



NCERT Chemistry Question Paper (Class - 11)

(Chemistry) Chapter 1 Some Basic Concepts Of Chemistry

NCERT Exercises Questions

Question 1.1 Calculate the molecular mass of the following :

- (i) H_2O
- (ii) CO_2
- (iii) CH_4

Question 1.2 Calculate the mass per cent of different elements present in sodium sulphate (Na_2SO_4).

Question 1.3 Determine the empirical formula of an oxide of iron which has 69.9% iron and 30.1% dioxygen by mass.

Question 1.4 Calculate the amount of carbon dioxide that could be produced when

- (i) 1 mole of carbon is burnt in air.
- (ii) 1 mole of carbon is burnt in 16 g of dioxygen.
- (iii) 2 moles of carbon are burnt in 16 g of dioxygen.

Question 1.5 Calculate the mass of sodium acetate (CH_3COONa) required to make 500 mL of 0.375 molar aqueous solution. Molar mass of sodium acetate is $82.0245 \text{ g mol}^{-1}$.

Question 1.6 Calculate the concentration of nitric acid in moles per litre in a sample which has a density, 1.41 g mL^{-1} and the mass per cent of nitric acid in it being 69%.

Question 1.7 How much copper can be obtained from 100 g of copper sulphate (CuSO_4) ?

Question 1.8 Determine the molecular formula of an oxide of iron in which the mass per cent of iron and oxygen are 69.9 and 30.1 respectively.

Question 1.9 Calculate the atomic mass (average) of chlorine using the following data :
% Natural Abundance Molar Mass
 ^{35}Cl 75.77 34.9689
 ^{37}Cl 24.23 36.9659

Question 1.10 In three moles of ethane (C_2H_6), calculate the following :

- (i) Number of moles of carbon atoms.
- (ii) Number of moles of hydrogen atoms.
- (iii) Number of molecules of ethane.

Question 1.11 What is the concentration of sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L?

Question 1.12 If the density of methanol is 0.793 kg L^{-1} , what is its volume needed for making 2.5 L of its 0.25 M solution?

Question 1.13 Pressure is determined as force per unit area of the surface. The SI unit of pressure, pascal is as shown below : $1 \text{ Pa} = 1 \text{ N m}^{-2}$ If mass of air at sea level is 1034 g cm^{-2} , calculate the pressure in pascal.

Question 1.14 What is the SI unit of mass? How is it defined?

Question 1.15 Match the following prefixes with their multiples: Prefixes Multiples

- (i) micro 10⁶
- (ii) deca 10⁹
- (iii) mega 10⁶
- (iv) giga 10¹⁵
- (v) femto 10⁻¹⁵

Question 1.16 What do you mean by significant figures ?

Question 1.17 A sample of drinking water was found to be severely contaminated with chloroform, CHCl₃, supposed to be carcinogenic in nature. The level of contamination was 15 ppm (by mass).

- (i) Express this in percent by mass.
- (ii) Determine the molality of chloroform in the water sample.

Question 1.18 Express the following in the scientific notation:

- (i) 0.0048
- (ii) 234,000
- (iii) 8008
- (iv) 500.0
- (v) 6.0012

Question 1.19 How many significant figures are present in the following?

- (i) 0.0025
- (ii) 208
- (iii) 5005
- (iv) 126,000
- (v) 500.0
- (vi) 2.0034

Question 1.20 Round up the following upto three significant figures:

- (i) 34.216
- (ii) 10.4107
- (iii) 0.04597
- (iv) 2808

Question 1.21 The following data are obtained when dinitrogen and dioxygen react together to form different compounds :
Mass of dinitrogen Mass of dioxygen

(i) 14 g 16 g

(ii) 14 g 32 g

(iii) 28 g 32 g

(iv) 28 g 80 g

(a) Which law of chemical combination is obeyed by the above experimental data? Give its statement.

(b) Fill in the blanks in the following conversions:

(i) 1 km = mm = pm

(ii) 1 mg = kg = ng

(iii) 1 mL = L = dm³

Question 1.22 If the speed of light is $3.0 \times 10^8 \text{ m s}^{-1}$, calculate the distance covered by light in 2.00 ns.

Question 1.23 In a reaction $A + B_2 \rightarrow AB_2$ Identify the limiting reagent, if any, in the following reaction mixtures.

(i) 300 atoms of A + 200 molecules of B

(ii) 2 mol A + 3 mol B

(iii) 100 atoms of A + 100 molecules of B

(iv) 5 mol A + 2.5 mol B

(v) 2.5 mol A + 5 mol B

Question 1.24 Dinitrogen and dihydrogen react with each other to produce ammonia according to the following chemical equation: $N_2(g) + H_2(g) \rightarrow 2NH_3(g)$

(i) Calculate the mass of ammonia produced if $2.00 \times 10^3 \text{ g}$ dinitrogen reacts with $1.00 \times 10^3 \text{ g}$ of dihydrogen.

(ii) Will any of the two reactants remain unreacted?

(iii) If yes, which one and what would be its mass?

Question 1.25 How are 0.50 mol Na_2CO_3 and 0.50 M Na_2CO_3 different?

Question 1.26 If ten volumes of dihydrogen gas reacts with five volumes of dioxygen gas, how many volumes of water vapour would be produced?

Question 1.27 Convert the following into basic units:

(i) 28.7 pm

- (ii) 15.15 pm
- (iii) 25365 mg

Question 1.28 Which one of the following will have largest number of atoms?

- (i) 1 g Au (s)
- (ii) 1 g Na (s)
- (iii) 1 g Li (s)
- (iv) 1 g of Cl₂(g)

Question 1.29 Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040.

Question 1.30 What will be the mass of one ¹²C atom in g ?

Question 1.31 How many significant figures should be present in the answer of the following calculations?

- (i) $0.02856 \times 29 \times 0.5$
- (ii) 5×5.364
- (iii) $0.0125 + 0.7864 + 0.0215$

Question 1.32 Use the data given in the following table to calculate the molar mass of naturally occurring argon isotopes:

Isotope	Isotopic molar mass	Abundance
³⁶ Ar	35.96755 g mol ⁻¹	0.337%
³⁸ Ar	37.96272 g mol ⁻¹	0.063%
⁴⁰ Ar	39.9624 g mol ⁻¹	99.600%

Question 1.33 Calculate the number of atoms in each of the following

- (i) 52 moles of Ar
- (ii) 52 u of He
- (iii) 52 g of He.

Question 1.34 A welding fuel gas contains carbon and hydrogen only. Burning a small sample of it in oxygen gives 3.38 g carbon dioxide, 0.690 g of water and no other products. A volume of 10.0 L (measured at STP) of this welding gas is found to weigh 11.6 g. Calculate (i) empirical formula, (ii) molar mass of the gas, and (iii) molecular formula.

Question 1.35 Calcium carbonate reacts with aqueous HCl to give CaCl₂ and CO₂ according to the reaction, CaCO₃ (s) + 2 HCl (aq) → CaCl₂ (aq) + CO₂(g) + H₂O(l) What mass of CaCO₃ is required

to react completely with 25 mL of 0.75 M HCl?

Question 1.36 Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction $4 \text{HCl} (\text{aq}) + \text{MnO}_2(\text{s}) \rightarrow 2\text{H}_2\text{O} (\text{l}) + \text{MnCl}_2(\text{aq}) + \text{Cl}_2 (\text{g})$ How many grams of HCl react with 5.0 g of manganese dioxide?

(Chemistry) Chapter 2 Structure Of Atom

NCERT Exercises Questions

Question 2.1 (i) Calculate the number of electrons which will together weigh one gram.
(ii) Calculate the mass and charge of one mole of electrons.

Question 2.2 (i) Calculate the total number of electrons present in one mole of methane.
(ii) Find (a) the total number and (b) the total mass of neutrons in 7 mg of ^{14}C . (Assume that mass of a neutron = 1.675×10^{-27} kg).
(iii) Find (a) the total number and (b) the total mass of protons in 34 mg of NH_3 at STP. Will the answer change if the temperature and pressure are changed ?

Question 2.3 How many neutrons and protons are there in the following nuclei ?
 $^{13}_6\text{C}$, $^{16}_8\text{O}$, $^{24}_{12}\text{Mg}$,

Question 2.4 Write the complete symbol for the atom with the given atomic number (Z) and atomic mass (A)
(i) $Z = 17$, $A = 35$.
(ii) $Z = 92$, $A = 233$.
(iii) $Z = 4$, $A = 9$.

Question 2.5 Yellow light emitted from a sodium lamp has a wavelength (λ) of 580 nm. Calculate the frequency (ν) and wavenumber ($\bar{\nu}$) of the yellow light.

Question 2.6 Find energy of each of the photons which

(i) correspond to light of frequency 3×10^{15} Hz.

(ii) have wavelength of 0.50 \AA .

Question 2.7 Calculate the wavelength, frequency and wavenumber of a light wave whose period is 2.0×10^{-10} s.

Question 2.8 What is the number of photons of light with a wavelength of 4000 pm that provide 1 J of energy?

Question 2.9 A photon of wavelength $4 \times 10^{-7} \text{ m}$ strikes on metal surface, the work function of the metal being 2.13 eV . Calculate (i) the energy of the photon (eV), (ii) the kinetic energy of the emission, and (iii) the velocity of the photoelectron ($1 \text{ eV} = 1.6020 \times 10^{-19} \text{ J}$).

Question 2.10 Electromagnetic radiation of wavelength 242 nm is just sufficient to ionise the sodium atom. Calculate the ionisation energy of sodium in kJ mol^{-1} .

Question 2.11 A 25 watt bulb emits monochromatic yellow light of wavelength of $0.57 \mu\text{m}$. Calculate the rate of emission of quanta per second.

Question 2.12 Electrons are emitted with zero velocity from a metal surface when it is exposed to radiation of wavelength 6800 \AA . Calculate threshold frequency (ν_0) and work function (W_0) of the metal.

Question 2.13 What is the wavelength of light emitted when the electron in a hydrogen atom undergoes transition from an energy level with $n = 4$ to an energy level with $n = 2$?

Question 2.14 How much energy is required to ionise a H atom if the electron occupies $n = 5$ orbit? Compare your answer with the ionization enthalpy of H atom (energy required to remove the electron from $n = 1$ orbit).

Question 2.15 What is the maximum number of emission lines when the excited electron of a H atom in $n = 6$ drops to the ground state?

Question 2.16 (i) The energy associated with the first orbit in the hydrogen atom is $-2.18 \times 10^{-18} \text{ J atom}^{-1}$. What is the energy associated with the fifth orbit?

(ii) Calculate the radius of Bohr's fifth orbit for hydrogen atom.

Question 2.17 Calculate the wavenumber for the longest wavelength transition in the Balmer series of atomic hydrogen.

Question 2.18 What is the energy in joules, required to shift the electron of the hydrogen atom from the first Bohr orbit to the fifth Bohr orbit and what is the wavelength of the light emitted when the electron returns to the ground state? The ground state electron energy is -2.18×10^{-11} ergs.

Question 2.19 The electron energy in hydrogen atom is given by $E_n = (-2.18 \times 10^{-18})/n^2$ J. Calculate the energy required to remove an electron completely from the $n = 2$ orbit. What is the longest wavelength of light in cm that can be used to cause this transition?

Question 2.20 Calculate the wavelength of an electron moving with a velocity of 2.05×10^7 m s⁻¹.

Question 2.21 The mass of an electron is 9.1×10^{-31} kg. If its K.E. is 3.0×10^{-25} J, calculate its wavelength.

Question 2.22 Which of the following are isoelectronic species i.e., those having the same number of electrons? Na⁺, K⁺, Mg²⁺, Ca²⁺, S²⁻, Ar.

Question 2.23 (i) Write the electronic configurations of the following ions:

- (a) H⁻
- (b) Na⁺
- (c) O²⁻
- (d) F⁻

(ii) What are the atomic numbers of elements whose outermost electrons are represented by

- (a) 3s¹
- (b) 2p³
- (c) 3p⁵ ?

(iii) Which atoms are indicated by the following configurations ?

- (a) [He] 2s¹
- (b) [Ne] 3s² 3p³
- (c) [Ar] 4s² 3d¹.

Question 2.24 What is the lowest value of n that allows g orbitals to exist?

Question 2.25 An electron is in one of the $3d$ orbitals. Give the possible values of n , l and m_l for this electron.

Question 2.26 An atom of an element contains 29 electrons and 35 neutrons. Deduce (i) the number of protons and (ii) the electronic configuration of the element.

Question 2.27 Give the number of electrons in the species H , H^+ and O^{2+} , O^+

Question 2.28 (i) An atomic orbital has $n = 3$. What are the possible values of l and m_l ?

(ii) List the quantum numbers (m_l and l) of electrons for $3d$ orbital.

(iii) Which of the following orbitals are possible? $1p$, $2s$, $2p$ and $3f$

Question 2.29 Using s , p , d notations, describe the orbital with the following quantum numbers.

(a) $n=1, l=0$;

(b) $n = 3; l=1$

(c) $n 4; l =2$;

(d) $n=4; l=3$.

Question 2.30 Explain, giving reasons, which of the following sets of quantum numbers are not possible.

(a) $n = 0, l = 0, m_l = 0, m_s = + \frac{1}{2}$

(b) $n = 1, l = 0, m_l = 0, m_s = - \frac{1}{2}$

(c) $n = 1, l = 1, m_l = 0, m_s = + \frac{1}{2}$

(d) $n = 2, l = 1, m_l = 0, m_s = - \frac{1}{2}$

(e) $n = 3, l = 3, m_l = -3, m_s = + \frac{1}{2}$

(f) $n = 3, l = 1, m_l = 0, m_s = + \frac{1}{2}$

Question 2.31 How many electrons in an atom may have the following quantum numbers?

(a) $n = 4, m_s = - \frac{1}{2}$

(b) $n = 3, l = 0$

Question 2.32 Show that the circumference of the Bohr orbit for the hydrogen atom is an integral multiple of the de Broglie wavelength associated with the electron revolving around the orbit.

Question 2.33 What transition in the hydrogen spectrum would have the same wavelength as the Balmer transition $n = 4$ to $n = 2$ of He^+ spectrum ?

Question 2.34 Calculate the energy required for the process $\text{He}^+ (\text{g}) \rightarrow \text{He}^{2+} (\text{g}) + \text{e}^-$. The ionization energy for the H atom in the ground state is $2.18 \times 10^{-18} \text{ J atom}^{-1}$

Question 2.35 If the diameter of a carbon atom is 0.15 nm, calculate the number of carbon atoms which can be placed side by side in a straight line across length of scale of length 20 cm long.

Question 2.36 2×10^8 atoms of carbon are arranged side by side. Calculate the radius of carbon atom if the length of this arrangement is 2.4 cm.

Question 2.37 The diameter of zinc atom is 2.6 Å. Calculate (a) radius of zinc atom in pm and (b) number of atoms present in a length of 1.6 cm if the zinc atoms are arranged side by side lengthwise.

Question 2.38 A certain particle carries $2.5 \times 10^{-16} \text{ C}$ of static electric charge. Calculate the number of electrons present in it.

Question 2.39 In Milikan's experiment, static electric charge on the oil drops has been obtained by shining X-rays. If the static electric charge on the oil drop is $-1.282 \times 10^{-18} \text{ C}$, calculate the number of electrons present on it.

Question 2.40 In Rutherford's experiment, generally the thin foil of heavy atoms, like gold, platinum etc. have been used to be bombarded by the α -particles. If the thin foil of light atoms like aluminium etc. is used, what difference would be observed from the above results ?

Question 2.41 Symbols ${}_{79}^{35}\text{Br}$ and ${}^{79}\text{Br}$ can be written, whereas symbols ${}^{35}{}_{79}\text{Br}$ and ${}^{35}\text{Br}$ are not acceptable. Answer briefly.

Question 2.42 An element with mass number 81 contains 31.7% more neutrons as compared to protons. Assign the atomic symbol.

Question 2.43 An ion with mass number 37 possesses one unit of negative charge. If the ion contains 11.1% more neutrons than the electrons, find the symbol of the ion.

Question 2.44 An ion with mass number 56 contains 3 units of positive charge and 30.4% more neutrons than electrons. Assign the symbol to this ion.

Question 2.45 Arrange the following type of radiations in increasing order of frequency:

(a) radiation from microwave oven (b) amber light from traffic signal (c) radiation from FM radio (d) cosmic rays from outer space and (e) X-rays.

Question 2.46 Nitrogen laser produces a radiation at a wavelength of 337.1 nm. If the number of photons emitted is 5.6×10^{24} , calculate the power of this laser.

Question 2.47 Neon gas is generally used in the sign boards. If it emits strongly at 616 nm, calculate

(a) the frequency of emission (b) distance traveled by this radiation in 30 s (c) energy of quantum and (d) number of quanta present if it produces 2 J of energy.

Question 2.48 In astronomical observations, signals observed from the distant stars are generally

(Chemistry) Chapter 3 Classification of Elements and Periodicity in Properties

NCERT Exercises Questions

Question 3.1 What is the basic theme of organisation in the periodic table?

Question 3.2 Which important property did Mendeleev use to classify the elements in his periodic table and did he stick to that?

Question 3.3 What is the basic difference in approach between the Mendeleev's Periodic Law and the Modern Periodic Law?

Question 3.4 On the basis of quantum numbers, justify that the sixth period of the periodic table should have 32 elements.

Question 3.5 In terms of period and group where would you locate the element with $Z = 114$?

Question 3.6 Write the atomic number of the element present in the third period and seventeenth group of the periodic table.

Question 3.7 Which element do you think would have been named by

(i) Lawrence Berkeley Laboratory

(ii) Seaborg's group?

Question 3.8 Why do elements in the same group have similar physical and chemical properties?

Question 3.9 What does atomic radius and ionic radius really mean to you?

Question 3.10 How do atomic radius vary in a period and in a group? How do you explain the variation?

Question 3.11 What do you understand by isoelectronic species? Name a species that will be isoelectronic with each of the following atoms or ions.

(i) F^-

(ii) Ar

(iii) Mg^{2+}

(iv) Rb^+

Question 3.12 Consider the following species : N_3^- , O_2^- , F^- , Na^+ , Mg^{2+} and Al^{3+}

(a) What is common in them?

(b) Arrange them in the order of increasing ionic radii.

Question 3.13 Explain why cation are smaller and anions larger in radii than their parent atoms?

Question 3.14 What is the significance of the terms — 'isolated gaseous atom' and 'ground state' while defining the ionization enthalpy and electron gain enthalpy? Hint : Requirements for comparison purposes.

Question 3.15 Energy of an electron in the ground state of the hydrogen atom is $-2.18 \times 10^{-18} J$. Calculate the ionization enthalpy of atomic hydrogen in terms of $J mol^{-1}$. Hint: Apply the idea of mole concept to derive the answer.

Question 3.16 Among the second period elements the actual ionization enthalpies are in the order $\text{Li} < \text{B} < \text{Be} < \text{C} < \text{O} < \text{N} < \text{F} < \text{Ne}$. Explain why

- (i) Be has higher $\Delta_i H$ than B
- (ii) O has lower $\Delta_i H$ than N and F?

Question 3.17 How would you explain the fact that the first ionization enthalpy of sodium is lower than that of magnesium but its second ionization enthalpy is higher than that of magnesium?

Question 3.18 What are the various factors due to which the ionization enthalpy of the main group elements tends to decrease down a group?

Question 3.19 The first ionization enthalpy values (in kJ mol^{-1}) of group 13 elements are : B Al Ga In Tl 801 577 579 558 589 How would you explain this deviation from the general trend ?

Question 3.20 Which of the following pairs of elements would have a more negative electron gain enthalpy?

- (i) O or F
- (ii) F or Cl

Question 3.21 Would you expect the second electron gain enthalpy of O as positive, more negative or less negative than the first? Justify your answer.

Question 3.22 What is the basic difference between the terms electron gain enthalpy and electronegativity?

Question 3.23 How would you react to the statement that the electronegativity of N on Pauling scale is Question 3.0 in all the nitrogen compounds?

Question 3.24 Describe the theory associated with the radius of an atom as it

- (a) gains an electron
- (b) loses an electron

Question 3.25 Would you expect the first ionization enthalpies for two isotopes of the same element to be the same or different? Justify your answer.

Question 3.26 What are the major differences between metals and non-metals?

Question 3.27 Use the periodic table to answer the following questions.

- (a) Identify an element with five electrons in the outer subshell.
- (b) Identify an element that would tend to lose two electrons.
- (c) Identify an element that would tend to gain two electrons.
- (d) Identify the group having metal, non-metal, liquid as well as gas at the room temperature.

Question 3.28 The increasing order of reactivity among group 1 elements is $\text{Li} < \text{Na} < \text{K} < \text{Rb} < \text{Cs} < \text{Fr}$. Explain.

Question 3.29 Write the general outer electronic configuration of s-, p-, d- and f- block elements.

Question 3.30 Assign the position of the element having outer electronic configuration

- (i) ns^2np^4 for $n=3$
- (ii) $(n-1)d^2ns^2$ for $n=4$, and
- (iii) $(n-2)f^7(n-1)d^1ns^2$ for $n=6$, in the periodic table.

Question 3.31 The first ($\Delta_i H_1$) and the second ($\Delta_i H_2$) ionization enthalpies (in kJ mol^{-1}) and the ($\Delta_{\text{eg}} H$) electron gain enthalpy (in kJ mol^{-1}) of a few elements are given below: Elements ΔH_1 ΔH_2 $\Delta_{\text{eg}} H$ I 520 7300 -60 II 419 3051 -48 III 1681 3374 -328 IV 1008 1846 -295 V 2372 5251 +48 VI 738 1451 -40 Which of the above elements is likely to be :

- (a) the least reactive element.
- (b) the most reactive metal.
- (c) the most reactive non-metal.
- (d) the least reactive non-metal.
- (e) the metal which can form a stable binary halide of the formula MX_2 ($X=\text{halogen}$).
- (f) the metal which can form a predominantly stable covalent halide of the formula MX ($X=\text{halogen}$)?

Question 3.32 Predict the formulas of the stable binary compounds that would be formed by the combination of the following pairs of elements.

- (a) Lithium and oxygen
- (b) Magnesium and nitrogen
- (c) Aluminium and iodine
- (d) Silicon and oxygen
- (e) Phosphorus and fluorine
- (f) Element 71 and fluorine

Question 3.33 In the modern periodic table, the period indicates the value of :

- (a) atomic number
- (b) atomic mass
- (c) principal quantum number
- (d) azimuthal quantum number.

Question 3.34 Which of the following statements related to the modern periodic table is incorrect?

- (a) The p-block has 6 columns, because a maximum of 6 electrons can occupy all the orbitals in a p-shell.
- (b) The d-block has 8 columns, because a maximum of 8 electrons can occupy all the orbitals in a d-subshell.
- (c) Each block contains a number of columns equal to the number of electrons that can occupy that subshell.
- (d) The block indicates value of azimuthal quantum number (l) for the last subshell that received electrons in building up the electronic configuration.

Question 3.35 Anything that influences the valence electrons will affect the chemistry of the element. Which one of the following factors does not affect the valence shell?

- (a) Valence principal quantum number (n)
- (b) Nuclear charge (Z)
- (c) Nuclear mass
- (d) Number of core electrons.

Question 3.36 The size of isoelectronic species — F^- , Ne and Na^+ is affected by

- (a) nuclear charge (Z)
- (b) valence principal quantum number (n)
- (c) electron-electron interaction in the outer orbitals
- (d) none of the factors because their size is the same.

Question 3.37 Which one of the following statements is incorrect in relation to ionization enthalpy?

- (a) Ionization enthalpy increases for each successive electron.
- (b) The greatest increase in ionization enthalpy is experienced on removal of electron from core noble gas configuration.
- (c) End of valence electrons is marked by a big jump in ionization enthalpy.
- (d) Removal of electron from orbitals bearing lower n value is easier than from orbital having higher

n value.

Question 3.38 Considering the elements B, Al, Mg, and K, the correct order of their metallic character is :

- (a) $B > Al > Mg > K$
- (b) $Al > Mg > B > K$
- (c) $Mg > Al > K > B$
- (d) $K > Mg > Al > B$

Question 3.39 Considering the elements B, C, N, F, and Si, the correct order of their non-metallic character is :

- (a) $B > C > Si > N > F$
- (b) $Si > C > B > N > F$
- (c) $F > N > C > B > Si$
- (d) $F > N > C > Si > B$

Question 3.40 Considering the elements F, Cl, O and N, the correct order of their chemical reactivity in terms of oxidizing property is :

- (a) $F > Cl > O > N$
- (b) $F > O > Cl > N$
- (c) $Cl > F > O > N$
- (d) $O > F > N > Cl$

(Chemistry) Chapter 4 Chemical Bonding and Molecular Structure

NCERT Exercises Questions

Question 4.1 Explain the formation of a chemical bond.

Question 4.2 Write Lewis dot symbols for atoms of the following elements :
Mg, Na, B, O, N, Br.

Question 4.3 Write Lewis symbols for the following atoms and ions:

S and S^{2-} ; Al and Al^{3+} ; H and H^{-}

Question 4.4 Draw the Lewis structures for the following molecules and ions :

H_2S , $SiCl_4$, BeF_2 , CO_3^{2-} , $HCOOH$

Question 4.5 Define octet rule. Write its significance and limitations.

Question 4.6 Write the favourable factors for the formation of ionic bond.

Question 4.7 Discuss the shape of the following molecules using the VSEPR model:

$BeCl_2$, BCl_3 , $SiCl_4$, AsF_5 , H_2S , PH_3

Question 4.8 Although geometries of NH_3 and H_2O molecules are distorted tetrahedral, bond angle in water is less than that of ammonia. Discuss.

Question 4.9 How do you express the bond strength in terms of bond order ?

Question 4.10 Define the bond length.

Question 4.11 Explain the important aspects of resonance with reference to the CO_3^{2-} ion.

Question 4.12 H_3PO_3 can be represented by structures 1 and 2 shown below. Can these two structures be taken as the canonical forms of the resonance hybrid representing H_3PO_3 ? If not, give reasons for the same.

Question 4.13 Write the resonance structures for SO_3 , NO_2 and $3NO^{-}$.

Question 4.14 Use Lewis symbols to show electron transfer between the following atoms to form cations and anions : (a) K and (b) Ca and O (c) Al and N.

Question 4.15 Although both CO_2 and H_2O are triatomic molecules, the shape of H_2O molecule is bent while that of CO_2 is linear. Explain this on the basis of dipole moment. 4.16 Write the significance/applications of dipole moment.

Question 4.17 Define electronegativity. How does it differ from electron gain enthalpy ?

Question 4.18 Explain with the help of suitable example polar covalent bond.

Question 4.19 Arrange the bonds in order of increasing ionic character in the molecules: LiF, K₂O, N₂, SO₂ and ClF₃.

Question 4.20 The skeletal structure of CH₃COOH as shown below is correct, but some of the bonds are shown incorrectly. Write the correct Lewis structure for acetic acid.

Question 4.21 Apart from tetrahedral geometry, another possible geometry for CH₄ is square planar with the four H atoms at the corners of the square and the C atom at its centre. Explain why CH₄ is not square planar ?

Question 4.22 Explain why BeH₂ molecule has a zero dipole moment although the Be–H bonds are polar.

Question 4.23 Which out of NH₃ and NF₃ has higher dipole moment and why ?

Question 4.24 What is meant by hybridisation of atomic orbitals? Describe the shapes of sp, sp², sp³ hybrid orbitals.

Question 4.25 Describe the change in hybridisation (if any) of the Al atom in the following reaction.
 $3 \text{AlCl}_3 + \text{Cl}^- \rightarrow \text{Al}$

Question 4.26 Is there any change in the hybridisation of B and N atoms as a result of the following reaction ? $3 \text{BF}_3 + \text{NH}_3 \rightarrow \text{F}_3\text{B-NH}_2$

Question 4.27 Draw diagrams showing the formation of a double bond and a triple bond between carbon atoms in C₂H₄ and C₂H₂ molecules.

Question 4.28 What is the total number of sigma and pi bonds in the following molecules ?

- (a) C₂H₂
- (b) C₂H₄

Question 4.29 Considering x-axis as the internuclear axis which out of the following will not form a sigma bond and why?

- (a) 1s and 1s
- (b) 1s and 2p_x ;

- (c) $2p_y$ and $2p_z$
- (d) $1s$ and $2s$.

Question 4.30 Which hybrid orbitals are used by carbon atoms in the following molecules ?

- (a) $\text{CH}_3\text{-CH}_3$;
- (b) $\text{CH}_3\text{-CH=CH}_2$;
- (c) $\text{CH}_3\text{-CH}_2\text{-OH}$;
- (d) $\text{CH}_3\text{-CHO}$
- (e) CH_3COOH

Question 4.31 What do you understand by bond pairs and lone pairs of electrons ? Illustrate by giving one example of each type.

Question 4.32 Distinguish between a sigma and a pi bond.

Question 4.33 Explain the formation of H_2 molecule on the basis of valence bond theory.

Question 4.34 Write the important conditions required for the linear combination of atomic orbitals to form molecular orbitals.

Question 4.35 Use molecular orbital theory to explain why the Be_2 molecule does not exist. 4.36 Compare the relative stability of the following species and indicate their magnetic properties; $2 2 2 \text{O}^+$, $2 2 2 \text{O}^-$ (superoxide), $2 2 \text{O}^-$ (peroxide)

Question 4.37 Write the significance of a plus and a minus sign shown in representing the orbitals.

Question 4.38 Describe the hybridisation in case of PCl_5 . Why are the axial bonds longer as compared to equatorial bonds ?

Question 4.39 Define hydrogen bond. Is it weaker or stronger than the van der Waals forces?

Question 4.40 What is meant by the term bond order ? Calculate the bond order of : N_2 , O_2 , O_2^+ and O_2^-

(Chemistry) Chapter 5 States Of Matter

NCERT Solutions Questions

Question 5.1 What will be the minimum pressure required to compress 500 dm³ of air at 1 bar to 200 dm³ at 30°C? 152 C:\ChemistryXI\Unit-5\Unit-5(4)-Lay-2.pmd 14.1.6 (Final), 17.1.6, 24.1.6

Question 5.2 A vessel of 120 mL capacity contains a certain amount of gas at 35 °C and 1.2 bar pressure. The gas is transferred to another vessel of volume 180 mL at 35 °C. What would be its pressure?

Question 5.3 Using the equation of state $pV=nRT$; show that at a given temperature density of a gas is proportional to gas pressure p .

Question 5.4 At 0°C, the density of a certain oxide of a gas at 2 bar is same as that of dinitrogen at 5 bar. What is the molecular mass of the oxide?

Question 5.5 Pressure of 1 g of an ideal gas A at 27 °C is found to be 2 bar. When 2 g of another ideal gas B is introduced in the same flask at same temperature the pressure becomes 3 bar. Find a relationship between their molecular masses.

Question 5.6 The drain cleaner, Drainex contains small bits of aluminum which react with caustic soda to produce dihydrogen. What volume of dihydrogen at 20 °C and one bar will be released when 0.15g of aluminum reacts?

Question 5.7 What will be the pressure exerted by a mixture of 3.2 g of methane and 4.4 g of carbon dioxide contained in a 9 dm³ flask at 27 °C ?

Question 5.8 What will be the pressure of the gaseous mixture when 0.5 L of H₂ at 0.8 bar and 2.0 L of dioxygen at 0.7 bar are introduced in a 1L vessel at 27°C?

Question 5.9 Density of a gas is found to be 5.46 g/dm³ at 27 °C at 2 bar pressure. What will be its density at STP?

Question 5.10 34.05 mL of phosphorus vapour weighs 0.0625 g at 546 °C and 0.1 bar pressure. What is the molar mass of phosphorus?

Question 5.11 A student forgot to add the reaction mixture to the round bottomed flask at 27 °C but instead he/she placed the flask on the flame. After a lapse of time, he realized his mistake, and using a pyrometer he found the temperature of the flask was 477 °C. What fraction of air would have been expelled out?

Question 5.12 Calculate the temperature of 4.0 mol of a gas occupying 5 dm³ at 3.32 bar. (R = 0.083 bar dm³ K⁻¹ mol⁻¹).

Question 5.13 Calculate the total number of electrons present in 1.4 g of dinitrogen gas.

Question 5.14 How much time would it take to distribute one Avogadro number of wheat grains, if 1010 grains are distributed each second ?

Question 5.15 Calculate the total pressure in a mixture of 8 g of dioxygen and 4 g of dihydrogen confined in a vessel of 1 dm³ at 27°C. R = 0.083 bar dm³ K⁻¹ mol⁻¹.

Question 5.16 Pay load is defined as the difference between the mass of displaced air and the mass of the balloon. Calculate the pay load when a balloon of radius 10 m, mass 100 kg is filled with helium at 1.66 bar at 27°C. (Density of air = 1.2 kg m⁻³ and R = 0.083 bar dm³ K⁻¹ mol⁻¹).

Question 5.17 Calculate the volume occupied by 8.8 g of CO₂ at 31.1°C and 1 bar pressure. R = 0.083 bar L K⁻¹ mol⁻¹.

Question 5.18 2.9 g of a gas at 95 °C occupied the same volume as 0.184 g of dihydrogen at 17 °C, at the same pressure. What is the molar mass of the gas?

Question 5.19 A mixture of dihydrogen and dioxygen at one bar pressure contains 20% by weight of dihydrogen. Calculate the partial pressure of dihydrogen.

Question 5.20 What would be the SI unit for the quantity $pV^{2/3}n$?

Question 5.21 In terms of Charles' law explain why -273 °C is the lowest possible temperature.

Question 5.22 Critical temperature for carbon dioxide and methane are 31.1 °C and -81.9 °C respectively. Which of these has stronger intermolecular forces and why?

Question 5.23 Explain the physical significance of van der Waals parameters.

(Chemistry) Chapter 6 Thermodynamics

NCERT Exercises Questions

Question 6. 1 Choose the correct answer. A thermodynamic state function is a quantity

- (i) used to determine heat changes
- (ii) whose value is independent of path
- (iii) used to determine pressure volume work
- (iv) whose value depends on temperature only.

Question 6. 2 For the process to occur under adiabatic conditions, the correct condition is:

- (i) $\Delta T = 0$
- (ii) $\Delta p = 0$
- (iii) $q = 0$
- (iv) $w = 0$

Question 6. 3 The enthalpies of all elements in their standard states are:

- (i) unity
- (ii) zero
- (iii) < 0
- (iv) different for each element

Question 6. 4 ΔU_0 of combustion of methane is $-X \text{ kJ mol}^{-1}$. The value of ΔH_0 is

- (i) $= \Delta U_0$
- (ii) $> \Delta U_0$
- (iii) $< \Delta U_0$
- (iv) $= 0$

Question 6. 5 The enthalpy of combustion of methane, graphite and dihydrogen at 298 K are, $-890.3 \text{ kJ mol}^{-1}$, $-393.5 \text{ kJ mol}^{-1}$, and $-285.8 \text{ kJ mol}^{-1}$ respectively. Enthalpy of formation of $\text{CH}_4(\text{g})$ will be

- (i) $-74.8 \text{ kJ mol}^{-1}$
- (ii) $-52.27 \text{ kJ mol}^{-1}$
- (iii) $+74.8 \text{ kJ mol}^{-1}$
- (iv) $+52.26 \text{ kJ mol}^{-1}$.

Question 6. 6 A reaction, $A + B \rightarrow C + D + q$ is found to have a positive entropy change. The reaction will be

- (i) possible at high temperature
- (ii) possible only at low temperature
- (iii) not possible at any temperature
- (v) possible at any temperature

Question 6. 7 In a process, 701 J of heat is absorbed by a system and 394 J of work is done by the system. What is the change in internal energy for the process?

Question 6. 8 The reaction of cyanamide, NH_2CN (s), with dioxygen was carried out in a bomb calorimeter, and ΔU was found to be $-742.7 \text{ kJ mol}^{-1}$ at 298 K. Calculate enthalpy change for the reaction at 298 K. $\text{NH}_2\text{CN}(\text{g}) + 3/2 \text{ O}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$

Question 6. 9 Calculate the number of kJ of heat necessary to raise the temperature of 60.0 g of aluminium from 35°C to 55°C . Molar heat capacity of Al is $24 \text{ J mol}^{-1} \text{ K}^{-1}$.

Question 6. 10 Calculate the enthalpy change on freezing of 1.0 mol of water at 10.0°C to ice at -10.0°C . $\Delta_{\text{fus}}H =$

Question 6. 03 kJ mol^{-1} at 0°C . $C_p [\text{H}_2\text{O}(\text{l})] = 75.3 \text{ J mol}^{-1} \text{ K}^{-1}$ $C_p [\text{H}_2\text{O}(\text{s})] = 3$
 $\text{mol}^{-1} \text{ K}^{-1}$

Question 6. 11 Enthalpy of combustion of carbon to CO_2 is $-393.5 \text{ kJ mol}^{-1}$. Calculate the heat released upon formation of 35.2 g of CO_2 from carbon and dioxygen gas.

Question 6. 12 Enthalpies of formation of $\text{CO}(\text{g})$, $\text{CO}_2(\text{g})$, $\text{N}_2\text{O}(\text{g})$ and $\text{N}_2\text{O}_4(\text{g})$ are -110 , -393 , 81 and 9.7 kJ mol^{-1} respectively. Find the value of $\Delta_r H$ for the reaction:
 $\text{N}_2\text{O}_4(\text{g}) + 3\text{CO}(\text{g}) \rightarrow \text{N}_2\text{O}(\text{g}) + 3\text{CO}_2(\text{g})$

Question 6. 13 Given $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightarrow 2\text{NH}_3(\text{g})$; $\Delta_r H^\circ = -92.4 \text{ kJ mol}^{-1}$ What is the standard

enthalpy of formation of NH₃ gas?

Question 6. 14 Calculate the standard enthalpy of formation of CH₃OH(l) from the following data:
CH₃OH (l) + 3/2 O₂(g) → CO₂(g) + 2H₂O(l) ; Δ_rH⁰ = -726 kJ mol⁻¹
C(g) + O₂(g) → CO₂(g) ; Δ_cH⁰ = -393 kJ mol⁻¹
H₂(g) + 1/2 O₂(g) → H₂O(l) ; Δ_fH⁰ = -286 kJ mol⁻¹.

Question 6. 15 Calculate the enthalpy change for the process CCl₄(g) → C(g) + 4 Cl(g) and calculate bond enthalpy of C – Cl in CCl₄(g). Δ_{vap}H⁰(CCl₄) = 30.5 kJ mol⁻¹. Δ_fH⁰(CCl₄) = -135.5 kJ mol⁻¹. Δ_aH⁰(C) = 715.0 kJ mol⁻¹ , where Δ_aH⁰ is enthalpy of atomisation Δ_aH⁰(Cl₂) = 242 kJ mol⁻¹
Question 6. 16 For an isolated system, ΔU = 0, what will be ΔS ?

Question 6. 17 For the reaction at 298 K, 2A + B → C ΔH = 400 kJ mol⁻¹ and ΔS = 0.2 kJ K⁻¹ mol⁻¹ At what temperature will the reaction become spontaneous considering ΔH and ΔS to be constant over the temperature range.

Question 6. 18 For the reaction, 2 Cl(g) → Cl₂(g), what are the signs of ΔH and ΔS ?

Question 6. 19 For the reaction 2 A(g) + B(g) → 2D(g) ΔU⁰ = -10.5 kJ and ΔS⁰ = -44.1 JK⁻¹. Calculate ΔG⁰ for the reaction, and predict whether the reaction may occur spontaneously.

Question 6. 20 The equilibrium constant for a reaction is 10. What will be the value of ΔG⁰ ? R = 8.314 JK⁻¹ mol⁻¹, T = 300 K.

Question 6. 21 Comment on the thermodynamic stability of NO(g), given 1/2 N₂(g) + 1/2 O₂(g) → NO(g) ; Δ_rH⁰ = 90 kJ mol⁻¹
NO(g) + 1/2 O₂(g) → NO₂(g) : Δ_rH⁰ = -74 kJ mol⁻¹

Question 6. 22 Calculate the entropy change in surroundings when 1.00 mol of H₂O(l) is formed under standard conditions. Δ_fH⁰ = -286 kJ mol⁻¹.

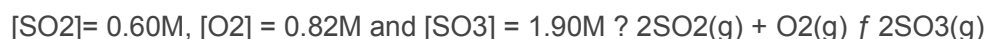
(Chemistry) Chapter 7 Equilibrium

NCERT Exercises Questions

Question 7.1 A liquid is in equilibrium with its vapour in a sealed container at a fixed temperature. The volume of the container is suddenly increased.

- What is the initial effect of the change on vapour pressure?
- How do rates of evaporation and condensation change initially?
- What happens when equilibrium is restored finally and what will be the final vapour pressure?

Question 7.2 What is K_c for the following equilibrium when the equilibrium concentration of each substance is:



Question 7.3 At a certain temperature and total pressure of 105Pa, iodine vapour contains 40% by volume of I atoms $I_2(g) \rightleftharpoons 2I(g)$ Calculate K_p for the equilibrium.

Question 7.4 Write the expression for the equilibrium constant, K_c for each of the following reactions:

- $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$
- $2Cu(NO_3)_2(s) \rightleftharpoons 2CuO(s) + 4NO_2(g) + O_2(g)$
- $CH_3COOC_2H_5(aq) + H_2O(l) \rightleftharpoons CH_3COOH(aq) + C_2H_5OH(aq)$
- $Fe^{3+}(aq) + 3OH^-(aq) \rightleftharpoons Fe(OH)_3(s)$ (v) $I_2(s) + 5F_2 \rightleftharpoons 2IF_5$

Question 7.5 Find out the value of K_c for each of the following equilibria from the value of K_p :

- $2NOCl(g) \rightleftharpoons 2NO(g) + Cl_2(g)$; $K_p = 1.8 \times 10^{-2}$ at 500 K
- $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$; $K_p = 167$ at 1073 K

Question 7.6 For the following equilibrium, $K_c = 6.3 \times 10^{14}$ at 1000 K $NO(g) + O_3(g) \rightleftharpoons NO_2(g) + O_2(g)$ Both the forward and reverse reactions in the equilibrium are elementary bimolecular reactions. What is K_c , for the reverse reaction?

Question 7.7 Explain why pure liquids and solids can be ignored while writing the equilibrium constant expression?

Question 7.8 Reaction between N_2 and O_2 takes place as follows: $2N_2(g) + O_2(g) \rightleftharpoons 2N_2O(g)$ If a mixture of 0.482 mol N_2 and 0.933 mol of O_2 is placed in a 10 L reaction vessel and allowed to form N_2O at a temperature for which $K_c = 2.0 \times 10^{-37}$, determine the composition of equilibrium mixture. \

Question 7.9 Nitric oxide reacts with Br₂ and gives nitrosyl bromide as per reaction given below:

$2\text{NO}(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{NOBr}(\text{g})$ When 0.087 mol of NO and 0.0437 mol of Br₂ are mixed in a closed container at constant temperature, 0.0518 mol of NOBr is obtained at equilibrium. Calculate equilibrium amount of NO and Br₂.

Question 7.10 At 450K, $K_p = 2.0 \times 10^{10}/\text{bar}$ for the given reaction at equilibrium. $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ What is K_c at this temperature ?

Question 7.11 A sample of HI(g) is placed in flask at a pressure of 0.2 atm. At equilibrium the partial pressure of HI(g) is 0.04 atm. What is K_p for the given equilibrium ? $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$

Question 7.12 A mixture of 1.57 mol of N₂, 1.92 mol of H₂ and 8.13 mol of NH₃ is introduced into a 20 L reaction vessel at 500 K. At this temperature, the equilibrium constant, K_c for the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ is 1.7×10^2 . Is the reaction mixture at equilibrium? If not, what is the direction of the net reaction?

Question 7.13 The equilibrium constant expression for a gas reaction is, $K = \frac{[\text{NH}_3]^2}{[\text{N}_2]^{1/4} [\text{H}_2]^{3/2}}$ Write the balanced chemical equation corresponding to this expression.

Question 7.14 One mole of H₂O and one mole of CO are taken in 10 L vessel and heated to 725 K. At equilibrium 40% of water (by mass) reacts with CO according to the equation, $\text{H}_2\text{O}(\text{g}) + \text{CO}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{CO}_2(\text{g})$ Calculate the equilibrium constant for the reaction.

Question 7.15 At 700 K, equilibrium constant for the reaction: $\text{H}_2(\text{g}) + \text{I}_2(\text{g}) \rightleftharpoons 2\text{HI}(\text{g})$ is 54.8. If 0.5 mol L⁻¹ of HI(g) is present at equilibrium at 700 K, what are the concentration of H₂(g) and I₂(g) assuming that we initially started with HI(g) and allowed it to reach equilibrium at 700K?

Question 7.16 What is the equilibrium concentration of each of the substances in the equilibrium when the initial concentration of ICl was 0.78 M ? $2\text{ICl}(\text{g}) \rightleftharpoons \text{I}_2(\text{g}) + \text{Cl}_2(\text{g}); K_c = 0.14$

Question 7.17 $K_p = 0.04 \text{ atm}$ at 899 K for the equilibrium shown below. What is the equilibrium concentration of C₂H₆ when it is placed in a flask at 4.0 atm pressure and allowed to come to equilibrium? $\text{C}_2\text{H}_6(\text{g}) \rightleftharpoons \text{C}_2\text{H}_4(\text{g}) + \text{H}_2(\text{g})$

Question 7.18 Ethyl acetate is formed by the reaction between ethanol and acetic acid and the equilibrium is represented as: $\text{CH}_3\text{COOH}(\text{l}) + \text{C}_2\text{H}_5\text{OH}(\text{l}) \rightleftharpoons \text{CH}_3\text{COOC}_2\text{H}_5(\text{l}) + \text{H}_2\text{O}(\text{l})$

(i) Write the concentration ratio (reaction quotient), Q_c , for this reaction (note: water is not in excess and is not a solvent in this reaction)

(ii) At 293 K, if one starts with 1.00 mol of acetic acid and 0.18 mol of ethanol, there is 0.171 mol of ethyl acetate in the final equilibrium mixture. Calculate the equilibrium constant.

(iii) Starting with 0.5 mol of ethanol and 1.0 mol of acetic acid and maintaining it at 293 K, 0.214 mol of ethyl acetate is found after sometime. Has equilibrium been reached?

Question 7.19 A sample of pure PCl_5 was introduced into an evacuated vessel at 473 K. After equilibrium was attained, concentration of PCl_5 was found to be $0.5 \times 10^{-1} \text{ mol L}^{-1}$. If value of K_c is 8.3×10^{-3} , what are the concentrations of PCl_3 and Cl_2 at equilibrium? $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$

Question 7.20 One of the reaction that takes place in producing steel from iron ore is the reduction of iron(II) oxide by carbon monoxide to give iron metal and CO_2 . $\text{FeO}(\text{s}) + \text{CO}(\text{g}) \rightleftharpoons \text{Fe}(\text{s}) + \text{CO}_2(\text{g})$; $K_p = 0.265 \text{ atm}$ at 1050K What are the equilibrium partial pressures of CO and CO_2 at 1050 K if the initial partial pressures are: $p_{\text{CO}} = 1.4 \text{ atm}$ and $p_{\text{CO}_2} = 0.80 \text{ atm}$?

Question 7.21 Equilibrium constant, K_c for the reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ at 500 K is 0.061 At a particular time, the analysis shows that composition of the reaction mixture is $3.0 \text{ mol L}^{-1} \text{ N}_2$, $2.0 \text{ mol L}^{-1} \text{ H}_2$ and $0.5 \text{ mol L}^{-1} \text{ NH}_3$. Is the reaction at equilibrium? If not in which direction does the reaction tend to proceed to reach equilibrium?

Question 7.22 Bromine monochloride, BrCl decomposes into bromine and chlorine and reaches the equilibrium: $2\text{BrCl}(\text{g}) \rightleftharpoons \text{Br}_2(\text{g}) + \text{Cl}_2(\text{g})$ for which $K_c = 32$ at 500 K. If initially pure BrCl is present at a concentration of $3.3 \times 10^{-3} \text{ mol L}^{-1}$, what is its molar concentration in the mixture at equilibrium?

Question 7.23 At 1127 K and 1 atm pressure, a gaseous mixture of CO and CO_2 in equilibrium with solid carbon has 90.55% CO by mass $\text{C}(\text{s}) + \text{CO}_2(\text{g}) \rightleftharpoons 2\text{CO}(\text{g})$ Calculate K_c for this reaction at the above temperature.

Question 7.24 Calculate a) ΔG° and b) the equilibrium constant for the formation of NO_2 from NO and O_2 at 298K $\text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightleftharpoons \text{NO}_2(\text{g})$ where $\Delta_f G^\circ(\text{NO}_2) = 52.0 \text{ kJ/mol}$ $\Delta_f G^\circ(\text{NO}) = 87.0 \text{ kJ/mol}$ $\Delta_f G^\circ(\text{O}_2) = 0 \text{ kJ/mol}$

Question 7.25 Does the number of moles of reaction products increase, decrease or remain same

when each of the following equilibria is subjected to a decrease in pressure by increasing the volume?

- (a) $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$
- (b) $\text{CaO}(\text{s}) + \text{CO}_2(\text{g}) \rightleftharpoons \text{CaCO}_3(\text{s})$
- (c) $3\text{Fe}(\text{s}) + 4\text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{Fe}_3\text{O}_4(\text{s}) + 4\text{H}_2(\text{g})$

Question 7.26 Which of the following reactions will get affected by increasing the pressure? Also, mention whether change will cause the reaction to go into forward or backward direction.

- (i) $\text{COCl}_2(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + \text{Cl}_2(\text{g})$
- (ii) $\text{CH}_4(\text{g}) + 2\text{S}_2(\text{g}) \rightleftharpoons \text{CS}_2(\text{g}) + 2\text{H}_2\text{S}(\text{g})$
- (iii) $\text{CO}_2(\text{g}) + \text{C}(\text{s}) \rightleftharpoons 2\text{CO}(\text{g})$ (iv) $2\text{H}_2(\text{g}) + \text{CO}(\text{g}) \rightleftharpoons \text{CH}_3\text{OH}(\text{g})$
- (v) $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
- (vi) $4\text{NH}_3(\text{g}) + 5\text{O}_2(\text{g}) \rightleftharpoons 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g})$

Question 7.27 The equilibrium constant for the following reaction is 1.6×10^5 at 1024K $\text{H}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{HBr}(\text{g})$ Find the equilibrium pressure of all gases if 10.0 bar of HBr is introduced into a sealed container at 1024K.

Question 28 Dihydrogen gas is obtained from natural gas by partial oxidation with steam as per following endothermic reaction: $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons \text{CO}(\text{g}) + 3\text{H}_2(\text{g})$

- (a) Write an expression for K_p for the above reaction.
- (b) How will the values of K_p and composition of equilibrium mixture be affected by
 - (i) increasing the pressure
 - (ii) increasing the temperature
 - (iii) using a catalyst ?

Question 7.29 Describe the effect of : a) addition of H_2 b) addition of CH_3OH c) removal of CO d) removal of CH_3OH on the equilibrium of the reaction:



Question 7.30 At 473 K, equilibrium constant K_c for decomposition of phosphorus pentachloride, PCl_5 is 8.3×10^{-3} . If decomposition is depicted as, $\text{PCl}_5(\text{g}) \rightleftharpoons \text{PCl}_3(\text{g}) + \text{Cl}_2(\text{g})$ $\Delta_r H^\circ = 124.0 \text{ kJ mol}^{-1}$

- a) write an expression for K_c for the reaction.
- b) what is the value of K_c for the reverse reaction at the same temperature ?
- c) what would be the effect on K_c if

- (i) more PCl_5 is added
- (ii) pressure is increased
- (iii) the temperature is increased ?

Question 7.31 Dihydrogen gas used in Haber's process is produced by reacting methane from natural gas with high temperature steam. The first stage of two stage reaction involves the formation of CO and H_2 . In second stage, CO formed in first stage is reacted with more steam in water gas shift reaction, $\text{CO (g)} + \text{H}_2\text{O (g)} \rightleftharpoons \text{CO}_2 \text{ (g)} + \text{H}_2 \text{ (g)}$ If a reaction vessel at 400°C is charged with an equimolar mixture of CO and steam such that $\text{CO} = \text{H}_2\text{O} = p = 4.0 \text{ bar}$, what will be the partial pressure of H_2 at equilibrium? $K_p = 10.1$ at 400°C

Question 7.32 Predict which of the following reaction will have appreciable concentration of reactants and products: a) $\text{Cl}_2 \text{ (g)} \rightleftharpoons 2\text{Cl (g)}$ $K_c = 5 \times 10^{-39}$ b) $\text{Cl}_2 \text{ (g)} + 2\text{NO (g)} \rightleftharpoons 2\text{NOCl (g)}$ $K_c = 3.7 \times 10^8$ c) $\text{Cl}_2 \text{ (g)} + 2\text{NO}_2 \text{ (g)} \rightleftharpoons 2\text{NO}_2\text{Cl (g)}$ $K_c = 1.8$

Question 7.33 The value of K_c for the reaction $3\text{O}_2 \text{ (g)} \rightleftharpoons 2\text{O}_3 \text{ (g)}$ is 2.0×10^{-50} at 25°C . If the equilibrium concentration of O_2 in air at 25°C is 1.6×10^{-2} , what is the concentration of O_3 ?

Question 7.34 The reaction, $\text{CO(g)} + 3\text{H}_2\text{(g)} \rightleftharpoons \text{CH}_4\text{(g)} + \text{H}_2\text{O(g)}$ is at equilibrium at 1300 K in a 1 L flask. It also contain 0.30 mol of CO , 0.10 mol of H_2 and 0.02 mol of H_2O and an unknown amount of CH_4 in the flask. Determine the concentration of CH_4 in the mixture. The equilibrium constant, K_c for the reaction at the given temperature is 3.90 .

Question 7.35 What is meant by the conjugate acid-base pair? Find the conjugate acid/base for the following species: HNO_2 , CN^- , HClO_4 , F^- , OH^- , CO_3^{2-} , and S^{2-}

Question 7.36 Which of the followings are Lewis acids? H_2O , BF_3 , H^+ , and NH_4^+ 7.37 What will be the conjugate bases for the Brønsted acids: HF , H_2SO_4 and HCO_3^- ? 7.38 Write the conjugate acids for the following Brønsted bases: NH_2^- , NH_3 and HCOO^- .

Question 7.39 The species: H_2O , HCO_3^- , HSO_4^- and NH_3 can act both as Brønsted acids and bases. For each case give the corresponding conjugate acid and base.

Question 7.40 Classify the following species into Lewis acids and Lewis bases and show how these act as Lewis acid/base:

- (a) OH^-

- (b) F⁻
- (c) H⁺
- (d) BCl₃ .

Question 7.41 The concentration of hydrogen ion in a sample of soft drink is 3.8×10^{-3} M. what is its pH?

Question 7.42 The pH of a sample of vinegar is 3.76. Calculate the concentration of hydrogen ion in it.

Question 7.43 The ionization constant of HF, HCOOH and HCN at 298K are 6.8×10^{-4} , 1.8×10^{-4} and 4.8×10^{-9} respectively. Calculate the ionization constants of the corresponding conjugate base.

Question 7.44 The ionization constant of phenol is 1.0×10^{-10} . What is the concentration of phenolate ion in 0.05 M solution of phenol? What will be its degree of ionization if the solution is also 0.01M in sodium phenolate?

Question 7.45 The first ionization constant of H₂S is 9.1×10^{-8} . Calculate the concentration of HS⁻ ion in its 0.1M solution. How will this concentration be affected if the solution is 0.1M in HCl also ? If the second dissociation constant of H₂S is 1.2×10^{-13} , calculate the concentration of S²⁻ under both conditions.

Question 7.46 The ionization constant of acetic acid is 1.74×10^{-5} . Calculate the degree of dissociation of acetic acid in its 0.05 M solution. Calculate the concentration of acetate ion in the solution and its pH.

Question 7.47 It has been found that the pH of a 0.01M solution of an organic acid is 4.15. Calculate the concentration of the anion, the ionization constant of the acid and its pK_a .

Question 7.48 Assuming complete dissociation, calculate the pH of the following solutions:

- (a) 0.003 M HCl
- (b) 0.005 M NaOH
- (c) 0.002 M HBr
- (d) 0.002 M KOH

Question 7.49 Calculate the pH of the following solutions:

- a) 2 g of TIOH dissolved in water to give 2 litre of solution.
- b) 0.3 g of Ca(OH)₂ dissolved in water to give 500 mL of solution.
- c) 0.3 g of NaOH dissolved in water to give 200 mL of solution.
- d) 1mL of 13.6 M HCl is diluted with water to give 1 litre of solution.

Question 7.50 The degree of ionization of a 0.1M bromoacetic acid solution is 0.132. Calculate the pH of the solution and the pKa of bromoacetic acid.

Question 7.51 The pH of 0.005M codeine (C₁₈H₂₁NO₃) solution is 9.95. Calculate its ionization constant and pKb.

Question 7.52 What is the pH of 0.001M aniline solution ? The ionization constant of aniline can be taken from Table7. Calculate the degree of ionization of aniline in the solution. Also calculate the ionization constant of the conjugate acid of aniline.

Question 7.53 Calculate the degree of ionization of 0.05M acetic acid if its pKa value is 4.74. How is the degree of dissociation affected when its solution also contains (a) 0.01M (b) 0.1M in HCl ?

Question 7.54 The ionization constant of dimethylamine is 5.4×10^{-4} . Calculate its degree of ionization in its 0.02M solution. What percentage of dimethylamine is ionized if the solution is also 0.1M in NaOH?

Question 7.55 Calculate the hydrogen ion concentration in the following biological fluids whose pH are given below:

- (a) Human muscle-fluid, 6.83
- (b) Human stomach fluid, 1.2
- (c) Human blood, 7.38
- (d) Human saliva, 6.4.

Question 7.56 The pH of milk, black coffee, tomato juice, lemon juice and egg white are 6.8, 5.0, 4.2, 2.2 and 7.8 respectively. Calculate corresponding hydrogen ion concentration in each.

Question 7.57 If 0.561 g of KOH is dissolved in water to give 200 mL of solution at 298 K. Calculate the concentrations of potassium, hydrogen and hydroxyl ions. What is its pH?

Question 7.58 The solubility of Sr(OH)₂ at 298 K is 19.23 g/L of solution. Calculate the

concentrations of strontium and hydroxyl ions and the pH of the solution. \

Question 7.59 The ionization constant of propanoic acid is 1.32×10^{-5} . Calculate the degree of ionization of the acid in its 0.05M solution and also its pH. What will be its degree of ionization if the solution is 0.01M in HCl also?

Question 7.60 The pH of 0.1M solution of cyanic acid (HCNO) is 2.34. Calculate the ionization constant of the acid and its degree of ionization in the solution.

Question 7.61 The ionization constant of nitrous acid is 4.5×10^{-4} . Calculate the pH of 0.04 M sodium nitrite solution and also its degree of hydrolysis.

Question 7.62 A 0.02M solution of pyridinium hydrochloride has pH = 3.44. Calculate the ionization constant of pyridine.

Question 7.63 Predict if the solutions of the following salts are neutral, acidic or basic: NaCl, KBr, NaCN, NH_4NO_3 , NaNO_2 and KF

Question 7.64 The ionization constant of chloroacetic acid is 1.35×10^{-3} . What will be the pH of 0.1M acid and its 0.1M sodium salt solution?

Question 7.65 Ionic product of water at 310 K is 2.7×10^{-14} . What is the pH of neutral water at this temperature?

Question 7.66 Calculate the pH of the resultant mixtures:

- a) 10 mL of 0.2M $\text{Ca}(\text{OH})_2$ + 25 mL of 0.1M HCl
- b) 10 mL of 0.01M H_2SO_4 + 10 mL of 0.01M $\text{Ca}(\text{OH})_2$
- c) 10 mL of 0.1M H_2SO_4 + 10 mL of 0.1M KOH

Question 7.67 Determine the solubilities of silver chromate, barium chromate, ferric hydroxide, lead chloride and mercurous iodide at 298K from their solubility product constants given in Table Determine also the molarities of individual ions.

Question 7.68 The solubility product constant of Ag_2CrO_4 and AgBr are 1.1×10^{-12} and 5.0×10^{-13} respectively. Calculate the ratio of the molarities of their saturated solutions.

Question 7.69 Equal volumes of 0.002 M solutions of sodium iodate and cupric chlorate are mixed together. Will it lead to precipitation of copper iodate? (For cupric iodate $K_{sp} = 4 \times 10^{-8}$).

Question 7.70 The ionization constant of benzoic acid is 6.46×10^{-5} and K_{sp} for silver benzoate is 2.5×10^{-13} . How many times is silver benzoate more soluble in a buffer of pH 3.19 compared to its solubility in pure water?

Question 7.71 What is the maximum concentration of equimolar solutions of ferrous sulphate and sodium sulphide so that when mixed in equal volumes, there is no precipitation of iron sulphide? (For iron sulphide, $K_{sp} = 6.3 \times 10^{-18}$).

Question 7.72 What is the minimum volume of water required to dissolve 1g of calcium sulphate at 298 K? (For calcium sulphate, K_{sp} is 9.1×10^{-6}).

Question 7.73 The concentration of sulphide ion in 0.1M HCl solution saturated with hydrogen sulphide is 1.0×10^{-19} M. If 10 mL of this is added to 5 mL of 0.04 M solution of the following: $FeSO_4$, $MnCl_2$, $ZnCl_2$ and $CdCl_2$. in which of these solutions precipitation will take place?

(Chemistry) Chapter 8 Redox Reactions

NCERT Exercises Questions

Question 8. 1 Assign oxidation number to the underlined elements in each of the following species:

- (a) NaH_2PO_4
- (b) $NaHSO_4$
- (c) $H_4P_2O_7$
- (d) K_2MnO_4
- (e) CaO_2
- (f) $NaBH_4$
- (g) $H_2S_2O_7$
- (h) $KAl(SO_4)_2 \cdot 12 H_2O$

Question 8. 2 What are the oxidation number of the underlined elements in each of the following

and how do you rationalise your results ?

- (a) KI₃
- (b) H₂S₄O₆
- (c) Fe₃O₄
- (d) CH₃CH₂OH
- (e) CH₃COOH

Question 8. 3 Justify that the following reactions are redox reactions:

- (a) $\text{CuO(s)} + \text{H}_2\text{(g)} \rightarrow \text{Cu(s)} + \text{H}_2\text{O(g)}$
- (b) $\text{Fe}_2\text{O}_3\text{(s)} + 3\text{CO(g)} \rightarrow 2\text{Fe(s)} + 3\text{CO}_2\text{(g)}$
- (c) $4\text{BCl}_3\text{(g)} + 3\text{LiAlH}_4\text{(s)} \rightarrow 2\text{B}_2\text{H}_6\text{(g)} + 3\text{LiCl(s)} + 3\text{AlCl}_3\text{(s)}$
- (d) $2\text{K(s)} + \text{F}_2\text{(g)} \rightarrow 2\text{K}^+\text{F}^-\text{(s)}$ (e) $4\text{NH}_3\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 4\text{NO(g)} + 6\text{H}_2\text{O(g)}$

Question 8. 4 Fluorine reacts with ice and results in the change: $\text{H}_2\text{O(s)} + \text{F}_2\text{(g)} \rightarrow \text{HF(g)} + \text{HOF(g)}$
Justify that this reaction is a redox reaction.

Question 8. 5 Calculate the oxidation number of sulphur, chromium and nitrogen in H₂SO₅, Cr₂O₇²⁻ and NO₃⁻. Suggest structure of these compounds. Count for the fallacy.

Question 8. 6 Write formulas for the following compounds:

- (a) Mercury(II) chloride
- (b) Nickel(II) sulphate
- (c) Tin(IV) oxide
- (d) Thallium(I) sulphate
- (e) Iron(III) sulphate
- (f) Chromium(III) oxide

Question 8. 7 Suggest a list of the substances where carbon can exhibit oxidation states from -4 to +4 and nitrogen from -3 to +5.

Question 8. 8 While sulphur dioxide and hydrogen peroxide can act as oxidising as well as reducing agents in their reactions, ozone and nitric acid act only as oxidants. Why ?

Question 8. 9 Consider the reactions:

(a) $6 \text{CO}_2(\text{g}) + 6\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{O}_2(\text{g})$ (b) $\text{O}_3(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + 2\text{O}_2(\text{g})$ Why it is more appropriate to write these reactions as : (a) $6\text{CO}_2(\text{g}) + 12\text{H}_2\text{O}(\text{l}) \rightarrow \text{C}_6\text{H}_{12}\text{O}_6(\text{aq}) + 6\text{H}_2\text{O}(\text{l}) + 6\text{O}_2(\text{g})$ (b) $\text{O}_3(\text{g}) + \text{H}_2\text{O}_2(\text{l}) \rightarrow \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) + \text{O}_2(\text{g})$ Also suggest a technique to investigate the path of the above (a) and (b) redox reactions.

Question 8. 10 The compound AgF_2 is unstable compound. However, if formed, the compound acts as a very strong oxidising agent. Why ?

Question 8. 11 Whenever a reaction between an oxidising agent and a reducing agent is carried out, a compound of lower oxidation state is formed if the reducing agent is in excess and a compound of higher oxidation state is formed if the oxidising agent is in excess. Justify this statement giving three illustrations.

Question 8. 12 How do you count for the following observations ?

(a) Though alkaline potassium permanganate and acidic potassium permanganate both are used as oxidants, yet in the manufacture of benzoic acid from toluene we use alcoholic potassium permanganate as an oxidant. Why ? Write a balanced redox equation for the reaction.
(b) When concentrated sulphuric acid is added to an inorganic mixture containing chloride, we get colourless pungent smelling gas HCl , but if the mixture contains bromide then we get red vapour of bromine. Why ?

Question 8. 13 Identify the substance oxidised reduced, oxidising agent and reducing agent for each of the following reactions:

(a) $2\text{AgBr}(\text{s}) + \text{C}_6\text{H}_6\text{O}_2(\text{aq}) \rightarrow 2\text{Ag}(\text{s}) + 2\text{HBr}(\text{aq}) + \text{C}_6\text{H}_4\text{O}_2(\text{aq})$
(b) $\text{HCHO}(\text{l}) + 2[\text{Ag}(\text{NH}_3)_2]^+(\text{aq}) + 3\text{OH}^-(\text{aq}) \rightarrow 2\text{Ag}(\text{s}) + \text{HCOO}^-(\text{aq}) + 4\text{NH}_3(\text{aq}) + 2\text{H}_2\text{O}(\text{l})$
(c) $\text{HCHO}(\text{l}) + 2\text{Cu}^{2+}(\text{aq}) + 5\text{OH}^-(\text{aq}) \rightarrow \text{Cu}_2\text{O}(\text{s}) + \text{HCOO}^-(\text{aq}) + 3\text{H}_2\text{O}(\text{l})$
(d) $\text{N}_2\text{H}_4(\text{l}) + 2\text{H}_2\text{O}_2(\text{l}) \rightarrow \text{N}_2(\text{g}) + 4\text{H}_2\text{O}(\text{l})$ (e) $\text{Pb}(\text{s}) + \text{PbO}_2(\text{s}) + 2\text{H}_2\text{SO}_4(\text{aq}) \rightarrow 2\text{PbSO}_4(\text{s}) + 2\text{H}_2\text{O}(\text{l})$

Question 8. 14 Consider the reactions : $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{s}) \rightarrow \text{S}_4\text{O}_6^{2-}(\text{aq}) + 2\text{I}^-(\text{aq})$ $\text{S}_2\text{O}_3^{2-}(\text{aq}) + 2\text{Br}_2(\text{l}) + 5\text{H}_2\text{O}(\text{l}) \rightarrow 2\text{SO}_4^{2-}(\text{aq}) + 4\text{Br}^-(\text{aq}) + 10\text{H}^+(\text{aq})$ Why does the same reductant, thiosulphate react differently with iodine and bromine ?

Question 8. 15 Justify giving reactions that among halogens, fluorine is the best oxidant and among hydrohalic compounds, hydroiodic acid is the best reductant. What inference do you draw about the

behaviour of Ag^+ and Cu^{2+} from these reactions ?

Question 8. 18 Balance the following redox reactions by ion – electron method :

- (a) $\text{MnO}_4^- (\text{aq}) + \text{I}^- (\text{aq}) \rightarrow \text{MnO}_2 (\text{s}) + \text{I}_2 (\text{s})$ (in basic medium)
- (b) $\text{MnO}_4^- (\text{aq}) + \text{SO}_2 (\text{g}) \rightarrow \text{Mn}^{2+} (\text{aq}) + \text{HSO}_4^- (\text{aq})$ (in acidic solution)
- (c) $\text{H}_2\text{O}_2 (\text{aq}) + \text{Fe}^{2+} (\text{aq}) \rightarrow \text{Fe}^{3+} (\text{aq}) + \text{H}_2\text{O} (\text{l})$ (in acidic solution)
- (d) $\text{Cr}_2\text{O}_7^{2-} + \text{SO}_2 (\text{g}) \rightarrow \text{Cr}^{3+} (\text{aq}) + \text{SO}_4^{2-} (\text{aq})$ (in acidic solution)

Question 8. 19 Balance the following equations in basic medium by ion-electron method and oxidation number methods and identify the oxidising agent and the reducing agent.

- (a) $\text{P}_4 (\text{s}) + \text{OH}^- (\text{aq}) \rightarrow \text{PH}_3 (\text{g}) + \text{HPO}_2^- (\text{aq})$
- (b) $\text{N}_2\text{H}_4 (\text{l}) + \text{ClO}_3^- (\text{aq}) \rightarrow \text{NO} (\text{g}) + \text{Cl}^- (\text{g})$
- (c) $\text{Cl}_2\text{O}_7 (\text{g}) + \text{H}_2\text{O}_2 (\text{aq}) \rightarrow \text{ClO}_2^- (\text{aq}) + \text{O}_2 (\text{g}) + \text{H}^+$

Question 8. 20 What sorts of informations can you draw from the following reaction ? $(\text{CN})_2 (\text{g}) + 2\text{OH}^- (\text{aq}) \rightarrow \text{CN}^- (\text{aq}) + \text{CNO}^- (\text{aq}) + \text{H}_2\text{O} (\text{l})$

Question 8. 21 The Mn^{3+} ion is unstable in solution and undergoes disproportionation to give Mn^{2+} , MnO_2 , and H^+ ion. Write a balanced ionic equation for the reaction.

Question 8. 22 Consider the elements : Cs, Ne, I and F

- (a) Identify the element that exhibits only negative oxidation state.
- (b) Identify the element that exhibits only positive oxidation state.
- (c) Identify the element that exhibits both positive and negative oxidation states.
- (d) Identify the element which exhibits neither the negative nor does the positive oxidation state.

Question 8. 23 Chlorine is used to purify drinking water. Excess of chlorine is harmful. The excess of chlorine is removed by treating with sulphur dioxide. Present a balanced equation for this redox change taking place in water.

Question 8. 24 Refer to the periodic table given in your book and now answer the following questions:

- (a) Select the possible non metals that can show disproportionation reaction.
- (b) Select three metals that can show disproportionation reaction.

Question 8. 25 In Ostwald's process for the manufacture of nitric acid, the first step involves the

oxidation of ammonia gas by oxygen gas to give nitric oxide gas and steam. What is the maximum weight of nitric oxide that can be obtained starting only with 10.00 g. of ammonia and 20.00 g of oxygen ?

Question 8. 26 Using the standard electrode potentials given in the Table , predict if the reaction between the following is feasible:

- (a) $\text{Fe}^{3+}(\text{aq})$ and $\text{I}^{-}(\text{aq})$
- (b) $\text{Ag}^{+}(\text{aq})$ and $\text{Cu}(\text{s})$
- (c) $\text{Fe}^{3+}(\text{aq})$ and $\text{Cu}(\text{s})$
- (d) $\text{Ag}(\text{s})$ and $\text{Fe}^{3+}(\text{aq})$
- (e) $\text{Br}_2(\text{aq})$ and $\text{Fe}^{2+}(\text{aq})$.

Question 8. 27 Predict the products of electrolysis in each of the following:

- (i) An aqueous solution of AgNO_3 with silver electrodes
- (ii) An aqueous solution AgNO_3 with platinum electrodes
- (iii) A dilute solution of H_2SO_4 with platinum electrodes
- (iv) An aqueous solution of CuCl_2 with platinum electrodes.

Question 8. 28 Arrange the following metals in the order in which they displace each other from the solution of their salts. Al, Cu, Fe, Mg and Zn.

Question 8. 29 Given the standard electrode potentials, $\text{K}^{+}/\text{K} = -2.93\text{V}$, $\text{Ag}^{+}/\text{Ag} = 0.80\text{V}$, $\text{Hg}_2^{2+}/\text{Hg} = 0.79\text{V}$ $\text{Mg}^{2+}/\text{Mg} = -2.37\text{V}$. $\text{Cr}^{3+}/\text{Cr} = -0.74\text{V}$ arrange these metals in their increasing order of reducing power.

Question 8. 30 Depict the galvanic cell in which the reaction $\text{Zn}(\text{s}) + 2\text{Ag}^{+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{Ag}(\text{s})$ takes place, Further show:

- (i) which of the electrode is negatively charged,
- (ii) the carriers of the current in the cell, and
- (iii) individual reaction at each electrode.

Chapter 9 Hydrogen

Question 9.1 Justify the position of hydrogen in the periodic table on the basis of its electronic configuration.

Question 9.2 Write the names of isotopes of hydrogen. What is the mass ratio of these isotopes?

Question 9.3 Why does hydrogen occur in a diatomic form rather than in a monoatomic form under normal conditions?

Question 9.4 How can the production of dihydrogen, obtained from 'coal gasification', be increased ?

Question 9.5 Describe the bulk preparation of dihydrogen by electrolytic method. What is the role of an electrolyte in this process ?

Question 9.6 Complete the following reactions:



Question 9.7 Discuss the consequences of high enthalpy of H–H bond in terms of chemical reactivity of dihydrogen.

Question 9.8 What do you understand by (i) electron-deficient, (ii) electron-precise, and (iii) electron-rich compounds of hydrogen? Provide justification with suitable examples.

Question 9.9 What characteristics do you expect from an electron-deficient hydride with respect to its structure and chemical reactions?

Question 9.10 Do you expect the carbon hydrides of the type $(\text{C}_n\text{H}_{2n+2})$ to act as 'Lewis' acid or base? Justify your answer.

Question 9.11 What do you understand by the term "non-stoichiometric hydrides"? Do you expect this type of the hydrides to be formed by alkali metals? Justify your answer.

Question 9.12 How do you expect the metallic hydrides to be useful for hydrogen storage? Explain.

Question 9.13 How does the atomic hydrogen or oxy-hydrogen torch function for cutting and

welding purposes ? Explain.

Question 9.14 Among NH_3 , H_2O and HF , which would you expect to have highest magnitude of hydrogen bonding and why?

Question 9.15 Saline hydrides are known to react with water violently producing fire. Can CO_2 , a well known fire extinguisher, be used in this case? Explain.

Question 9.16 Arrange the following

- (i) CaH_2 , BeH_2 and TiH_2 in order of increasing electrical conductance.
- (ii) LiH , NaH and CsH in order of increasing ionic character.
- (iii) H-H , D-D and F-F in order of increasing bond dissociation enthalpy.
- (iv) NaH , MgH_2 and H_2O in order of increasing reducing property.

Question 9.17 Compare the structures of H_2O and H_2O_2 .

Question 9.18 What do you understand by the term 'auto-protolysis' of water? What is its significance?

Question 9.19 Consider the reaction of water with F_2 and suggest, in terms of oxidation and reduction, which species are oxidised/reduced.

Question 9.20 Complete the following chemical reactions.

- (i) $() () 2 2 \text{PbS s} + \text{H O aq} \rightarrow$
- (ii) $() () 4 2 2 \text{MnO aq} + \text{H O aq} \rightarrow$
- (iii) $() () 2 \text{CaO s} + \text{H O g} \rightarrow$
- (v) $() () 3 2 \text{AlCl g} + \text{H O l} \rightarrow$
- (vi) $() () 3 2 2 \text{Ca N s} + \text{H O l} \rightarrow$ Classify the above into
 - (a) hydrolysis,
 - (b) redox and
 - (c) hydration reactions.

Question 9.21 Describe the structure of the common form of ice.

Question 9.22 What causes the temporary and permanent hardness of water ?

Question 9.23 Discuss the principle and method of softening of hard water by synthetic ionexchange resins.

Question 9.24 Write chemical reactions to show the amphoteric nature of water.

Question 9.25 Write chemical reactions to justify that hydrogen peroxide can function as an oxidising as well as reducing agent.

Question 9.26 What is meant by 'demineralised' water and how can it be obtained ?

Question 9.27 Is demineralised or distilled water useful for drinking purposes? If not, how can it be made useful?

Question 9.28 Describe the usefulness of water in biosphere and biological systems.

Question 9.29 What properties of water make it useful as a solvent? What types of compound can it (i) dissolve, and (ii) hydrolyse ?

Question 9.30 Knowing the properties of H₂O and D₂O, do you think that D₂O can be used for drinking purposes?

Question 9.31 What is the difference between the terms 'hydrolysis' and 'hydration' ?

Question 9.32 How can saline hydrides remove traces of water from organic compounds?

Question 9.33 What do you expect the nature of hydrides is, if formed by elements of atomic numbers 15, 19, 23 and 44 with dihydrogen? Compare their behaviour towards water.

Question 9.34 Do you expect different products in solution when aluminium(III) chloride and potassium chloride treated separately with (i) normal water (ii) acidified water, and (iii) alkaline water? Write equations wherever necessary.

Question 9.35 How does H₂O₂ behave as a bleaching agent?

Question 9.36 What do you understand by the terms:

- (i) hydrogen economy
- (ii) hydrogenation

- (iii) 'syngas'
- (iv) water-gas shift reaction
- (v) fuel-cell ?

(Chemistry) Chapter 10 The S -Block Elements

NCERT Exercises Questions

Question 10.1 What are the common physical and chemical features of alkali metals ?

Question 10.2 Discuss the general characteristics and gradation in properties of alkaline earth metals.

Question 10.3 Why are alkali metals not found in nature ?

Question 10.4 Find out the oxidation state of sodium in Na_2O_2 . 5.5 Explain why is sodium less reactive than potassium.

Question 10.6 Compare the alkali metals and alkaline earth metals with respect to (i) ionisation enthalpy (ii) basicity of oxides and (iii) solubility of hydroxides.

Question 10.7 In what ways lithium shows similarities to magnesium in its chemical behaviour?

Question 10.8 Explain why can alkali and alkaline earth metals not be obtained by chemical reduction methods?

Question 10.9 Why are potassium and caesium, rather than lithium used in photoelectric cells?

Question 10.10 When an alkali metal dissolves in liquid ammonia the solution can acquire different colours. Explain the reasons for this type of colour change.

Question 10.11 Beryllium and magnesium do not give colour to flame whereas other alkaline earth metals do so. Why ?

Question 10.12 Discuss the various reactions that occur in the Solvay process.

Question 10.13 Potassium carbonate cannot be prepared by Solvay process. Why ?

Question 10.14 Why is Li_2CO_3 decomposed at a lower temperature whereas Na_2CO_3 at higher temperature?

Question 10.15 Compare the solubility and thermal stability of the following compounds of the alkali metals with those of the alkaline earth metals.

- (a) Nitrates
- (b) Carbonates
- (c) Sulphates.

Question 10.16 Starting with sodium chloride how would you proceed to prepare

- (i) sodium metal
- (ii) sodium hydroxide
- (iii) sodium peroxide
- (iv) sodium carbonate ?

Question 10.17 What happens when

- (i) magnesium is burnt in air
- (ii) quick lime is heated with silica
- (iii) chlorine reacts with slaked lime
- (iv) calcium nitrate is heated ?

Question 10.18 Describe two important uses of each of the following :

- (i) caustic soda
- (ii) sodium carbonate
- (iii) quicklime.

Question 10.19 Draw the structure of

- (i) BeCl_2 (vapour)
- (ii) BeCl_2 (solid).

Question 10.20 The hydroxides and carbonates of sodium and potassium are easily soluble in water while the corresponding salts of magnesium and calcium are sparingly soluble in water.

Explain.

Question 10.21 Describe the importance of the following :

- (i) limestone
- (ii) cement
- (iii) plaster of paris.

Question 10.22 Why are lithium salts commonly hydrated and those of the other alkali ions usually anhydrous?

Question 10.23 Why is LiF almost insoluble in water whereas LiCl soluble not only in water but also in acetone ?

Question 10.24 Explain the significance of sodium, potassium, magnesium and calcium in biological fluids.

approx 1:10 What happens when

- (i) sodium metal is dropped in water ?
- (ii) sodium metal is heated in free supply of air ?
- (iii) sodium peroxide dissolves in water ?

Question 10.26 Comment on each of the following observations:

- (a) The mobilities of the alkali metal ions in aqueous solution are $\text{Li}^+ < \text{Na}^+ < \text{K}^+ < \text{Rb}^+ < \text{Cs}^+$
- (b) Lithium is the only alkali metal to form a nitride directly.
- (c) E^0 for $\text{M}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{M}(\text{s})$ (where $\text{M} = \text{Ca}, \text{Sr}$ or Ba) is nearly constant.

Question 10.27 State as to why

- (a) a solution of Na_2CO_3 is alkaline ?
- (b) alkali metals are prepared by electrolysis of their fused chlorides ?
- (c) sodium is found to be more useful than potassium ?

Question 10.28 Write balanced equations for reactions between

- (a) Na_2O_2 and water
- (b) KO_2 and water
- (c) Na_2O and CO_2 .

Question 10.29 How would you explain the following observations?

- (i) BeO is almost insoluble but BeSO₄ is soluble in water,
- (ii) BaO is soluble but BaSO₄ is insoluble in water,
- (iii) LiI is more soluble than KI in ethanol.

Question 10.30 Which of the alkali metal is having least melting point ?

- (a) Na
- (b) K
- (c) Rb
- (d) Cs

Question 10.31 Which one of the following alkali metals gives hydrated salts ?

- (a) Li
- (b) Na
- (c) K
- (d) Cs

Question 10.32 Which one of the alkaline earth metal carbonates is thermally the most stable ?

- (a) MgCO₃
- (b) CaCO₃
- (c) SrCO₃
- (d) BaCO₃

(Chemistry) Chapter 11 The P -Block Elements

NCERT Exercises Questions

Question 11.1 Discuss the pattern of variation in the oxidation states of (i) B to Tl and (ii) C to Pb.

Question 11.2 How can you explain higher stability of BCl₃ as compared to TlCl₃ ?

Question 11.3 Why does boron trifluoride behave as a Lewis acid ?

Question 11.4 Consider the compounds, BCl₃ and CCl₄. How will they behave with water ? Justify.

Question 11.5 Is boric acid a protic acid ? Explain.

Question 11.6 Explain what happens when boric acid is heated .

Question 11.7 Describe the shapes of BF_3 and BH_4^- . Assign the hybridisation of boron in these species.

Question 11.8 Write reactions to justify amphoteric nature of aluminium.

Question 11.9 What are electron deficient compounds ? Are BCl_3 and SiCl_4 electron deficient species ? Explain.

Question 11.10 Write the resonance structures of CO_3^{2-} and HCO_3^- .

Question 11.11 What is the state of hybridisation of carbon in (a) CO_3^{2-} (b) diamond (c) graphite?

Question 11.12 Explain the difference in properties of diamond and graphite on the basis of their structures.

Question 11.13 Rationalise the given statements and give chemical reactions : • Lead(II) chloride reacts with Cl_2 to give PbCl_4 . • Lead(IV) chloride is highly unstable towards heat. • Lead is known not to form an iodide, PbI_4 .

Question 11.14 Suggest reasons why the B–F bond lengths in BF_3 (130 pm) and BF_4^- (143 pm) differ.

Question 11.15 If B–Cl bond has a dipole moment, explain why BCl_3 molecule has zero dipole moment.

Question 11.16 Aluminium trifluoride is insoluble in anhydrous HF but dissolves on addition of NaF. Aluminium trifluoride precipitates out of the resulting solution when gaseous BF_3 is bubbled through. Give reasons.

Question 11.17 Suggest a reason as to why CO is poisonous.

Question 11.18 How is excessive content of CO_2 responsible for global warming ?

Question 11.19 Explain structures of diborane and boric acid.

Question 11.20 What happens when (a) Borax is heated strongly,

- (b) Boric acid is added to water,
- (c) Aluminium is treated with dilute NaOH,
- (d) BF_3 is reacted with ammonia ?

Question 11.21 Explain the following reactions

- (a) Silicon is heated with methyl chloride at high temperature in the presence of copper;
- (b) Silicon dioxide is treated with hydrogen fluoride;
- (c) CO is heated with ZnO;
- (d) Hydrated alumina is treated with aqueous NaOH solution.

Question 11.22 Give reasons :

- (i) Conc. HNO_3 can be transported in aluminium container.
- (ii) A mixture of dilute NaOH and aluminium pieces is used to open drain.
- (iii) Graphite is used as lubricant.
- (iv) Diamond is used as an abrasive.
- (v) Aluminium alloys are used to make aircraft body.
- (vi) Aluminium utensils should not be kept in water overnight.
- (vii) Aluminium wire is used to make transmission cables.

Question 11.23 Explain why is there a phenomenal decrease in ionization enthalpy from carbon to silicon ?

Question 11.24 How would you explain the lower atomic radius of Ga as compared to Al ?

Question 11.25 What are allotropes? Sketch the structure of two allotropes of carbon namely diamond and graphite. What is the impact of structure on physical properties of two allotropes?

Question 11.26 (a) Classify following oxides as neutral, acidic, basic or amphoteric: CO, B_2O_3 , SiO_2 , CO_2 , Al_2O_3 , PbO_2 , Ti_2O_3 (b) Write suitable chemical equations to show their nature.

Question 11.27 In some of the reactions thallium resembles aluminium, whereas in others it resembles with group I metals. Support this statement by giving some evidences.

Question 11.28 When metal X is treated with sodium hydroxide, a white precipitate (A) is obtained, which is soluble in excess of NaOH to give soluble complex (B). Compound (A) is soluble in dilute HCl to form compound (C). The compound (A) when heated strongly gives (D), which is used to extract metal. Identify (X), (A), (B), (C) and (D). Write suitable equations to support their identities.

Question 11.29 What do you understand by (a) inert pair effect (b) allotropy and (c) catenation?

Question 11.30 A certain salt X, gives the following results.

- (i) Its aqueous solution is alkaline to litmus.
- (ii) It swells up to a glassy material Y on strong heating.
- (iii) When conc. H_2SO_4 is added to a hot solution of X, white crystal of an acid Z separates out. Write equations for all the above reactions and identify X, Y and Z.

Question 11.31 Write balanced equations for:

- (i) $\text{BF}_3 + \text{LiH} \rightarrow$
- (ii) $\text{B}_2\text{H}_6 + \text{H}_2\text{O} \rightarrow$
- (iii) $\text{NaH} + \text{B}_2\text{H}_6 \rightarrow$
- (iv) $\text{H}_3\text{BO}_3 \xrightarrow{\Delta}$
- (v) $\text{Al} + \text{NaOH} \rightarrow$
- (vi) $\text{B}_2\text{H}_6 + \text{NH}_3 \rightarrow$

Question 11.32. Give one method for industrial preparation and one for laboratory preparation of CO and CO_2 each.

Question 11.33 An aqueous solution of borax is

- (a) neutral
- (b) amphoteric
- (c) basic
- (d) acidic

Question 11.34 Boric acid is polymeric due to

- (a) its acidic nature
- (b) the presence of hydrogen bonds
- (c) its monobasic nature
- (d) its geometry

Question 11.35 The type of hybridisation of boron in diborane is

- (a) sp
- (b) sp²
- (c) sp³
- (d) dsp²

Question 11.36 Thermodynamically the most stable form of carbon is

- (a) diamond
- (b) graphite
- (c) fullerenes
- (d) coal

Question 11.37 Elements of group 14

- (a) exhibit oxidation state of +4 only
- (b) exhibit oxidation state of +2 and +4
- (c) form M²⁻ and M⁴⁺ ion
- (d) form M²⁺ and M⁴⁺ ions

Question 11.38 If the starting material for the manufacture of silicones is RSiCl₃, write the structure of the product formed

(Chemistry) Chapter 12 Organic Chemistry – Some Basic Principles And Techniques

NCERT Exercises Questions

Question 12.1 What are hybridisation states of each carbon atom in the following compounds ?
CH₂=C=O, CH₃CH=CH₂, (CH₃)₂CO, CH₂=CHCN, C₆H₆

Question 12.2 Indicate the σ and π bonds in the following molecules : C₆H₆, C₆H₁₂, CH₂Cl₂,
CH₂=C=CH₂, CH₃NO₂, HCONHCH₃

Question 12.3 Write bond line formulas for : Isopropyl alcohol, 2,3-Dimethyl butanal, Heptan-4- one.

Question 12.4 Give the IUPAC names of the following compounds :

Question 12.5 Which of the following represents the correct IUPAC name for the compounds concerned ?

- (a) 2,2-Dimethylpentane or 2-Dimethylpentane
- (b) 2,4,7- Trimethyloctane or 2,5,7-Trimethyloctane
- (c) 2-Chloro-4-methylpentane or 4-Chloro-2-methylpentane
- (d) But-3-yn-1-ol or But-4-ol-1-yne.

Question 12.6 Draw formulas for the first five members of each homologous series beginning with the following compounds

- (a) H-COOH
- (b) CH₃COCH₃
- (c) H-CH=CH₂

Question 12.7 Give condensed and bond line structural formulas and identify the functional group(s) present, if any, for :

- (a) 2,2,4-Trimethylpentane
- (b) 2-Hydroxy-1,2,3-propanetricarboxylic acid
- (c) Hexanedial

Question 12.8 Identify the functional groups in the following compounds

Question 12.9 Which of the two: O₂NCH₂CH₂O⁻ or CH₃CH₂O⁻ is expected to be more stable and why ?

Question 12.10 Explain why alkyl groups act as electron donors when attached to a π system.

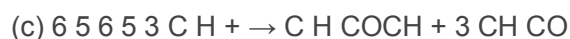
Question 12.11 Draw the resonance structures for the following compounds. Show the electron shift using curved-arrow notation.

- (a) C₆H₅OH
- (b) C₆H₅NO₂
- (c) CH₃CH=CHCHO
- (d) C₆H₅-CHO
- (e) $\text{C}_6\text{H}_5\text{CH}^+\text{CH}_2^-$

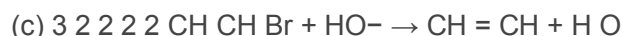
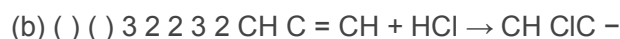


Question 12.12 What are electrophiles and nucleophiles ? Explain with examples.

Question 12.13 Identify the reagents shown in bold in the following equations as nucleophiles or electrophiles :



Question 12.14 Classify the following reactions in one of the reaction type studied in this unit.



Question 12.15 What is the relationship between the members of following pairs of structures ? Are they structural or geometrical isomers or resonance contributors ?

Question 12.16 For the following bond cleavages, use curved-arrows to show the electron flow and classify each as homolysis or heterolysis. Identify reactive intermediate produced as free radical, carbocation and carbanion.

Question 12.17 Explain the terms Inductive and Electromeric effects. Which electron displacement effect explains the following correct orders of acidity of the carboxylic acids?



Question 12.18 Give a brief description of the principles of the following techniques taking an example in each case.

(a) Crystallisation

(b) Distillation

(c) Chromatography

Question 12.19 Describe the method, which can be used to separate two compounds with different solubilities in a solvent S.

Question 12.20 What is the difference between distillation, distillation under reduced pressure and steam distillation ?

Question 12.21 Discuss the chemistry of Lassaigne's test 12.22 Differentiate between the principle of estimation of nitrogen in an organic compound by (i) Dumas method and (ii) Kjeldahl's method.

Question 12.23 Discuss the principle of estimation of halogens, sulphur and phosphorus present in an organic compound.

Question 12.24 Explain the principle of paper chromatography.

Question 12.25 Why is nitric acid added to sodium extract before adding silver nitrate for testing halogens?

Question 12.26 Explain the reason for the fusion of an organic compound with metallic sodium for testing nitrogen, sulphur and halogens.

Question 12.27 Name a suitable technique of separation of the components from a mixture of calcium sulphate and camphor.

Question 12.28 Explain, why an organic liquid vaporises at a temperature below its boiling point in its steam distillation ?

Question 12.29 Will CCl_4 give white precipitate of AgCl on heating it with silver nitrate? Give reason for your answer.

Question 12.30 Why is a solution of potassium hydroxide used to absorb carbon dioxide evolved during the estimation of carbon present in an organic compound?

Question 12.31 Why is it necessary to use acetic acid and not sulphuric acid for acidification of sodium extract for testing sulphur by lead acetate test?

Question 12.32 An organic compound contains 69% carbon and 4.8% hydrogen, the remainder being oxygen. Calculate the masses of carbon dioxide and water produced when 0.20 g of this substance is subjected to complete combustion.

Question 12.33 A sample of 0.50 g of an organic compound was treated according to Kjeldahl's method. The ammonia evolved was absorbed in 50 ml of 0.5 M H_2SO_4 . The residual acid required 60 mL of 0.5 M solution of NaOH for neutralisation. Find the percentage composition of nitrogen in the compound.

Question 12.34 0.3780 g of an organic chloro compound gave 0.5740 g of silver chloride in Carius estimation. Calculate the percentage of chlorine present in the compound.

Question 12.35 In the estimation of sulphur by Carius method, 0.468 g of an organic sulphur compound afforded 0.668 g of barium sulphate. Find out the percentage of sulphur in the given compound.

Question 12.36 In the organic compound $\text{CH}_2 = \text{CH} - \text{CH}_2 - \text{CH}_2 - \text{C} \equiv \text{CH}$, the pair of hybridised orbitals involved in the formation of: $\text{C}_2 - \text{C}_3$ bond is:

- (a) $sp - sp^2$
- (b) $sp - sp^3$
- (c) $sp^2 - sp^3$
- (d) $sp^3 - sp^3$

Question 12.37 In the Lassaigne's test for nitrogen in an organic compound, the Prussian blue colour is obtained due to the formation of:

- (a) $\text{Na}_4[\text{Fe}(\text{CN})_6]$
- (b) $\text{Fe}_4[\text{Fe}(\text{CN})_6]_3$
- (c) $\text{Fe}_2[\text{Fe}(\text{CN})_6]$
- (d) $\text{Fe}_3[\text{Fe}(\text{CN})_6]_4$

Question 12.38 Which of the following carbocation is most stable ?

- (a) $(\text{CH}_3)_3\text{C}^+ + \text{C}_2\text{H}_5$
- (b) $(\text{CH}_3)_3\text{C}^+ + \text{C}$
- (c) $\text{CH}_3\text{CH}_2^+ + \text{C}_2\text{H}_5$
- (d) $\text{CH}_3^+ + \text{C}_2\text{H}_5\text{CH}_2$

Question 12.39 The best and latest technique for isolation, purification and separation of organic compounds is:

- (a) Crystallisation

- (b) Distillation
- (c) Sublimation
- (d) Chromatography

Question 12.40 The reaction: $\text{CH}_3\text{CH}_2\text{I} + \text{KOH}(\text{aq}) \rightarrow \text{CH}_3\text{CH}_2\text{OH} + \text{KI}$ is classified as :

- (a) electrophilic substitution
- (b) nucleophilic substitution
- (c) elimination
- (d) addition

(Chemistry) Chapter 13 Hydrocarbons

NCERT Exercises Questions

Question 13.1 How do you account for the formation of ethane during chlorination of methane ?

Question 13.2 Write IUPAC names of the following compounds :

- (a) $\text{CH}_3\text{CH}=\text{C}(\text{CH}_3)_2$
- (b) $\text{CH}_2=\text{CH}-\text{C}\equiv\text{C}-\text{CH}_3$

Question 13.3 For the following compounds, write structural formulas and IUPAC names for all possible isomers having the number of double or triple bond as indicated :

- (a) C_4H_8 (one double bond)
- (b) C_5H_8 (one triple bond)

Question 13.4 Write IUPAC names of the products obtained by the ozonolysis of the following compounds :

- (i) Pent-2-ene
- (ii) 3,4-Dimethyl-hept-3-ene
- (iii) 2-Ethylbut-1-ene
- (iv) 1-Phenylbut-1-ene

Question 13.5 An alkene 'A' on ozonolysis gives a mixture of ethanal and pentan-3-one. Write structure and IUPAC name of 'A'.

Question 13.6 An alkene 'A' contains three C – C, eight C – H σ bonds and one C – C π bond. 'A' on ozonolysis gives two moles of an aldehyde of molar mass 44 u. Write IUPAC name of 'A'.

Question 13.7 Propanal and pentan-3-one are the ozonolysis products of an alkene? What is the structural formula of the alkene?

Question 13.8 Write chemical equations for combustion reaction of the following hydrocarbons:

- (i) Butane
- (ii) Pentene
- (iii) Hexyne
- (iv) Toluene

Question 13.9 Draw the cis and trans structures of hex-2-ene. Which isomer will have higher b.p. and why?

Question 13.10 Why is benzene extra ordinarily stable though it contains three double bonds?

Question 13.11 What are the necessary conditions for any system to be aromatic? 13.12 Explain why the following systems are not aromatic?

Question 13.13 How will you convert benzene into

- (i) p-nitrobromobenzene
- (ii) m- nitrochlorobenzene
- (iii) p - nitrotoluene
- (iv) acetophenone?

Question 13.14 In the alkane $\text{H}_3\text{C} - \text{CH}_2 - \text{C}(\text{CH}_3)_2 - \text{CH}_2 - \text{CH}(\text{CH}_3)_2$, identify $1^\circ, 2^\circ, 3^\circ$ carbon atoms and give the number of H atoms bonded to each one of these.

Question 13.15 What effect does branching of an alkane chain has on its boiling point?

Question 13.16 Addition of HBr to propene yields 2-bromopropane, while in the presence of benzoyl peroxide, the same reaction yields 1-bromopropane. Explain and give mechanism.

Question 13.17 Write down the products of ozonolysis of 1,2-dimethylbenzene (o-xylene). How

does the result support Kekulé structure for benzene?

Question 13.18 Arrange benzene, n-hexane and ethyne in decreasing order of acidic behaviour. Also give reason for this behaviour.

Question 13.19 Why does benzene undergo electrophilic substitution reactions easily and nucleophilic substitutions with difficulty?

Question 13.20 How would you convert the following compounds into benzene?

- (i) Ethyne
- (ii) Ethene
- (iii) Hexane

Question 13.21 Write structures of all the alkenes which on hydrogenation give 2-methylbutane.

Question 13.22 Arrange the following set of compounds in order of their decreasing relative reactivity with an electrophile, E⁺

- (a) Chlorobenzene, 2,4-dinitrochlorobenzene, p-nitrochlorobenzene
 - (b) Toluene, p-H₃C – C₆H₄ – NO₂, p-O₂N – C₆H₄ – NO₂.
- 13.23 Out of benzene, m-dinitrobenzene and toluene which will undergo nitration most easily and why?

Question 13.24 Suggest the name of a Lewis acid other than anhydrous aluminium chloride which can be used during ethylation of benzene.

Question 13.25 Why is Wurtz reaction not preferred for the preparation of alkanes containing odd number of carbon atoms? Illustrate your answer by taking one example.

(Chemistry) Chapter 14 Environmental Chemistry

NCERT Exercises Questions

Question 14. 1 Define environmental chemistry.

Question 14. 2 Explain tropospheric pollution in 100 words. 1

Question 14. 3 Carbon monoxide gas is more dangerous than carbon dioxide gas. Why?

Question 14. 4 List gases which are responsible for greenhouse effect.

Question 14. 5 Statues and monuments in India are affected by acid rain. How?

Question 14. 6 What is smog? How is classical smog different from photochemical smogs?

Question 14. 7 Write down the reactions involved during the formation of photochemical smog.

Question 14. 8 What are the harmful effects of photochemical smog and how can they be controlled?

Question 14. 9 What are the reactions involved for ozone layer depletion in the stratosphere?

Question 14. 10 What do you mean by ozone hole? What are its consequences?

Question 14. 11 What are the major causes of water pollution? Explain.

Question 14. 12 Have you ever observed any water pollution in your area? What measures would you suggest to control it?

Question 14. 13 What do you mean by Biochemical Oxygen Demand (BOD)?

Question 14. 14 Do you observe any soil pollution in your neighbourhood? What efforts will you make for controlling the soil pollution?

Question 14. 15 What are pesticides and herbicides? Explain giving examples.

Question 14. 16 What do you mean by green chemistry? How will it help decrease environmental pollution?

Question 14. 17 What would have happened if the greenhouse gases were totally missing in the earth's atmosphere? Discuss. 1

Question 14.18 A large number of fish are suddenly found floating dead on a lake. There is no

evidence of toxic dumping but you find an abundance of phytoplankton. Suggest a reason for the fish kill.

Question 14. 19 How can domestic waste be used as manure?

Question 14. 20 For your agricultural field or garden you have developed a compost producing pit. Discuss the process in the light of bad odour, flies and recycling of wastes for a good produce.