chatelier's principle, how temperature change affects an exothermic reversible reaction.

A: In an exothermic forward reaction, the reverse reaction is endothermic. Raising the temperature favours the endothermic direction, so the equilibrium shifts towards the reactants and the yield of product falls. Lowering the temperature favours the exothermic forward direction and increases the yield, but the reaction becomes slower, so in practice a compromise temperature is used.

Q2. Describe what happens, in terms of reaction rates, when a reversible reaction reaches equilibrium in a closed container.

A: At first, only reactants are present, so the forward reaction is fast and the reverse reaction is slow. As products build up, the forward rate falls and the reverse rate rises. When the two rates become equal, equilibrium is reached: reactants keep turning into products and products keep turning back, but the overall amounts no longer change.

MCQs

1. A reversible reaction is shown by the symbol: a) → b) ⇌ c) = d) ↑ - **Answer: b**

2. At dynamic equilibrium, the concentrations of reactants and products: a) Are zero b) Keep changing c) Stay constant d) Become equal - **Answer: c**

3. Increasing the temperature of an exothermic reversible reaction shifts equilibrium towards the: a) Products b) Reactants c) Neither side d) Catalyst - **Answer: b**

4. Which change does NOT affect the position of equilibrium? a) Temperature b) Pressure c) Adding a catalyst d) Concentration - **Answer: c**

5. Equilibrium can only be reached in a: a) Open system b) Closed system c) Very hot system d) Very cold system - **Answer: b**

Answer key (Ch. 6): 1-b, 2-c, 3-b, 4-c, 5-b.

Chapter 7: Acid Base Chemistry

Summary

Acids are substances that release hydrogen ions (H^+) in water and have a sour taste; common examples are hydrochloric acid, sulphuric acid and citric acid in citrus fruits. Bases are substances that neutralise acids; alkalis are bases that dissolve in water and release hydroxide ions (OH^-), such as sodium hydroxide. The pH scale measures how acidic or alkaline a solution is, from 0 to 14: values below 7 are acidic, 7 is neutral, and values above 7 are alkaline. Indicators, such as litmus paper and universal indicator, change colour to show whether a solution is acidic or alkaline.

When an acid reacts with a base, a neutralisation reaction takes place, producing a salt and water. Acids also react with metals to produce a salt and hydrogen gas, and with carbonates to produce a salt, water and carbon dioxide gas. These reactions explain everyday chemistry, from indigestion remedies that neutralise excess stomach acid to the fizzing when vinegar meets baking soda. Strong acids and alkalis are fully ionised in water, while weak acids and alkalis are only partly ionised.

Key definitions

- **Acid:** A substance that produces hydrogen ions (H^+) when dissolved in water.
- **Base:** A substance that neutralises an acid to form a salt and water.
- **Alkali:** A base that dissolves in water and produces hydroxide ions (OH^-).
- **pH scale:** A scale from 0 to 14 measuring how acidic or alkaline a solution is.
- **Neutralisation:** The reaction between an acid and a base to form a salt and water.
- **Salt:** A compound formed when the hydrogen ions of an acid are replaced by a metal or ammonium ion.
- **Indicator:** A substance that changes colour in acidic and alkaline solutions.

Important short questions

Q1. What is the difference between a base and an alkali?

A: A base is any substance that neutralises an acid, while an alkali is a base that dissolves in water, such as sodium hydroxide. All alkalis are bases, but not all bases are alkalis.

Q2. What are the products of the reaction between an acid and a metal carbonate?

A: A salt, water and carbon dioxide gas.

Q3. What does a pH of 3 tell you about a solution?

A: It tells you the solution is acidic, because its pH is below 7.

Q4. Why is a salt formed when hydrochloric acid reacts with sodium hydroxide?

A: The hydrogen ions from the acid combine with the hydroxide ions from the base to form water, and the remaining sodium and chloride ions form the salt, sodium chloride.

Important long questions

Q1. Describe how you would use an indicator to find out whether an unknown solution is acidic, neutral or alkaline.

A: Add a few drops of universal indicator to a sample of the solution and compare the colour with the pH colour chart. A red, orange or yellow colour shows an acidic solution, green shows a neutral solution, and blue or purple shows an alkaline solution. Litmus paper can also be used: blue litmus turns red in acid, and red litmus turns blue in alkali.

Q2. Write word equations for: (a) an acid reacting with a metal, (b) an acid reacting with a base, (c) an acid reacting with a carbonate, and give one example of each.

A: (a) Acid + metal → salt + hydrogen; for example, hydrochloric acid + magnesium → magnesium chloride + hydrogen. (b) Acid + base → salt + water; for example, hydrochloric acid + sodium hydroxide → sodium chloride + water. (c) Acid + carbonate → salt + water + carbon dioxide; for example, hydrochloric acid + calcium carbonate → calcium chloride + water + carbon dioxide.

MCQs

1. The pH of a neutral solution is: a) 0 b) 7 c) 14 d) 1 - **Answer: b**

2. Which of the following is an alkali? a) Copper oxide b) Sodium hydroxide c) Zinc carbonate d) Iron - **Answer: b**

3. An acid reacts with a metal to produce: a) Oxygen b) Carbon dioxide c) Hydrogen d) Nitrogen - **Answer: c**

4. Blue litmus paper turns red in: a) An alkali b) An acid c) Pure water d) A salt solution - **Answer: b**

5. The products of neutralisation are: a) Salt and hydrogen b) Salt and water c) Salt and oxygen d) Salt and carbon dioxide - **Answer: b**

Answer key (Ch. 7): 1-b, 2-b, 3-c, 4-b, 5-b.

Chapter 8: Periodic Table and Periodicity

Summary

The periodic table arranges the elements in order of increasing atomic number, so that elements with similar properties fall into the same vertical columns, called groups, while horizontal rows are called periods. This arrangement works because elements in the same group have the same number of electrons in their outer shell, and that outer-shell number largely decides how an element behaves in reactions. For example, group 1 elements (the alkali metals) all have one outer electron, which they readily lose, while group 7 elements (the halogens) have seven outer electrons and readily gain one.

Properties also change in regular patterns across the table, which is called periodicity. Moving across a period from left to right, atomic size generally decreases because the increasing nuclear charge pulls the outer electrons in more strongly, and the elements change gradually from metals to non-metals. Moving down a group, atomic size increases because each step adds a new electron shell. Reactivity also follows patterns: group 1 metals become more reactive down the group, while halogens become less reactive down the group. The periodic table also separates metals, which are found mostly on the left and in the centre, from non-metals on the right.

Key definitions

- **Group:** A vertical column of the periodic table, containing elements with the same number of outer-shell electrons.
- **Period:** A horizontal row of the periodic table, containing elements whose atoms have the same number of occupied electron shells.
- **Periodicity:** The repeating pattern of properties when elements are arranged by atomic number.
- **Atomic radius:** A measure of the size of an atom, usually taken as the distance from the nucleus to the outer shell.
- **Metal:** An element that tends to lose electrons to form positive ions, and is usually shiny, a good conductor, and malleable.
- **Non-metal:** An element that tends to gain or share electrons, and is usually a poor conductor of heat and electricity.

Important short questions

Q1. Why do elements in the same group have similar chemical properties?

A: Because their atoms have the same number of electrons in the outer shell, and it is the outer electrons that control how an element reacts.

Q2. What is the difference between a group and a period?

A: A group is a vertical column of elements with the same number of outer electrons; a period is a horizontal row of elements with the same number of electron shells.

Q3. Why does atomic size decrease across a period?

A: Across a period, the nuclear charge increases while electrons are added to the same shell, so the nucleus pulls the outer electrons in more strongly and the atom gets smaller.

Q4. Why do group 1 metals become more reactive down the group?

A: Down the group, the outer electron is further from the nucleus and more shielded by inner shells, so it is lost more easily and the metal reacts more readily.

Important long questions

Q1. Explain how the electronic structure of elements explains the shape of the periodic table.

A: Elements are arranged by atomic number, and as atomic number increases, electrons fill shells in a regular pattern. Elements with the same number of outer electrons fall in the same group, which is why they share chemical properties, and elements with the same number of occupied shells fall in the same period. The table is therefore a map of electronic structure, not just a list.

Q2. Describe and explain the trends in atomic size and metallic character across a period and down a group.

A: Across a period, atomic size decreases because of increasing nuclear charge on the same shell, and metallic character decreases as elements become less willing to lose electrons. Down a group, atomic size increases because a new shell is added at each step, and metallic character increases because outer electrons are held less tightly and are more easily lost.

MCQs

1. Elements in the same group have the same number of: a) Protons b) Neutrons c) Outer-shell electrons d) Shells - **Answer: c**

2. Moving across a period from left to right, atomic size generally: a) Increases b) Decreases c) Stays the same d) Doubles - **Answer: b**

3. Which group contains the alkali metals? a) Group 1 b) Group 7 c) Group 8 d) Group 0 - **Answer: a**

4. Which of the following is a halogen? a) Sodium b) Chlorine c) Calcium d) Neon - **Answer: b**

5. Down a group, atomic size increases because: a) Nuclear charge decreases b) A new electron shell is added c) Electrons are removed d) Atoms gain protons only - **Answer: b**

Answer key (Ch. 8): 1-c, 2-b, 3-a, 4-b, 5-b.

Chapter 9: Group Properties and Elements

Summary

Each group of the periodic table has its own character, and this chapter looks closely at the most important groups. Group 1, the alkali metals (lithium, sodium and potassium), are soft metals that react vigorously with water to form an alkaline solution and hydrogen gas, and their reactivity increases down the group. Group 7, the halogens (fluorine, chlorine, bromine and iodine), are reactive non-metals whose reactivity decreases down the group; a more reactive halogen can displace a less reactive halogen from its salt solution. Group 8 (also called group 0), the noble gases (helium, neon and argon), have full outer shells and are therefore very unreactive, which is why they are used where an inert atmosphere is needed, such as in light bulbs and airships.

The transition metals, found in the middle of the table, are hard, strong metals with high melting points that often form coloured compounds and act as catalysts. Comparing groups side by side shows the same underlying rule: the number of outer electrons decides the pattern of reactions, and the size of the atom and the shielding of the nucleus decide how strongly those outer electrons are held.

Key definitions

- **Alkali metals:** The reactive metals of group 1, with one outer electron, which form alkaline solutions when they react with water.
- **Halogens:** The reactive non-metals of group 7, with seven outer electrons, which form salts with metals.
- **Noble gases:** The very unreactive gases of group 8 (0), with full outer shells.
- **Displacement reaction:** A reaction in which a more reactive element takes the place of a less reactive element in a compound.
- **Transition metals:** The block of metals in the centre of the periodic table, typically hard and strong, with high melting points, often forming coloured compounds and acting as catalysts.

Important short questions

Q1. Why are the noble gases unreactive?

A: Because their atoms have full outer shells, so they have no tendency to gain, lose or share electrons.

Q2. What happens when sodium reacts with water?

A: Sodium reacts vigorously, producing sodium hydroxide solution (an alkali) and hydrogen gas, which may catch fire.

Q3. What is a displacement reaction? Give an example involving halogens.

A: A more reactive element displaces a less reactive one from its compound - for example, chlorine displaces bromine from potassium bromide solution, forming potassium chloride and bromine.

Q4. Why are halogens less reactive down the group?

A: Down the group, the outer shell is further from the nucleus and more shielded, so the atom attracts an extra electron less strongly.

Important long questions

Q1. Compare the reactivity trends of group 1 metals and group 7 non-metals down their groups, and explain both trends.

A: Group 1 reactivity increases down the group because the single outer electron becomes easier to lose as it gets further from the nucleus and more shielded. Group 7 reactivity decreases down the group because gaining an electron becomes harder for the same reason - the incoming electron would join a shell further from the nucleus and more shielded from its charge.

Q2. Describe two uses of noble gases and explain why their lack of reactivity makes them suitable.

A: Argon is used to fill light bulbs because it does not react with the hot filament, so the filament lasts longer. Helium is used in balloons and airships because it is light and, unlike hydrogen, does not burn. Both uses depend on the noble gases' full outer shells, which make them chemically inert.

MCQs

1. Which of the following is an alkali metal? a) Chlorine b) Sodium c) Neon d) Iron - **Answer: b**

2. The reactivity of the halogens down the group: a) Increases b) Decreases c) Stays the same d) Becomes zero at bromine - **Answer: b**

3. Noble gases are unreactive because they have: a) No electrons b) Full outer shells c) No nucleus d) Only one shell - **Answer: b**

4. Chlorine displaces bromine from potassium bromide because chlorine is: a) Less reactive b) More reactive c) A metal d) Larger - **Answer: b**

5. Which property is typical of transition metals? a) Low melting point b) Soft and easily cut c) Often form coloured compounds d) Very reactive with water - **Answer: c**

Answer key (Ch. 9): 1-b, 2-b, 3-b, 4-b, 5-c.

Chapter 10: Environmental Chemistry

Summary

Chemistry matters far beyond the laboratory because the air, water and soil around us are chemical systems. The atmosphere is a mixture of gases, mainly nitrogen and oxygen, with smaller amounts of carbon dioxide, water vapour and other gases. Burning fossil fuels releases carbon dioxide, which adds to the greenhouse effect and warms the planet, and also releases gases such as sulphur dioxide and nitrogen oxides, which dissolve in rainwater and fall as acid rain, damaging buildings, forests and lakes. Other air pollutants include carbon monoxide, which is poisonous because it reduces the blood's ability to carry oxygen, and tiny particles from smoke and dust.

Water pollution comes from sewage, industrial waste and farm chemicals, and it harms aquatic life and can make water unsafe to drink. Water can be made safe by treatment steps such as filtration, which removes solid particles, and chlorination, which kills harmful microbes. Chemistry also helps protect the environment: catalytic converters reduce harmful gases from vehicles, and understanding the ozone layer explains why certain chemicals that damage it have been restricted. The chapter's message is practical - the same chemical ideas used in class explain the biggest environmental problems and their solutions.

Key definitions

- **Greenhouse effect:** The warming of the Earth caused by gases in the atmosphere trapping heat radiated from the surface.
- **Acid rain:** Rain made acidic by dissolved gases such as sulphur dioxide and nitrogen oxides.
- **Pollutant:** A substance introduced into the environment in harmful amounts.
- **Filtration:** A method of separating an insoluble solid from a liquid by passing the mixture through a filter.
- **Chlorination:** The addition of chlorine to water to kill harmful bacteria and other microbes.
- **Ozone layer:** The region of the upper atmosphere containing ozone, which absorbs much of the Sun's harmful ultraviolet radiation.

Important short questions

Q1. What causes acid rain?

A: Gases such as sulphur dioxide and nitrogen oxides, released mainly by burning fossil fuels, dissolve in rainwater and make it acidic.

Q2. How does carbon dioxide contribute to global warming?

A: Carbon dioxide in the atmosphere traps heat radiated from the Earth's surface, strengthening the greenhouse effect and raising global temperatures.

Q3. Why is carbon monoxide dangerous?

A: It binds to the haemoglobin in red blood cells more strongly than oxygen does, reducing the blood's ability to carry oxygen around the body.

Q4. Name two steps in treating water to make it safe to drink, and state what each removes.

A: Filtration removes suspended solid particles, and chlorination kills harmful bacteria and microbes.

Important long questions

Q1. Explain how burning fossil fuels leads to both acid rain and an enhanced greenhouse effect.

A: Burning fossil fuels releases carbon dioxide, which builds up in the atmosphere and traps outgoing heat, strengthening the greenhouse effect. The same fuels often contain sulphur, and high-temperature burning in engines produces nitrogen oxides; these gases dissolve in atmospheric water to form acids, which fall as acid rain and damage buildings, soil, crops and aquatic life.

Q2. Describe the effects of water pollution and the main steps used to treat polluted water.

A: Sewage, industrial chemicals and farm fertilisers pollute rivers and lakes, killing aquatic life, spreading disease and making water unsafe for drinking and irrigation. Treatment includes screening and settling to remove large solids, filtration to remove fine particles, and chlorination or other disinfection to kill microbes, before the water is supplied for use.

MCQs

1. The main gas responsible for the enhanced greenhouse effect from burning fossil fuels is: a) Oxygen b) Nitrogen c) Carbon dioxide d) Helium - **Answer: c**

2. Acid rain is mainly caused by: a) Oxygen and nitrogen b) Sulphur dioxide and nitrogen oxides c) Carbon monoxide only d) Water vapour - **Answer: b**

3. Chlorination of water is done to: a) Remove sand b) Kill harmful microbes c) Add flavour d) Remove salts - **Answer: b**

4. The ozone layer protects life by absorbing: a) Visible light b) Infrared radiation c) Ultraviolet radiation d) Sound - **Answer: c**

5. Which of the following is a poisonous gas from incomplete burning of fuels? a) Carbon dioxide b) Carbon monoxide c) Nitrogen d) Oxygen - **Answer: b**

Answer key (Ch. 10): 1-c, 2-b, 3-b, 4-c, 5-b.

Chapter 11: Hydrocarbons

Summary

Hydrocarbons are compounds made up of hydrogen and carbon only, and they are the main components of fossil fuels such as natural gas, petrol and diesel. The simplest hydrocarbons are the alkanes, which have only single bonds between carbon atoms; methane, ethane and propane are the first members, and they are relatively unreactive. The alkenes have at least one double bond between carbon atoms, which makes them more reactive: they can take part in addition reactions, in which the double bond opens up and atoms add across it, and they decolourise bromine water, which is a test that distinguishes them from alkanes.

Hydrocarbons are obtained from crude oil by fractional distillation, which separates the mixture into fractions according to boiling point. Smaller fractions, such as petrol, are more useful as fuels, and larger fractions can be broken down into smaller, more useful molecules by cracking. When hydrocarbons burn completely in plenty of oxygen, they produce carbon dioxide and water; when they burn incompletely, they produce carbon monoxide and soot, which is why good ventilation matters for gas heaters and engines. Hydrocarbons are also the starting materials for plastics and many other useful products.

Key definitions

- **Hydrocarbon:** A compound containing hydrogen and carbon only.
- **Alkane:** A hydrocarbon with only single bonds between carbon atoms (a saturated hydrocarbon).
- **Alkene:** A hydrocarbon with at least one double bond between carbon atoms (an unsaturated hydrocarbon).
- **Saturated compound:** A compound whose carbon atoms are joined only by single bonds, holding the maximum number of hydrogen atoms.
- **Fractional distillation:** The separation of a mixture of liquids with different boiling points, using a fractionating column.
- **Cracking:** The breaking down of large hydrocarbon molecules into smaller, more useful ones, using heat and a catalyst.
- **Addition reaction:** A reaction in which atoms add across a carbon-carbon double bond, turning an unsaturated compound into a saturated one.

Important short questions

Q1. What is the difference between an alkane and an alkene?

A: Alkanes have only single bonds between carbon atoms and are saturated, while alkenes have at least one double bond and are unsaturated, making them more reactive.

Q2. How can you test whether a hydrocarbon is an alkene?

A: Add bromine water: an alkene decolourises the orange bromine water because an addition reaction takes place, while an alkane does not.

Q3. What are the products of complete combustion of a hydrocarbon?

A: Carbon dioxide and water.

Q4. Why is incomplete combustion dangerous?

A: It produces carbon monoxide, a poisonous gas, and soot, instead of only carbon dioxide and water.

Important long questions

Q1. Describe how fractional distillation separates crude oil into fractions, and explain why the fractions come off at different levels.

A: Crude oil is heated until it vaporises, and the vapour rises up a fractionating column that is hot at the bottom and cooler at the top. Fractions with low boiling points, such as petrol gases, stay as vapour and rise to the top, while fractions with high boiling points, such as bitumen, condense near the bottom. Each fraction condenses at the level where the temperature matches its boiling range, so the mixture is separated by boiling point.

Q2. Explain what cracking is and why refineries carry it out.

A: Cracking breaks large, less useful hydrocarbon molecules into smaller ones by heating them with a catalyst. Refineries do this because the demand for small molecules, such as petrol, is greater than the amount present in crude oil, and because cracking also produces alkenes, which are valuable starting materials for plastics and other chemicals.

MCQs

1. A hydrocarbon contains: a) Carbon and oxygen only b) Hydrogen and carbon only c) Hydrogen and oxygen only d) Carbon, hydrogen and nitrogen - **Answer: b**

2. Which of the following is an alkane? a) Ethene b) Methane c) Propene d) Butene - **Answer: b**

3. Alkenes decolourise bromine water because they undergo: a) Substitution b) Addition c) Neutralisation d) Combustion - **Answer: b**

4. The complete combustion of a hydrocarbon produces: a) Carbon monoxide and soot b) Carbon dioxide and water c) Carbon and hydrogen d) Oxygen and water - **Answer: b**

5. Cracking is used to: a) Join small molecules into large ones b) Break large hydrocarbons into smaller ones c) Remove salt from crude oil d) Burn fuels completely - **Answer: b**

Answer key (Ch. 11): 1-b, 2-b, 3-b, 4-b, 5-b.

Chapter 12: Empirical Data Collection and Analysis

Summary

Science depends on evidence, and this chapter is about collecting that evidence well and making sense of it. A good investigation starts with a clear question and a prediction, called a hypothesis, and then identifies the variables: the independent variable is the one the experimenter changes, the dependent variable is the one that is measured, and control variables are the ones kept the same so that the test is fair. Measurements should be repeated so that an average can be taken and unusual results, called anomalies, can be spotted. The range of an instrument and how precisely it can be read decide how reliable the measurements are.

Once data are collected, they are organised in tables and shown as charts or graphs. A line graph is used when both variables are continuous, and the line of best fit shows the overall trend rather than joining every point. A straight-line graph through the origin shows direct proportion: doubling one variable doubles the other. Drawing conclusions means describing what the pattern shows, judging whether the evidence supports the hypothesis, and suggesting improvements, such as taking more readings or controlling a variable more carefully. Honest reporting matters: results are recorded as found, not adjusted to fit what was expected.

Key definitions

- **Hypothesis:** A testable prediction, based on scientific ideas, about what will happen in an investigation.
- **Independent variable:** The variable that the experimenter deliberately changes.
- **Dependent variable:** The variable that is measured to see the effect of the change.
- **Control variable:** A variable kept the same throughout an investigation so the test is fair.
- **Anomaly (outlier):** A result that does not fit the overall pattern of the data.
- **Line of best fit:** A smooth line drawn through or near plotted points to show the overall trend of the data.

Important short questions

Q1. What is the difference between an independent and a dependent variable?

A: The independent variable is the one the experimenter changes, while the dependent variable is the one measured to observe the effect of that change.

Q2. Why should measurements be repeated in an investigation?

A: Repeating measurements allows an average to be calculated, reduces the effect of random errors, and helps anomalies to be identified.

Q3. What does a straight line through the origin on a graph show?

A: It shows direct proportion between the two variables - when one doubles, the other doubles too.

Q4. What is an anomaly, and what should be done with it?

A: An anomaly is a result that does not fit the pattern. It should be noted, checked by repeating the reading if possible, and not simply deleted without reason.

Important long questions

Q1. A student investigates how temperature affects how quickly sugar dissolves in water. Identify the independent, dependent and control variables, and explain how the student could make the test fair.

A: The independent variable is the temperature of the water, and the dependent variable is the time taken for the sugar to dissolve. Control variables include the mass of sugar, the volume of water, the stirring method and the size of the sugar grains, all kept the same. The test is fair only if temperature is the one thing changed, so any difference in dissolving time can be attributed to temperature.

Q2. Describe how to present a set of experimental results as a graph and draw a valid conclusion from it.

A: First, organise the results in a table with the independent variable in the first column and repeats and averages shown. Then plot the averages on a graph with the independent variable on the horizontal axis and the dependent variable on the vertical axis, using sensible scales, and draw a line of best fit rather than joining points dot to dot. The conclusion describes the pattern shown by the line, states whether it supports the hypothesis, and notes any anomalies or limits of the evidence.

MCQs

1. The variable that the experimenter changes is the: a) Dependent variable b) Independent variable c) Control variable d) Result - **Answer: b**

2. Repeating a measurement helps to: a) Increase the range b) Identify anomalies and calculate an average c) Change the hypothesis d) Remove the control variables - **Answer: b**

3. A line graph is most suitable when: a) Both variables are continuous b) Data are categories only c) Only one reading exists d) No pattern is expected - **Answer: a**

4. A result that does not fit the pattern is called: a) A hypothesis b) An anomaly c) A control d) A conclusion - **Answer: b**

5. In a fair test, the control variables are: a) Changed each time b) Measured as results c) Kept the same d) Ignored - **Answer: c**

Answer key (Ch. 12): 1-b, 2-b, 3-a, 4-b, 5-c.

Chapter 13: Laboratory and Practical Skills

Summary

Working safely and skilfully in the laboratory is part of learning chemistry, not an extra. Before any practical work, the experimenter reads the instructions, checks hazard warnings on chemical labels, and wears eye protection; long hair is tied back and bags are kept out of walkways. Common apparatus includes beakers and conical flasks for holding and mixing liquids, measuring cylinders for approximate volumes, burettes and pipettes for accurate volumes, test tubes for small-scale reactions, and the Bunsen burner for heating, with its air hole controlling whether the flame is yellow and sooty or blue and hotter.

Good practical skills also include measuring accurately, reading a scale at eye level and from the bottom of the meniscus for liquids, recording results immediately in a table, and cleaning up spills and disposing of waste as instructed rather than pouring chemicals down the sink without permission. When heating, test tubes are pointed away from people, and hot apparatus is handled with tongs or a holder. If a chemical splashes into the eyes, they are washed with plenty of water and a teacher is told at once. These habits protect the student, other people in the laboratory, and the reliability of the results.

Key definitions

- **Hazard:** Something that could cause harm, such as a corrosive chemical or a hot flame.
- **Risk assessment:** Thinking through the hazards of an experiment and the precautions that reduce the risk before starting work.
- **Meniscus:** The curved surface of a liquid in a container; volumes are read from its lowest point for most liquids.
- **Burette:** A long, graduated tube with a tap, used to deliver accurate, variable volumes of liquid.
- **Pipette:** A tube used to measure and transfer a fixed, accurate volume of liquid.
- **Eye protection:** Safety spectacles or goggles worn to protect the eyes from splashes and flying fragments.

Important short questions

Q1. Why should eye protection be worn in the laboratory?

A: To protect the eyes from chemical splashes, hot liquids and fragments if glassware breaks.

Q2. What is the difference between a measuring cylinder and a burette?

A: A measuring cylinder measures approximate volumes quickly, while a burette delivers accurate, variable volumes, drop by drop, through a tap.

Q3. How should the volume of a liquid in a measuring cylinder be read?

A: At eye level, reading from the bottom of the meniscus, the curved surface of the liquid.

Q4. What should you do immediately if a chemical splashes into your eyes?

A: Wash the eyes with plenty of clean water and tell the teacher at once.

Important long questions

Q1. Describe the safety precautions you would take before and during heating a liquid in a test tube over a Bunsen burner.

A: Before starting, tie back long hair, wear eye protection, check that the test tube is not cracked and that the bench is clear. Light the burner safely, adjust the air hole for a blue flame, and hold the test tube with a holder, pointing its mouth away from yourself and others. Heat gently, moving the tube through the flame, and never look down into the mouth of the tube.

Q2. Explain why accurate measurement and immediate recording matter in practical work.

A: Accurate measurement with the right apparatus, such as a pipette or burette for small volumes, reduces errors that could change the conclusion. Recording results immediately, in an organised table, prevents forgotten or misremembered readings, makes anomalies easier to spot, and lets others check and repeat the work.

MCQs

1. Which piece of apparatus is best for measuring exactly 25.0 cm^3 of a solution? a) Beaker b) Measuring cylinder c) Pipette d) Conical flask - **Answer: c**

2. The volume of a liquid in a measuring cylinder is read from the: a) Top of the meniscus b) Bottom of the meniscus c) Side of the cylinder d) Top of the liquid's edge - **Answer: b**

3. A blue Bunsen flame is produced when the air hole is: a) Closed b) Open c) Blocked with a stopper d) Half covered with paper - **Answer: b**

4. The first action if a chemical splashes into the eyes is to: a) Rub the eyes b) Wash with plenty of water c) Wait for it to dry d) Cover the eyes - **Answer: b**

5. Which of the following is a safety rule in the laboratory? a) Point test tubes towards your face b) Taste chemicals to identify them c) Wear eye protection when heating chemicals d) Pour waste chemicals anywhere - **Answer: c**

Answer key (Ch. 13): 1-c, 2-b, 3-b, 4-b, 5-c.