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NCERT Solutions for Class 11th : Chapter - 01

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Super Easy Chemistry By

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Some Basic Concepts of Chemistry

" The Learning Process Continues until You Achieve the Goal. "

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1.1. Calculate the molecular mass of the following : (i) H_2O (ii) CO_2 (iii) CH_4

Ans. (i) H_2O = (2×At. mass of H) + (1×At. mass of O) = [2(1.0084) + 1(16.00)] amu = 2.016 u + 16.00 amu = 18.016 amu

(ii) CO_2 = (1 × At. mass of C) + (2 × At. mass of O) = [1(12.011) + 2 (16.00)] amu = 12.011 amu + 32.00 u = 44.01 amu

(iii) CH_4 = (1 × At. mass of C) + (4 × At. mass of H) = [1(12.011) + 4 (1.008)] amu = 12.011 amu + 4.032 amu = 16.043 amu

1.2. Calculate the mass percent of different elements present in Sodium Sulphate (Na_2SO_4).

Ans. Molar mass of Na_2SO_4 = [(2 × 23.0) + (32.00) + 4 (16.00)] = 142 g

Mass percent of an element = (Mass of that element in compound/Molar mass of that compound) × 100

∴ Mass % of sodium (Na) : (46/142) × 100 = 32.39% & Mass % of sulphur(S) : (32/142) × 100 = 22.54%

Mass % of oxygen:(O) : (64/142) × 100 = 45.07%

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

Element	% of Element	At. Mass	Relative moles = %mass/At.wt.	Simplest molar Ratio	Simplest whole no. Ratio
Fe	69.9 %	55.85 u	69.9/55.85 = 1.25	1.25/1.25	1 or x 2 = 2
O	30.1 %	16.00 u	30.1/16 = 1.88	1.88/1.25	1.5 or x 2 = 3

∴ The empirical formula of the iron oxide is Fe_2O_3.

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 C is burnt in 16 g of dioxygen.(iii) 2 moles of C are burnt in 16 g of dioxygen.

Ans. The balanced reaction of combustion of carbon in dioxygens is:

$$\underset{1\text{mole}}{C(s)} + \underset{1\text{mole}(32g)}{O_2(g)} \rightarrow \underset{1\text{mole}(44g)}{CO_2\ (g)}$$

(i) In dioxygen, combustion is complete. Therefore 1 mole of carbon dioxide produced by burning 1 mole of carbon.

(ii) O_2 acts as a limiting reagent as only 16 g of O_2 is available. Hence, it will react with 0.5 mole of carbon to give 22 g of CO_2 .

(iii) Here, O_2 acts as a limiting reagent as only 16 g of dioxygen is available. It is a limiting reactant. Thus, 16 g of dioxygen can combine with only 0.5 mole of carbon to give 22 g of carbon dioxide.

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 82.0245g mol^{-1}.

Ans. 0.375 M aqueous solution of sodium acetate means that 1000 mL of solution containing 0.375 moles of sodium acetate.

∴ No. of moles of sodium acetate in 500 mL = (0.375/1000)×500 = 0.375/2 = 0.1875

∴ Mass of sodium acetate acquired = 0.1875×82.0245 g = 15.380g [∵ Molar mass of sodium acetate = 82.0245g mol^{-1}]

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%.

Ans. Mass % of 69% means that 100g of nitric acid solution contain 69 g of nitric acid by mass.
Molar mass of nitric acid (HNO_3) = 1+14+48 = 63g mol^{-1} So, Number of moles in 69 g of HNO_3 = 69/63 moles = 1.095 moles
Volume of 100g nitric acid solution = 100/1.41 mL = 70.92 mL = 0.07092 L
∴ Conc. of HNO_3 in moles per litre = 1.095/0.07092 = 15.44 M

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

Ans. 1 mole of $CuSO_4$ contains 1 mole of copper. & Molar mass of CuSO4 = (63.5) + (32.00) + 4(16.00) = 159.5 g
159.5 g of CuSO4 contains 63.5 g of copper.
∴ copper can be obtained from 100 g of copper sulphate = (63.5/159.5)×100 = 39.81g

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. Given that the molar mass of the oxide is 159.69 g mol^{-1}

Ans. From Question no. 1.3. ∴ The empirical formula of the iron oxide is Fe_2O_3.
Mass of Fe_2O_3 = (2×55.85) + (3×16.00) = 159.7 g mol^{-1} ∵ n = Molar mass/Empirical formula mass = 159.7/159.6 = 1
Thus, Molecular formula is same as Empirical Formula i.e. Fe_2O_3.

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

Ans. Fractional Abundance of ^{35}Cl = 0.7577 & M.M = 34.9689 Fractional Abundance of ^{37}Cl = 0.2423 & M.M = 36.9659
∴ Average Atomic mass = (0.7577×34.9689)amu + (0.2423×36.9659) = 26.4959 + 8.9568 = 35.4527

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.

Ans. (i) 1 mole of C_2H_6 contains 2 moles of Carbon atoms ∴ 3 moles of of C_2H_6 will contain 6 moles of Carbon atoms
(ii) 1 mole of C_2H_6 contains 6 moles of Hydrogen atoms ∴ 3 moles of of C_2H_6 will contain 18 moles of Hydrogen atoms
(iii) 1 mole of C_2H_6 contains Avogadro's no. 6.02×10^{23} molecules
∴ 3 moles of of C_2H_6 will contain ethane molecule = $3\times6.02\times10^{23}$= 18.06×10^{23} molecules.

1.11. What is the conc. of sugar ($C_{12}H_{22}O_{11}$) in mol L^{-1} if its 20 g are dissolved in enough water to make a final volume up to 2L?

Ans. Molar mass of sugar ($C_{12}H_{22}O_{11}$) = (12×12)+(1×22)+(11×16) = 342 g mol^{-1}
No. of moles in 20g of sugar = 20/342 = 0.0585 mole & Volume of Solution = 2L (given)
Molar concentration = Moles of solute/Volume of solution in L = 0.0585mol/2L = 0.0293 mol L^{-1} = 0.0293 M

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?

Ans. Molar mass of methanol (CH_3OH) = (1×12)+(4×1)+(1×16) = 32 g mol^{-1} = 0.032 kg mol^{-1}
Molarity of the solution = 0.793/0.032 = 24.78 mol L^{-1}
Applying, M_1V_1 (Given Solution) = M_2V_2 (Solution to be prepared)
$24.78\times V_1 = 0.25\times2.5$ L ∴ V_1= 0.02522 L = 25.22 mL

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

Ans. Pressure is the force (i.e. weigh) acting per unit area.
P = F/A = $1034g\times9.8ms^{-2}/cm^2$ = $1034g\times9.8ms^{-2}/cm^2 \times$ 1kg/1000g × 100cm/1m × 100cm/1m = 1.01332×10^5 N
Now, 1Pa = 1N m^{-2} ∴ 1.01332×10^5 $N\times m^{-2}$ = 1.01332×10^5 Pa

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

Ans. The SI unit of mass is kilogram (kg).
The kg is defined as the mass of platinum-iridium (Pt-Ir) cylinder that is stored in an air-tight jar at International Bureau of Weigh and Measures in France.

1.15. Match the following prefixes with their multiples:

Prefixes	Multiples
(i) micro	10^6
(ii) deca	10^9
(iii) mega	10^{-6}
(iv) giga	10^{-15}
(v) femto	10

Ans.

(i) micro	10^{-6}
(ii) deca	10
(iii) mega	10^6
(iv) giga	10^9
(v) femto	10^{-15}

1.16. What do you mean by significant figures ?

Ans. Significant figures are meaningful digits which are known with certainty including the last digit whose value is uncertain.
For example : In 11.2546 g, there are 6 significant figures but here 11.254 is certain and 6 is uncertain and the uncertainty would be ±1 in the last digit. Hence last uncertain digit is also included in Significant figures.

1.17. A sample of drinking water was found to be severely contaminated with chloroform, $CHCl_3$, 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.

Ans. (i) 15 ppm means 5 parts in million(10^6) parts. ∴ % by mass = $15/10^6 \times 100$ = 15×10^{-4} = 1.5×10^{-3} %
(ii) Molar mass of chloroform($CHCl_3$) = 12+1+(3×35.5) = 118.5 g mol^{-1}
100g of the sample contain chloroform = 1.5×10^{-3}g
∴ 1000 g(1 kg) of the sample will contain chloroform = 1.5×10^{-2} g = 1.5×10^{-2}/118.65 mole = 1.266×10^{-4} mole
∴ Molality = 1.266×10^{-4} m.

1.18. Express the following in the scientific notation:
(i) 0.0048 (ii) 234,000 (iii) 8008 (iv) 500.0 (v) 6.0012

Ans. (i) 0.0048 = 4.8×10^{-3} (ii) 234, 000 = 2.34×10^5 (iii) 8008 = 8.008×10^3 (iv) 500.0 = 5.000×10^2 (v) 6.0012 = 6.0012×10^0

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

Ans.(i) 2 (ii) 3 (iii) 4 (iv) 3 (v) 4 (vi) 5

1.20. Round up the following upto three significant figures:

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

Ans.(i) 34.2 (ii) 10.4 (iii) 0.046 (iv) 2810

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 = ... dm^3

Ans. (a) Fixing the mass of dinitrogen as 28 g, masses of dioxygen combined will be 32, 64, 32 and 80 g in the given four oxides. These masses of O_2 bears a simple whole number ratio as 2:4 : 2:5. Hence, the data given will obey the law of multiple proportions. The statement is as follows two elements always combine in a fixed mass of other bearing a simple ratio to another to form two or more chemical compounds.

(b) (i) 1 km = 1km×1000m/1km×100cm/1m/10mm/1cm = 10^6 mm & 1 km = 1km×1000m/1km×1pm/10^{-12}m = 10^{15} pm

(ii) 1 mg = 1mg×1g/1000mg×1kg/1000g = 10^{-6} kg & 1 mg = 1mg×1g/1000mg×1ng/10^{-9}g = 10^{-6} ng

(iii) 1 mL = 1mL×1L/1000mL = 10^{-3} L & 1 mL = 1cm^3 = 1cm^3×(1dm×1dm×1dm/10cm×10cm×10cm) = 10^3dm^3

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

Ans. Distance covered = Speed×Time = $3.0 \times 10^8 ms^{-1} \times 2.00$ ns = $3.0\times10^8ms^{-1}\times2.00$ ns$\times10^{-9}$s/1ns = 6.00×10^{-1}m = 0.600m

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

Ans. (i) According to the reaction, 1 atom of A reacts with 1 molecule of B.

∴ 200 molecules of B will react with 200 atoms of A, thereby leaving 100 atoms of A unreacted. Hence, B is the L.R.

(ii) According to the reaction, 1 mol of A reacts with 1 mol of B.

∴ 2 mol of A will react with only 2 mol of B leaving 1 mol of B. Hence, A is the limiting reagent.

(iii) 1 atom of A combines with 1 molecule of B.

∴ All 100 atoms of A will combine with all 100 molecules of B. Hence, the mixture is stoichiometric & there is no L.R

(iv) 1 mol of atom A combines with 1 mol of molecule B.

∴ 2.5 mol of B will combine with only 2.5 mol of A. and 2.5 mol of A will be left unreacted. Hence, B is the limiting reagent.

(v) 1 mol of atom A combines with 1 mol of molecule B.

∴ 2.5 mol of A will combine with only 2.5 mol of B and the remaining 2.5 mol of B will be left. Hence, A is the L.R.

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 2.00×10^3g dinitrogen reacts with 1.00×10^3g of dihydrogen.

(ii) Will any of the two reactants remain unreacted ? (iii) If yes, which one and what would be its mass ?

Ans. 1 mole of dinitrogen (28g) reacts with 3 mole of dihydrogen (6g) to give 2 mole of ammonia (34g).

∴ 2000 g of N_2 will react with H_2 = 6/28 ×200g = 428.6g. Thus, here N_2 is the limiting reagent while H_2 is in excess.

28g of N_2 produce 34g of NH_3. ∴ 2000g of N_2 will produce = 34/28×2000g = 2428.57 g of NH_3.

(ii) N_2 is the limiting reagent and H_2 is the excess reagent. Hence, H_2 will remain unreacted.

(iii) Mass of dihydrogen left unreacted = 1000g - 428.6g = 571.4 g

1.25. How are 0.50 mol $Na_2CO_3$and 0.50 M Na_2CO_3 different?

Ans. Molar mass of Na_2CO_3 = (2×23)+12.00+(3×16) = 106 g mol^{-1}

∴ 0.50 mol Na_2CO_3 means 0.50×106g = 53g

0.50 M Na_2CO_3 means 0.50 mol of Na_2CO_3 i.e. 53g of Na_2CO_3 are present in 1litre of the solution.

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

Ans. Dihydrogen gas reacts with dioxygen gas as, $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

Thus, two volumes of dihydrogen react with one volume of dihydrogen to produce two volumes of water vapour.

Hence, ten volumes of dihydrogen will react with five volumes of dioxygen to produce ten volumes of water vapour.

1.27. Convert the following into basic units: (i) 28.7 pm (ii) 15.15 pm (iii) 25365 mg

Ans. (i) 1 pm = 10^{-12} m ∴ 28.7 pm = 28.7×10^{-12} m = 2.87×10^{-11} m

(ii) 1 pm = 10^{-12} m ∴ 15.15 pm = 15.15×10^{-12} m = 1.515×10^{-11} m

(iii) 1 mg = 10^{-3} g 25365 mg =$2.5365\times10^4\times10^{-3}$ g = 2.5365×10 g =$2.5365\times10\times10^{-3}$ kg ∴ 25365 mg = 2.5365×10^{-2} kg

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 Cl_2 (g)

Ans. (i) 1 g Au = 1/197 mol = $1/197\times6.022\times10^{23}$ atoms (ii) 1 g Na = 1/23 mol = $1/23\times6.022\times10^{23}$ atoms

(iii) 1 g Li = 1/7 mol = $1/7\times6.022\times10^{23}$ atoms (iv) 1 g Cl_2 = 1/71 mol = $1/71\times6.022\times10^{23}$ atoms

Thus, 1 g of Li has the largest number of atoms.

1.29. Calculate the molarity of a solution of ethanol in water in which the mole fraction of ethanol is 0.040 (density of H_2O is 1).

Ans. Mole fraction of C_2H_5OH = No. of moles of C_2H_5OH/No. of moles of solution

n_{C2H5OH} = $n(C_2H_5OH)/(C_2H_5OH)+n(H_2O)$ = 0.040 (Given) -------------------1

We have to find the number of moles of ethanol in 1L of the solution but the solution is dilute. Therefore, water is approx. 1L.
No. of moles in 1L of water = $1000g/18g\ mol^{-1}$ = 55.55 moles
Substituting $n(H_2O)$ = 55.55 in equation 1 ∴ $n(C_2H_5OH)/(C_2H_5OH) + 55.55 = 0.040$
⇒ $0.96\ n(C_2H_5OH) = 55.55 \times 0.040$ ⇒ $n(C_2H_5OH) = 2.31$ mol Hence, molarity of the solution = 2.31M

1.30. What will be the mass of one ^{12}C atom in g ?

Ans. 1 mol of ^{12}C atoms = 6.022×10^{23} atoms = 12g ∴ Mass of 1 atom $^{12}C = 12/6.022\times10^{23}$ g = 1.9927×10^{-23} g

1.31. How many significant figures should be present in the **Ans.** of the following calculations?
(i) 0.02856×298.15×0.112/0.5785 (ii) 5 × 5.364 (iii) 0.0125 + 0.7864 + 0.0215

Ans. (i) Least precise term i.e. 0.112 is having 3 significant digits. ∴ There will be 3 significant figures in the calculation.
(ii) 5.364 is having 4 significant figures. ∴ There will be 4 significant figures in the calculation.
(iii) Least number of decimal places in each term is 4. ∴ There will be 4 significant figures in the calculation.

1.32. Use the data given in the following table to calculate the molar mass of naturally occuring 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%

Ans. Molar mass of Ar = $\sum p_iA_i$ = (0.00337×35.96755)+(0.00063×37.96272)+(0.99600×39.9624) = 39.948 g mol^{-1}

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.

Ans. (i) 1 mol of Ar = 6.022×10^{23}atoms ∴ 52 mol of Ar = $52\times6.022\times10^{23}$atoms = 3.131×10^{25}atoms
(ii) 1 atom of He = 4 u of He 4 u of He = 1 Atom of He ∴ 52 u of He = 1/4 × 52 = 13 atoms
(iii) 1 mol of He = 4 g = 6.022×10^{23}atoms ∴ 52 g of He = $(6.022\times10^{23}/4) \times 52$ atoms = 7.8286×10^{24}atoms

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 11.6g. Calculate (i) empirical formula (ii) molar mass of the gas and (iii) molecular formula.

Ans. Amount of C in 3.38 g of CO_2 = 12/44 × 3.38 g = 0.9218 g & Amount of H in 0.690 g H_2O = 2/18 × 0.690 g = 0.0767 g
The compound contains only C and H, therefore total mass of the compound = 0.9218 + 0.0767 = 0.9985 g
∴ % of C in the compound = (0.9218/0.9985)×100 = 92.32 & % of H in the compound = (0.0767/0.9985)×100 = 7.68
(i) Calculation of empirical formula,
Moles of C in the compound = 92.32/12 = 7.69 & Moles of H in the compound = 7.68/1 = 7.68
Simplest molar ratio = 7.69 : 7.68 = 1(approx) ∴ Empirical formula CH
(ii) 10.0 L of the gas at STP weigh = 11.6 g
∴ 22.4 L of the gas at STP = 11.6/10.0 × 22.4 = 25.984 = 26 (approx) ∴ Molar mass of gass = 26 g mol^{-1}
(iii) Mass of empirical formula CH = 12+1 = 13
∴ n = Molecular Mass/Empirical Formula = 26/13 = 2 ∴ Molecular Formula = C_2H_2

1.35. Calcium carbonate reacts with aqueous HCl to give $CaCl_2$ and CO_2 according to the reaction,
$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
What mass of $CaCO_3$ is required to react completely with 25 mL of 0.75 M HCl ?

Ans. 1000 mL of 0.75 M HCl have 0.75 mol of HCl = 0.75×36.5 g = 24.375 g
∴ Mass of HCl in 25mL of 0.75 M HCl = 24.375/1000 × 25 g = 0.6844 g
From the given chemical equation, $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$
2 mol of HCl i.e. 73 g HCl react completely with 1 mol of $CaCO_3$ i.e. 100g
∴ 0.6844 g HCl reacts completely with $CaCO_3$ = 100/73 × 0.6844 g = 0.938 g

1.36. Chlorine is prepared in the laboratory by treating manganese dioxide (MnO_2) with aqueous hydrochloric acid according to the reaction $4HCl(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$
How many grams of HCl react with 5.0 g of manganese dioxide?

Ans. 1 mol of MnO_2 = 55+32 g = 87 g
87 g of MnO_2 react with 4 moles of HCl i.e. 4×36.5 g = 146 g of HCl.
∴ 5.0 g of MnO_2 will react with HCl = 146/87×5.0 g = 8.40 g.

HOT [High Order Thinking] Problems

1. If the concentration of glucose ($C_6H_{12}O_6$) in blood is 0.9 g L^{-1}, what will be the molarity of glucose in blood?
2. If the density of a solution is 3.12 g mL^1, the mass of 1.5 mL solution in significant figures is________
3. Which of the following reactions is not correct according to the law of conservation of mass?
(a) $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ **(b)** $C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$
(c) $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$ (d) $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
4. What will be the mass of one atom of C-12 in grams?
5. The density of 3 molal solution of NaOH is 1.110 g mL^{-1}. Calculate the molarity of the solution.
6. 4 L of dinitrogen reacted with 22.7 L of dioxygen and 45.4 L of nitrous oxide was formed. The reaction is given below:
$2N_2(g) + O_2(g) \rightarrow 2N_2O(g)$ Which law is being obeyed in this experiment? Write the statement of the law.
7. Which of the following solutions have the same concentration?
(a) 20 g of NaOH in 200 mL of solution (b) 0.5 mol of KC1 in 200 mL of solution
(c) 40 g of NaOH in 100 mL of solution (d) 20 g of KOH in 200 mL of solution
8. A gaseous hydrocarbon gives upon combustion 0.72 g of water and 3.08 g of CO_2. The empirical formula of the hydrocarbon is: (a) C_3H_4 (b) C_6H_5 **(c)** C_7H_8 (d) C_2H_4
9. 29.2 % (w/w) HCl stock solution has density of 1.25 g mL^1. The molecular weight of HCl is 36.5 g mol^{-1}. The volume (mL) of stock solution required to prepare a 200 mL solution of 0.4 M HCl is-
10. Dissolving 120 g of urea (mol. wt. 60) in 1000 g of water gave a solution of density 1.15 g/mL. The molarity of the solution is -