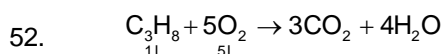


WEEKLY TEST MEDICAL PLUS -04 TEST - 01 RAJPUR
SOLUTION Date 21-07-2019

[CHEMISTRY]

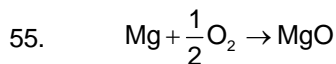
46.
$$\text{Molarity} = \frac{w}{M_B} \times \frac{1000}{V(\text{in mL})}$$
- $$w[\text{Ca(OH)}_2] = \frac{0.5 \times 74 \times 500}{1000} = 18.5\text{g}$$
- $$\text{Ca(OH)}_2 + \text{CO}_2 \rightarrow \text{CaCO}_3 + \text{H}_2\text{O}$$
- $$74\text{g Ca(OH)}_2 = 100\text{g CaCO}_3$$
- $$18.5\text{g Ca(OH)}_2 = \frac{100 \times 8.5}{74} = 25\text{g CaCO}_3$$
47. Molar mass of $\text{C}_{60}\text{H}_{122} = 842\text{g}$
- Mass of one molecule = $\frac{842}{6.02 \times 10^{23}} = 842 \times 1.66 \times 10^{-24} = 1.4 \times 10^{-21}\text{g}$
48.
$$15\text{ L H}_2(\text{g}) \text{ at STP} = \frac{15}{22.4} \times 6.02 \times 10^{23} = 4.03 \times 10^{23} \text{ molecules}$$
- $$15\text{ L N}_2(\text{g}) \text{ at STP} = \frac{15}{22.4} \times 6.02 \times 10^{23} = 1.34 \times 10^{23} \text{ molecules}$$
- $$0.5\text{ g H}_2(\text{g}) \text{ at STP} = \frac{0.5}{2} \times 6.02 \times 10^{23} = 1.5 \times 10^{23} \text{ molecules}$$
- $$10\text{ g O}_2(\text{g}) \text{ at STP} = \frac{10}{32} \times 6.02 \times 10^{23} = 1.88 \times 10^{23} \text{ molecules}$$
49. Average atomic weight = $\frac{(200 \times 90) + (199 \times 8) + (202 \times 2)}{100} = 199.96 = 200\text{amu}$
50.
$$\text{CH}_3\text{OH} + \frac{3}{2}\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}; \Delta H = -723\text{kJ}$$
- $$1.5\text{ mol O}_2 = 723\text{ kJ (evolved)}$$
- $$1\text{ mole O}_2 = \frac{723}{1.5} = 482\text{ kJ}$$
51.
$$100\text{amu} = (100) \left(\frac{1\text{g}}{6.022 \times 10^{23}} \right) = 1.66 \times 10^{-22}\text{g}$$
- Mass of 7.0×10^{22} molecules = $\frac{7.0 \times 10^{22}}{6.022 \times 10^{23}} \times 46 = 5.35\text{g}$
- Mass of 8.0×10^{-1} mol = $0.8 \times 46\text{g} = 36.8\text{g}$



53. Ratio of atoms C : H : : $\frac{85.6}{12} : \frac{14.4}{1} : : 7.13 : 14.4 : : 1 : 2$

Simplest formula : CH_2

54. Number of atoms = $3 \times$ Number of moles \times Avogadro Number
 $= 3 \times 0.1 \times 6.02 \times 10^{23} = 1.806 \times 10^{23}$



16 g oxygen = 24 g Mg

0.56 g oxygen = $\frac{24 \times 0.56}{16} = 0.84$ g Mg

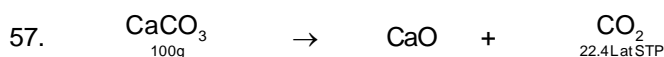
Given mass of Mg is 1.0 g which is surplus by $1.0 - 0.84 = 0.16$ g (Left)

56. Pressure exerted by H_2 = mole fraction of $H_2 \times$ total pressure
 Suppose w gram of both CH_4 and H_2 were taken.

Moles of $H_2 = \frac{w}{M.W} = \frac{w}{2}$; Moles of $CH_4 = \frac{w}{16}$

Mole fraction $H_2 = \frac{w/2}{\frac{w}{2} + \frac{w}{16}} = \frac{8}{9}$

Pressure exerted by $H_2 = \frac{8}{9} \times$ total pressure



22.4 L $CO_2 \equiv 100$ g $CaCO_3$

44.8 L $CO_2 = \frac{100 \times 44.8}{22.4} = 200$ g $CaCO_3$

For the use of 80 g $CaCO_3$, the amount taken = 100 g

For the use of 200 g $CaCO_3$, the amount taken = $\frac{100 \times 200}{80} = 250$ g

58. The average isotopic mass or atomic mass = $\sum m_i \times \frac{x_i}{100}$

where m_i = mass of i^{th} isotope, x_i = abundance of i^{th} isotope

\therefore Atomic mass = $54 \times \frac{5}{100} + 56 \times \frac{90}{100} + 57 \times \frac{5}{100}$

= 55.95

59. Mass of Fe in one mole of haemoglobin = 0.33% of 67200

= $\frac{0.33}{100} \times 67200 = 22.176$ g

No. of moles of Fe atoms per mole of haemoglobin = $\frac{22.176}{56}$

= 3.96 = 4 (whole number)

60. 490 mg $H_2SO_4 = 490 \times 10^{-3}$ g $H_2SO_4 = \frac{490 \times 10^{-3}}{98}$ mol

= $\frac{490 \times 10^{-3} \times 6.02 \times 10^{23}}{98}$ molecules = 3.01×10^{21} molecules

Molecules left over = $(3.01 \times 10^{21}) - (10^{20}) = 3.01 \times 10^{21} - 0.1 \times 10^{21}$
 $= (3.01 - 0.1) \times 10^{21} = 2.91 \times 10^{21}$

61. $\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$
 22400 mL of methane requires = 20 mL of oxygen.
 This means that 20 mL of methane will burn completely using 20 mL of oxygen.
 \therefore Volume of the gas left will be of oxygen only = $(50 - 20) = 30$ mL
62.
$$m = \frac{m}{d - M(M_B \text{ kg})} = \frac{0.5}{1.02 - 0.5 \times \frac{40}{1000}} = \frac{0.5}{1.02 - 0.02} = 0.5$$
63.
$$u_{\text{urea}} = \frac{15}{60} = \frac{1}{4} = 0.25$$

$$u_{\text{H}_2\text{O}} = \frac{175.5}{18} = 9.75$$

$$\chi_{\text{urea}} = \frac{0.25}{0.25 + 9.75} = \frac{0.25}{10} = 0.025$$
64. 11.11 moles of urea in 1000 g water, i.e., 55.55 moles of H_2O .

$$\chi_{\text{urea}} = \frac{11.11}{11.11 + 55.55} = \frac{1}{6} = 0.17$$
65.
$$M = \frac{10x\%d}{M_B}$$

$$\Rightarrow d = \frac{MM_B}{10x\%} = \frac{3.6 \times 98}{10 \times 29} = 1.216 \text{ g mL}^{-1}$$
66.
$$\text{Molarity} = \frac{10xd}{M_B} = \frac{10 \times 98 \times 1.96}{98} = 19.6 \text{ M}$$

 Normality of $\text{H}_2\text{SO}_4 = 2 \times \text{Molarity} = 2 \times 19.6 = 39.2 \text{ N}$
67. 1 L or 1000 mL of 0.001 M HCl solution contains 0.001 mole of Cl^- ions
 \therefore 100 mL of 0.001 M HCl solution will contain = $\frac{0.001}{10}$ mol of Cl^- ions
 1 mol of Cl^- ions $\equiv 6.023 \times 10^{23}$ Cl^- ions [\therefore Avogadro's law]
 \therefore 10^{-4} mol of $\text{Cl}^- \equiv 6.022 \times 10^{23} \times 10^{-4}$ Cl^- ions
 6.022×10^{19} Cl^- ions
68. Let the mass of $\text{O}_2 = x$ and that of $\text{N}_2 = 4x$
 No. of molecules of $\text{O}_2 = \frac{x}{32}$
 No. of molecules of $\text{N}_2 = \frac{4x}{28} = \frac{x}{7}$
 Ration $\frac{x}{32} : \frac{x}{7}$ or 7 : 32
69. The ratio of number of molecules is the same as the ratio of number of their moles,
 For the same weight x, ratio of number of molecules of O_2 and SO_2 will be
70. 300 mL of a gas weighs 0.368 g
 1 mL of a gas will weigh = $\frac{0.368}{300}$ g
 22400 mL of a gas will weight = $\frac{0.368}{300} \times 22400 = 27.477 \approx 27.5$ g
71. Gram molecular mass of NH_3 is 7 g.
 \therefore No. of molecules in 4.25 g of $\text{NH}_3 = \frac{4.25}{17} N_A = \frac{N_A}{4}$
 Now, one molecule of NH_3 contains 4 atoms

$$\therefore \frac{N_A}{4} \text{ molecules contain } \frac{N_A}{4} \times 4 = N_A \text{ atoms}$$

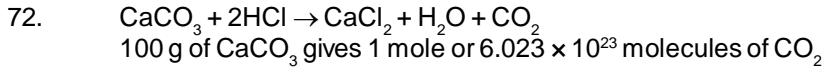
Again, 32 g of $O_2 = N_A$ molecules = $2N_A$ atoms

$$\therefore 8 \text{ g of } O_2 = \frac{N_A}{32} \times 8 = \frac{N_A}{4} \text{ molecules } \frac{2N_A}{32} \times 8 = \frac{N_A}{2} \text{ atoms}$$

On the other hand,

2g of $H_2 = N_A$ molecules = $2N_A$ atoms

4g of He = N_A atoms [\because gram atomic mass of He = 4g]



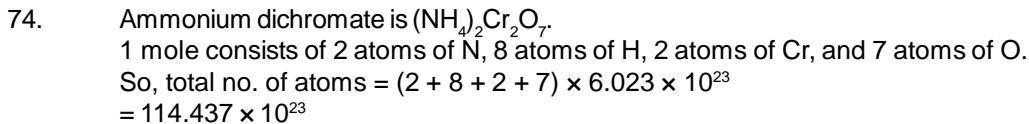
$$10^{-3} \text{ g of } CaCO_3 \text{ gives } = \frac{6.023 \times 10^{23}}{100} \times 10^{-3}$$

$$= 6.023 \times 10^{18} \text{ molecules of } CO_2$$

73. Number of atoms in 800 mg of Ca = $\frac{800 \times 10^{-3}}{40} \times N_A = 0.02N_A$ atoms

N_A atom of neon are present in 22.4 L

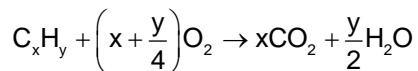
$$\therefore 0.02 N_A \text{ atoms of neon are present in } = \frac{22.4}{N_A} \times 0.02 \times N_A = 0.448L = 448cm^3$$



75. Moles of water produced = $\frac{0.72}{18} = 0.04$

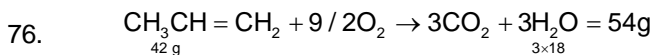
Moles of CO_2 produced = $\frac{3.08}{44} = 0.07$

Equation for combustion of an unknown hydrocarbon, C_xH_y is



$$\Rightarrow x = 0.07 \text{ and } \frac{y}{2} = 0.04 \Rightarrow y = 0.08 \text{ and } \frac{x}{y} = \frac{0.07}{0.08} = \frac{7}{8}$$

\therefore The empirical formula of the hydrocarbon is C_7H_8



54g pf $H_2P = 42$ g of propene

$$\therefore 24 \text{ g of } H_2O = \frac{42}{54} \times 27 = 21g$$



Mol. of mass of NaOH = 40 g mol⁻¹

$$\text{No. of moles in 0.064 g of NaOH} = \frac{0.064}{40} = 0.0016$$

$$\text{No. of mole of oxalic acid} = \frac{0.0016}{2} = 8 \times 10^{-4}$$

$$\text{Volume of solution (in L)} = \frac{25}{1000}$$

$$\text{Hence, molarity} = \frac{\text{No. of moles of solute}}{\text{Volume of solution (in L)}}$$

$$= 8 \times 10^{-4} \times \frac{1000}{25} = 0.032M$$

78. Normality = Molarity \times acidity of base $\text{Ca(OH)}_2 = N_1 = 0.1 \times 2 = 0.2$; $N_2 = 0.1$

$$N_1 V_1 = N_2 V_2$$

$\text{Ca(OH)}_2 \quad \text{HCl}$

$$0.2 \times V_1 = 0.1 \times 10 \Rightarrow V_1 = \frac{0.1 \times 10}{0.2} = 5 \text{ mL}$$

79. Number of gram equivalents of HCl = $\frac{\text{Normality} \times V}{1000} = \frac{0.1 \times 100}{1000} = 0.01$

Number of gram equivalents of metal carbonate = number of gram equivalents of HCl

$$\frac{W}{E} = 0.01 \Rightarrow \frac{2}{E} = 0.01 \Rightarrow E = 200$$

80. Mw_2 of $\text{CaCO}_3 = 40 + 12 + 48 = 100$

$$\text{Moles of } \text{CaCO}_3 \text{ in } 10\text{g} = \frac{10}{100} = 0.1 \text{ mol} = 0.1 \text{ g tom}$$

81. $N_1 V_1 + N_2 V_2 + N_3 V_3 = N_4 V_4$
($V_4 = V_1 + V_2 + V_3 + V_4$) or $V_4 = \text{Final volume} = 1 \text{ L} \equiv 1000 \text{ mL}$

$$5 \times N + 20 \times \frac{N}{2} + 30 \times \frac{N}{3} = N_4 = 100$$

$$\therefore N_4 = \frac{N}{40}$$

82. Weight of 6.023×10^{23} (Avogadro's number) = Mw of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 249 \text{ g}$
= 1 mol of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

$$\text{Weight of } 1 \times 10^{22} \text{ molecules of } \text{CuSO}_4 \cdot 5\text{H}_2\text{O} = \frac{249 \times 1 \times 10^{22}}{6.023 \times 10^{23}} = 4.14 \text{ g}$$

83. $M_1 = 1.0 \text{ M}$, $M_2 = 0.25 \text{ M}$

Let V_1 and V_2 are volumes required.

$$(1.0 \times V_1 + 0.25 \times V_2) = 0.75 (V_1 + V_2)$$

$$\Rightarrow 0.25V_1 = 0.5V_2, \Rightarrow V_1 : V_2 = 2 : 1$$

84. Molarity, normality, and formality are calculated against the volume of the solution. The volume of the solution changes with change in temperature; therefore, these quantities do not remain constant with temperature

$$\text{Molality (m)} = \frac{\text{Moles of solute}}{\text{Mas of solvent in kg}}$$

The molality of a solution remains independent of temperature because it involves only mass, which is independent of temperature.

$$M = 0.875$$

85. D

86. The normality of oxalic acid dihydrate is

$$\frac{6.3}{63} \times \frac{1}{250} \times 100 = 0.4$$

[Ew for $(\text{COOH})_2 \cdot 2\text{H}_2\text{O}$ is 63]

$$NV(\text{acid}) = N_2 V_2 (\text{vase})$$

$$\text{or } 0.4 \times 10 = 0.1 \times V_2$$

$$\text{or } V_2 = 40 \text{ mL}$$

87. B

88. C

89. D

90. B