**Chapter 1: Mole Concept-I**

**Review Questions**

**1.** Define mass and weight. How are they measured?

**Solution**

Mass is defined as the amount of matter present in the object. Mass is measured with the help of an instrument called a measuring balance.

The weight of an object is a measure of the effect of gravity on the object. Weight is determined by using an instrument called a weighing scale, which measures force against a spring. This means that mass is independent of the location of an object but weight is not. The SI unit of mass is kilogram (kg) and that of weight is newton (N).

**2.** What does symbol SI signify? Name the seven basic SI units.

**Solution**

SI signifies International System of Units or *Système Internationale* (in French). The seven basic SI units are meter, kilogram, second, ampere, kelvin, mole, and candela.

**3.** Explain why chemists might find it more useful to describe concentrated $H_2SO_4$ as a solution that contains 18.0 mol of solute per liter instead of as a solution that is 96.0% $H_2SO_4$ by mass.

**Solution**

Mole is a more convenient unit. Percent mass requires a conversion to moles.

**4.** What is the difference between the mass of a molecule and gram molecular mass?

**Solution**

Mass of a molecule is the actual mass of a single molecule expressed in grams, whereas gram molecular mass is the mass in gram of Avogadro's number of molecules.

**5.** Calculate the amount of carbon dioxide that could be produced when

(A) 1 mol of carbon is burnt in air.

(B) 1 mol of carbon is burnt in 16 g of dioxygen.

(C) 2 mol of carbon are burnt in 16 g of dioxygen.

**Solution**

The balanced reaction of combustion of carbon can be written as:

$$\begin{array}{ccc} C(s) + O_2(s) & \rightarrow & CO_2(s) \\ 1\,\text{mol} \;\; 1\,\text{mol} & & 1\,\text{mol} \\ (32\text{ g}) & & (44\text{ g}) \end{array}$$

(A) As per the balanced equation, 1 mol of carbon burns in 1 mol of dioxygen (air) to produce 1 mol of carbon dioxide.

(B) According to the question, only 16 g of dioxygen is available. Hence, it will react with 0.5 mol of carbon to give 22 g of carbon dioxide. Hence, dioxygen is the limiting reagent.

(C) Since only 16 g of dioxygen is available, which acts as a limiting reagent, so 16 g of dioxygen can combine with only 0.5 mol of carbon to give 22 g of carbon dioxide.

**6.** $n$ gram of a substance X reacts with $m$ g of substance Y to form $p$ g of substance R and $q$ g of substance S. This reaction can be represented as follows: $X + Y = R + S$. What is the relation that can be established in the amounts of the reactants and the products?

**Solution**

$nX + mY = pR + qS$

$n + m = p + q$

**7.** A gold-colored metal object has a mass of 365 g and a volume of 22.12 $cm^3$. Is the object composed of pure gold?

**Solution**

We know that $\text{Density} = \frac{\text{Mass}}{\text{Volume}}$

So, $\text{Density of the object} = \frac{365}{22.12} = 16.5\ \text{g cm}^{-3}$

The object is not composed of pure gold since the density of gold is 19.3 g cm$^{-3}$.

**8.** Why must measurements always be written with units? What is the meaning of each of the following prefixes?

(A) centi- (B) milli- (C) kilo- (D) micro-
(E) nano- (F) pico- (G) mega-

**Solution**

Measurements involve a comparison. The unit gives the number a meaning. The meaning of the each of the given prefixes is:

| | | |
|---|---|---|
| (A) | centi | $10^{-2}$ |
| (B) | milli | $10^{-3}$ |
| (C) | kilo | $10^{3}$ |
| (D) | micro | $10^{-6}$ |
| (E) | nano | $10^{-9}$ |
| (F) | pico | $10^{-12}$ |
| (G) | mega | $10^{6}$ |

**9.** State the rules to follow in rounding off a number.

**Solution**

While rounding off the measurement values, we use the following rules by convention:

(A) If the digit to be dropped is less than 5, then the preceding digit is left unchanged. For example, $x$ = 9.82 is rounded off to 9.8 and $x$ = 5.94 is rounded off to 5.9.

(B) If the digit to be dropped is more than 5, then the preceding digit is raised by one. For example, $x$ = 6.87 is rounded off to 6.9 and $x$ = 12.78 is rounded off to 12.8.

(C) If the digit to be dropped is 5 followed by digits other than zero, then the preceding digit is raised by one. For example, $x$ = 16.351 is rounded off to 16.4 and $x$ = 6.758 is rounded off to 6.8.

(D) If the digit to be dropped is 5 or 5 followed by zeros, then the preceding digit if it is even, is left unchanged. For example, $x$ = 3.250 becomes 3.2 on rounding off and $x$ = 12.650 becomes 12.6 on rounding off.

(E) If the digit to be dropped is 5 or 5 followed by zeros, then the preceding digit if it is odd, is raised by one. For example, $x$ = 3.750 is rounded off to 3.8, again $x$ = 16.150 is rounded off to 16.2.

**10.** Suppose 2.0 mol of $H_2$(g) is mixed with 1.0 mol of $O_2$(g) and allowed to react as shown below.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

(A) How many atoms of H and O are initially present?

(B) How many atoms of H and O will there be in the product?

(C) How many molecules of $H_2$ and $O_2$ were initially present?

(D) How many molecules of $H_2O$ were formed?

**Solution**

(A) According to the reaction four atoms of $H_2$ reacts with two atoms of oxygen to give two molecules of water.

(B) In the product, there are four atoms of hydrogen and two atoms of oxygen.

(C) Two molecules of $H_2$ and one molecule of $O_2$ was present.

(D) Two molecules of $H_2O$ were formed.

**11.** The following data are obtained when dinitrogen and dioxygen react together to form different compounds:

| **Mass of dinitrogen** | **Mass of dioxygen** |
|---|---|

| | |
|---|---|
| (A) 14 g | 16 g |
| (B) 14 g | 32 g |
| (C) 28 g | 32 g |
| (D) 28 g | 80 g |

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

**Solution**

If we fix the mass of dinitrogen at 28 g, then the masses of dioxygen that will combine with the fixed mass of dinitrogen are 32 g, 64 g, 32 g, and 80 g in the given four oxides. The masses of dioxygen bear a simple ratio 2:4:2:5. Hence, the given experimental data obeys the law of multiple proportions. The law states that whenever two elements form more than one compound, the different masses of one element that combine with the same mass of the other element bear a simple ratio to one another.

**12.** Describe what you need to do in the laboratory to test (A) the law of conservation of mass, (B) the law of definite proportions, and (C) the law of multiple proportions.

**Solution**

(A) To test the law of conservation of mass, a reaction would have to be carried out in which the mass of the reactants and the mass of the products are weighed and shown to be the same.

(B) The law of definite proportions could be shown by demonstrating that no matter how a compound is made, the same proportions by mass are used. This could be done by decomposing a compound and showing that the masses of the elements are always in the same ratio.

(C) To test the law of multiple proportions, two different compounds made up of the same elements would have to be decomposed. The amount used would have to keep mass of one of the elements constant, and then the masses of other the element from the different samples would have to be in a ratio of small whole numbers.

**13.** What is meant by the statement "1 g of Mg is burnt in a closed vessel that contains 0.5 g of $O_2$"?

**Solution**

1 g Mg is burnt in 0.5 g of $O_2$

That is, 24 g of Mg is burnt in $0.5 \times 32$ g of oxygen

24 g of Mg is burnt in 16 g of oxygen

Therefore, equivalent of MgO is equal to that of oxygen = 0.5/8

$$\underset{2\times 24\text{ g}}{2\text{Mg}} + \underset{\substack{32\text{ g}\\0.5\times 32\text{ g}\\=16\text{ g}}}{\text{O}_2} \longrightarrow \underset{80\text{ g}}{2\text{MgO}}$$

Moles of oxygen = Moles of magnesium = Moles of MgO in the reaction

Therefore, $1\,\text{mol} = 6.023\times 10^{23}$ atoms or molecules.

**14.** State true or false and give reasons: "Average atomic mass of an element depends mainly on the heavier isotope."

**Solution**

False.

Average atomic mass of an element does not depend mainly on the heavier isotope. For example, chlorine contains two types of atoms having relative masses 35 u and 37 u. The relative abundance of these isotopes in nature is in the ratio 3:1. Thus, atomic mass of chlorine is the average of these different relative masses as described below:

$$\text{Atomic mass of chlorine} = \frac{(35\times 3 + 37\times 1)}{4} = 35.5\text{ u}$$

**15.** Give reasons for the following:

(A) Both 106 g of sodium carbonate and 12 g of carbon have same number of carbon atoms.

(B) The volume occupied by one molecule of water, if density of water is $1.0 \times 10^3$ kg m$^{-3}$, is $2.99 \times 10^{-23}$ mL.

**Solution**

(A) Both contain 1 g atom of carbon that contains $6.023 \times 10^{23}$ carbon atoms

(B) $6.023 \times 10^{23}$ molecules = 1 mol

So, 1 molecule = $1/(6.023 \times 10^{23})$ mol

Hence, 1 mol of $H_2O$ = 18 g

**16.** Identify the compound and mixture from the following: cinnabar and brass.

**Solution**

Cinnabar is a chemical compound, whereas brass is a mixture. This is because cinnabar always contains 6.25 times as much mercury as sulphur by weight. Brass can be made with widely different ratios of copper and zinc.

17. What is the number of significant digits in Avogadro's constant?

**Solution**

$6.023 \times 10^{23}$: It has four significant figures. The exponential term does not add to the number of significant figures.

**Numerical Problems**

**1.** The interatomic distance in the hydrogen molecule is given as 0.74 Å. What is this distance in nanometers, micrometers, and millimeters?

**Solution**

Interatomic distance in the hydrogen molecule, $D$ = 0.74 Å

Interatomic distance in the hydrogen molecule in meters will be

$$D = 0.74\ \text{Å} \times \frac{10^{-10}\ \text{m}}{1\ \text{Å}} = 0.74 \times 10^{-10}\ \text{m}$$

Interatomic distance in the hydrogen molecule in nanometers will be

$$D = 0.74 \times 10^{-10}\ \text{m} \times \frac{1\ \text{nm}}{10^{-9}\ \text{m}} = 0.74 \times 10^{-1}\ \text{nm} = 0.074\ \text{nm}$$

Interatomic distance in the hydrogen molecule in micrometers will be

$$D = 0.74 \times 10^{-10}\ \text{m} \times \frac{1\ \mu\text{m}}{10^{-6}\ \text{m}} = 0.74 \times 10^{-4}\ \mu\text{m}$$

Interatomic distance in the hydrogen molecule in millimeters will be

$$D = 0.74 \times 10^{-10}\ \text{m} \times \frac{1\ \text{mm}}{10^{-3}\ \text{m}} = 0.74 \times 10^{-7}\ \text{mm}$$

**2.** Calculate the number of moles of nitrogen dioxide $NO_2$ that could be prepared from 0.35 mol of nitrogen oxide and 0.25 mol of oxygen.

$$2NO(g) + O_2 \rightarrow 2NO_2(g)$$

**Solution**

The reaction is $2NO(g) + O_2 \rightarrow 2NO_2(g)$

$$0.35\ \text{mol NO} \times \frac{2\ \text{mol } NO_2}{2\ \text{mol NO}} = 0.35\ \text{mol } NO_2$$

$$0.25\ \text{mol } O_2 \times \frac{2\ \text{mol } NO_2}{1\ \text{mol } O_2} = 0.50\ \text{mol } NO_2$$

NO produces fewer moles of $NO_2$; therefore, it is the limiting reagent. If the amount of NO were increased, the yield of $NO_2$ would increase. If the amount of $O_2$ were increased, there would be no change in the yield of $NO_2$.

**3.** $\beta$-Carotene is the pro-vitamin from which nature builds vitamin A. It is widely distributed in the plant and animal kingdoms, always occurring in plants together with chlorophyll. Calculate the molecular formula for $\beta$-carotene if the compound is 89.49% C and 10.51% H by mass and its molecular weight is 536.89 g mol$^{-1}$.

**Solution**

Based on the given data, we have

| Element | Percentage Composition | Atomic Mass | Moles of the Element | Simplest Molar Ratio | Simplest Whole Number Ratio |
|---|---|---|---|---|---|
| C | 89.49 | 12.011 | 89.49/12.011 = 7.451 | 7.451/7.451 = 1 | 1 × 5 = 5 |
| H | 10.51 | 1.008 | 10.51/1.008 = 10.43 | 10.43/7.451 = 1.400 | 1.4 × 5 = 7 |

Hence, the empirical formula is $C_5H_7$, which has a mass of 67.110 g mol$^{-1}$. Furthermore, the molar mass is some multiple of the empirical mass. In this case

$$n = \frac{\text{Molar mass of gas}}{\text{Empirical formula mass of gas}} = \frac{536.89}{67.11} = 8$$

Therefore, the molecular formula for $\beta$-carotene is $C_{40}H_{56}$.

$$\frac{5}{2}\times 5\times 10^{-4} = 12.5\times 10^{-4}\,\text{mol of } C_2O_4^{2-}$$
$$= 12.5\times 10^{-4}\,\text{mol of } Na_2C_2O_4$$
$$= 12.5\times 10^{-4}\times(46+28+64)\,\text{g}$$
$$= 12.5\times 10^{-4}\times 134\,\text{g} = 0.1675\,\text{g}$$

**4.** Urea ($H_2NCONH_2$) is manufactured by passing $CO_2$(g) through ammonia solution followed by crystallization. $CO_2$ for the above reaction is prepared by combustion of hydrocarbon. If combustion of 236 kg of a saturated hydrocarbon ($C_nH_{2n+2}$) produces as much $CO_2$ as required for production of 999.6 kg urea, then what is the molecular formula of hydrocarbon?

**Solution**

(B)

$$\underset{\substack{44\,\text{g}\\ ?}}{CO_2} + \underset{34\,\text{g}}{2NH_3} \rightarrow \underset{\substack{60\,\text{g}\\ 999.6\times 10^3\,\text{g}}}{NH_2CONH_2} + H_2O$$

$$\frac{44}{60}\times 999.6\times 10^3 = 733\,\text{kg of } CO_2$$

$$\underset{\substack{(12n+2n+2)\\(14n+2)\,\text{g}}}{C_nH_{2n+2}} + \left(\frac{3n+2}{2}\right)O_2 \rightarrow \underset{(44n)\,\text{g}}{nCO_2} + (n+1)H_2O$$

$$\frac{44n\,\text{g}\,CO_2}{(14n+2)}\times 236\times 10^3\,\text{g} = 733\times 10^3\,\text{g}\,CO_2$$

$$10384n = 10262n + 1466$$
$$122n = 1466$$
$$n = 12$$

So, $xC_nH_{2n+2} \Rightarrow C_{12}H_{26}$

**5.** In an experiment, 10 cm$^3$ of an organic compound in the gaseous state were sparked with an excess of oxygen. 20 cm$^3$ of carbon dioxide and 5 cm$^3$ of nitrogen were obtained among the products. All gas volumes were measured at the same temperature and pressure. What are the possible molecular formulas that would fit these data?

**Solution**

The possible formulas are:

(A) $C_2H_7N$ (B) $C_2H_3N$

$$\underset{\text{Organic compound}}{10\,\text{cm}^3\text{of}} + \underset{O_2}{\text{Excess}} \xrightarrow{\Delta} \underset{CO_2}{20\,\text{cm}^3\text{ of}} + \underset{N_2}{5\,\text{cm}^3\text{ of}}$$

From Avogadro's hypothesis, equal volume of all gases contain equal number of moles (or molecules) at same temperature and pressure. Therefore

$$\underset{\text{organic compound}}{1\,\text{mol of}} + \underset{O_2}{\text{Excess}} \xrightarrow{\Delta} \underset{CO_2}{2\text{ mol of}} + \underset{N_2}{1/2\text{ mol of}}$$

For

(A) $C_2H_7N + \frac{7.5}{2}O_2 \xrightarrow{\Delta} 2CO_2 + \frac{1}{2}N_2 + \frac{7}{2}H_2O$

(B) $C_2H_3N + \frac{5.5}{2}O_2 \xrightarrow{\Delta} 2CO_2 + \frac{1}{2}N_2 + \frac{3}{2}H_2O$

(C) $C_2H_6N_2 + \frac{7}{2}O_2 \xrightarrow{\Delta} 2CO_2 + N_2 + 3H_2O$

(D) $CH_5N + \frac{4.5}{2}O_2 \xrightarrow{\Delta} CO_2 + \frac{1}{2}N_2 + \frac{5}{2}H_2O$

Only (A) and (B) produce 2 mol of $CO_2$ and 0.5 mol of $N_2$.

**6.** Sulphuric acid reacts with sodium hydroxide as follows:

$$H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$$

When 1 L of 0.1 M sulphuric acid solution is allowed to react with 1 L of 0.1 M sodium hydroxide solution, what is the amount of sodium sulphate formed and its molarity in the solution?

**Solution**

$$\underset{\substack{1\,\text{L},\,0.1\,\text{M}\\0.1\,\text{mol}}}{H_2SO_4} + \underset{\substack{1\,\text{L},\,0.1\,\text{M}\\0.1\,\text{mol}}}{2NaOH} \rightarrow Na_2SO_4 + 2H_2O$$

But only 0.05 mol of $H_2SO_4$ is involved in the reaction. Therefore

$$\text{Number of moles of } Na_2SO_4 = 0.05\text{ mol} = 0.05 \times 142\text{ g} = 7.10\text{ g}$$

$$\text{Molarity} = \frac{\text{Number of moles of } Na_2SO_4}{\text{Liters of solution}} = \frac{0.05\text{ mol}}{2\text{ L}} = 0.025\text{ M}$$

**7.** If haemoglobin (molecular weight ≅ 67200) contains 0.33% of iron by weight, then what is the number of iron atoms present in one molecule of haemoglobin?

**Solution**
100 g haemoglobin has 0.33 g of Fe. Therefore

$$67200 \text{ g haemoglobin} = \frac{0.33 \times 67200}{100} = 221 \text{ g Fe} = \frac{221}{56} \text{ g atom}$$

So, Fe = 4 g atom Fe
Therefore, 1 molecule of haemoglobin has 4 atoms of Fe.

**8.** When 1 L of $CO_2$ is heated with graphite, the volume of the gases collected is 1.5 L. Calculate the number of moles of CO produced at STP.
**Solution**
When 1 L of $CO_2$ is heated with graphite, the volume collected is 1.5 L.

$$\text{Number of moles of CO} = \frac{1}{22.4}$$

**9.** X and Y are two elements that form $X_2Y_3$ and $X_3Y_4$. If 0.20 mol of $X_2Y_3$ weighs 32.0 g and 0.4 mol of $X_3Y_4$ weighs 92.8 g, what are the atomic masses of X and Y?
**Solution**
0.20 mol $X_2Y_3$ weight 32.0 g, so 1 mol $X_2Y_3$ weight 160 g
0.40 mol $X_3Y_4$ weight 92.8 g, so 1 mol $X_3Y_4$ weight 232 g
Thus, we have two linear equations as

$$M_X \times 2 + M_y \times 3 = 160$$
$$M_X \times 3 + M_y \times 4 = 232$$

Solving, we get $M_X$ = 56 u and $M_Y$ =16 u.

**10.** What will be the composition of residual mixture if 30 g of Mg combines with 30 g of oxygen?
(A) 40 g MgO + 20 g $O_2$ (B) 45 g MgO + 15 g $O_2$
(C) 50 g MgO + 10 g $O_2$ (D) 60 g MgO only
**Solution**
The reaction involved is

$$\begin{array}{cccc} 2\text{Mg} & + \ O_2 \rightarrow & & 2\text{MgO} \\ 24 \times 2 & 32 & \rightarrow & 80\text{ g} \\ 48\text{ g} & 32\text{ g} & & 80\text{ g} \\ 30\text{ g} & 30\text{g} & & \end{array}$$

From the reaction, we can see that 32 g of $O_2$ reacts with 48 g of Mg to form 80 g of MgO
30 g of $O_2$ will react with $\frac{48}{32} \times 30 = 45$ g of Mg
But the given amount of Mg is 30 g, so Mg is the limiting reactant.
Now, 48 g of Mg forms 80 g of MgO
So, 30 g of Mg will form $\frac{80}{48} \times 30 = 50$ g of MgO
Also, 48 g of Mg reacts completely with 32 g of $O_2$
So, 30 g of Mg will react completely with $\frac{32}{48} \times 30 = 40$ g of $O_2$
Hence, the amountof $O_2$ left unreacted is 40 – 30 = 10 g
and thus, the residual mixture contains 50 g of MgO and 10 g of $O_2$

**11.** Manganese trifluoride can be prepared by the following reaction:

$$MnI_2(s) + F_2(g) \rightarrow MnF_3 + IF_5$$

What is the minimum number of grams of $F_2$ that must be used to react with 12.0 g of $MnI_2$ if overall yield of $MnF_3$ is no more than 75%?

**Solution**

$$\underset{\substack{2(55+2\times127)\text{ g} \\ =618\text{ g} \\ 12\text{ g}}}{2MnI_2} + \underset{\substack{(13\times38)\text{ g} \\ =494\text{ g} \\ ?}}{13F_2} \rightarrow 2MnF_3 + 4IF_5$$

618 g of $MnI_2$ reacts with 494 g of $F_2$ to form $MnF_3$.
12 g of $MnI_2$ reacts with $x$ g to form $MnF_3$.

$$x = \frac{494\times12}{618} = 9.59\text{ g of } F_2$$

Given that yield = 75%, so the minimum number of grams of $F_2$ is

$$\text{Actual requirement} = \frac{9.59}{75}\times100 = 12.78\text{ g of } F_2$$

**Additional Questions**

**Single Correct Choice Type**

**Single Correct Choice Type**

**1.** How much quantity of zinc will have to be reacted with excess of dilute HCl solution to produce sufficient hydrogen gas for completely reacting with the oxygen obtained by decomposing 5.104 g of potassium chlorate?

(A) 8.124 g (B) 81.24 g (C) 0.08 g (D) 812.4 g

**Solution**

(A) The reactions are



$$\underset{\substack{245\text{ g} \\ 5.104\text{ g}}}{2KClO_3} \rightarrow 2KCl + \underset{\substack{96\text{ g} \\ ?(x)\text{ g}}}{3O_2}$$

$$x = \frac{96\text{ g}}{245\text{ g}}\times5.104\text{ g} = 2\text{ g of } O_2$$

$$\underset{\substack{4\text{ g} \\ ?}}{2H_2} + \underset{\substack{32\text{ g} \\ 2\text{ g}}}{O_2} \rightarrow 2H_2O$$

$$x = \frac{4\text{ g}}{32\text{ g}}\times2\text{ g} = 0.25\text{ g of } H_2$$

$$\underset{\substack{65\text{ g} \\ ?}}{Zn} + 2HCl \rightarrow ZnCl_2 + \underset{\substack{2\text{ g} \\ 0.25\text{ g}}}{H_2}\uparrow$$

$$x = \frac{65\text{ g}}{2\text{ g}}\times0.25\text{ g} = 8.125\text{ g}$$

**2.** Octane is a component of gasoline. Incomplete combustion of octane produces some CO along with $CO_2$ and $H_2O$, which reduces the efficiency of the engine. In a certain test run, 1.0 gallon of octane is burned and total mass of CO, $CO_2$, and $H_2O$ produced was found to be 11.53 kg. The efficiency of the engine would be (density of octane is 2.65 kg gallon$^{-1}$)

(A) 80% (B) 95.5% (C) 40.2% (D) 51.2%

**Solution**

(B) Given that density of octane = 2.65 kg gallon$^{-1}$
1 gallon of octane weighs 2.65 kg (2650 g).

$$\text{Number of moles of octane } (C_8H_8) = \frac{2650}{114} = 23.246\text{ mol}$$

$$C_8H_8 + nO_2 \rightarrow xCO + (8-x)CO_2 + 9H_2O$$

Here, $n = \text{Number of moles of oxygen} = \frac{1}{2}(x+16-2x+19) = \frac{1}{2}(25-x)$

Total mass of the product is

$$23.246 \times 28x + (8-x) \times 44 \times 23.246 + 9 \times 23.246 \times 18 = 11530$$

$$650.9x + 8182.6 - 1022.8x + 3765.8 = 11530$$

$$-371.9x + 11948.4 = 11530$$

$$-371.9x = -418.4$$

$$x = 1.125$$

Hence, $n = \frac{1}{2}(25-1.125) = 11.9375$

In case of complete combustion, $C_8H_8 + \frac{25}{2}O_2 \rightarrow 8CO_2 + 9H_2O$

$$\text{Efficiency} = \frac{11.9375}{12.5} \times 100 = 95.5\%$$

**3.** The number of moles of $CaCl_2$ needed to react with excess of $AgNO_3$ to produce 4.31 g of AgCl.

(A) 0.030 (B) 0.015 (C) 0.045 (D) 0.060

**Solution**

(B) $CaCl_2 + 2AgNO_3 \rightarrow 2AgCl + Ca(NO_3)_2$

We have $n = \frac{4.31}{143.5}$ and so, moles of $CaCl_2 = \frac{4.31}{143.5} \times \frac{1}{2} = 0.015\,\text{mol}$

**4.** Chlorine is prepared in the laboratory by treating manganese dioxide ($MnO_2$) with aqueous hydrochloric acid according to the following reaction:

$$4HCl\,(aq) + MnO_2(s) \rightarrow 2H_2O(l) + MnCl_2(aq) + Cl_2(g)$$

The grams of HCl that react with 5.0 g of manganese dioxide will be (at mass of Mn = 55)

(A) 84 g (B) 0.84 g (C) 8.4 g (D) 4.2 g

**Solution**

(C) $4HCl + MnO_2 \rightarrow MnCl_2 + Cl_2 + 2H_2O$

We have $n = \frac{5}{87}$, so

$$\text{Moles of HCl reacted} = \frac{5}{87} \times 4 = 0.05747\,\text{mol}$$

Mass of HCl = 0.05747 × 36.5 = 8.4 g

**5.** Copper reacts with dilute nitric acid according to the following equation:

$$3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$$

If a copper penny weighing 3.045 g is dissolved in a small amount of nitric acid and the resultant solution is diluted to 50.0 mL with water, what is the molarity of the $Cu(NO_3)_2$?

(A) 0.958 M (B) 0.278 M (C) 0.145 M (D) 0.312 M

**Solution**

(A) $\text{Molarity} = \frac{\text{Molecular weight} \times \text{Volume required}}{\text{Weight required} \times 1000} = \frac{57.8 \times 50}{3.045 \times 1000} = 0.958\,\text{M}$

**6.** 8 g of sulphur is burnt to form $SO_2$ that is oxidized by $Cl_2$ water. The solution is treated with $BaCl_2$ solution. The amount of $BaSO_4$ precipitate is

(A) 1 mol (B) 0.5 mol (C) 0.24 mol (D) 0.25 mol

**Solution**

(D) $S + O_2 \rightarrow SO_2$

Moles of $SO_2$ produced from 8 g of sulphur = 16/64 = 0.25

$SO_2 \ + \ Cl_2 \text{ water} \rightarrow SO_4^{2-} + 2Cl^-$

Moles of $SO_4^{2-}$ produced = Moles of $SO_2$ = 0.25

Now $SO_4^{2-} + BaCl_2 \rightarrow BaSO_4 + 2Cl^-$

Moles of $BaSO_4$ produced = Moles of $SO_4^{2-} = 0.25$

**7.** The sterile saline solution used to rinse contact lenses can be made by dissolving 400 mg of NaCl in sterile water and diluting to 100 mL. What is the molarity of the solution?

(A) 0.0685 M (B) 0.0312 M (C) 0.0212 M (D) 0.0418 M

**Solution**

(A) Molecular mass of NaCl is 58.5.

$1 \text{ molal contains } 58.5 \text{ g in } 1000 \text{ L}$

$1 \text{ molal contains } 5.85 \text{ g in } 100 \text{ mL}$

Now, $1$ molal contains $5.85$ g

Given that $x$ molal contains $400 \times 10^{-3}$ g

Therefore, $x = \dfrac{400 \times 10^{-3}}{5.85} = 0.0685 \text{ M}$

**8.** Titanium metal is obtained from the mineral rutile, $TiO_2$. How many kilograms of rutile are needed to produce 100.0 kg of Ti?

(A) 166.8 kg (B) 132.4 kg (C) 11.4 kg (D) 14.3 kg

**Solution**

(A) $\underset{80\text{ g}}{TiO_2} \rightarrow \underset{\substack{48\text{ g} \\ 100.0\text{ kg}}}{Ti} + \underset{32\text{ g}}{O_2}$

80 g $TiO_2$ gives 48 g of Ti and $x$ g $TiO_2$ gives $100 \times 10^3$ g of Ti. So,

$$x = \frac{80 \times 100 \times 10^3}{48} = 166.7 \text{ g}$$

**9.** A sample of tap water contains 366 ppm of $HCO_3^-$ ions with $Ca^{2+}$ ion. Now it is removed by Clark's method by addition of $Ca(OH)_2$. Then what minimum mass of $Ca(OH)_2$ will be required to remove $HCO_3^-$ ions completely from 500 g of same tap water sample.

(A) 1 g (B) 0.4 g (C) 0.222 g (D) 0.111 g

**Solution**

(D) 366 ppm of $HCO_3^-$ ions means 366 mg of $HCO_3^-$ is present per liter.

Therefore, $\dfrac{366}{61} m$ mol of $HCO_3^-$ ions per liter solution = 6 $m$ mol of $HCO_3^-$ ions per liter solution

or 3 $m$ mol of $Ca(HCO_3)_2$ = 3 mol of $Ca(OH)_2$ is required as $Ca^{2+}$: $HCO_3^-$ = 1:2

$3 \times 74 = 3 \times 74$ mg of $Ca(OH)_2$ per liter reqd $= \dfrac{3 \times 74}{2}$ mg of $Ca(OH)_2$ per 500 mL solution is reqd

= 0.111 g

**10.** An inorganic substance has the following composition: N = 35%, H = 5%, O = 60%. On being heated, it yielded a gaseous compound containing N = 63.3% and O = 36.37%. Which of the following reaction can be suggested based on the given data?

(A) $2HNO_3 \rightarrow N_2O_5 + H_2O$ (B) $NH_2OH + HONO \rightarrow N_2O + 2H_2O$

(C) $NH_4NO_3 \rightarrow N_2O + 2H_2O$ (D) $2HNO_2 \rightarrow N_2O_3 + H_2O$

**Solution**

(C) From the question, we have

| Element | Symbol | Percentage | Relative no. of moles | Simplest ratio of moles | Simplest whole number ratio |
|---|---|---|---|---|---|
| Nitrogen | N | 35 | $\frac{35}{14}=2.5$ | $\frac{2.5}{2.5}=1\times 2$ | 2 |
| Oxygen | O | 60 | $\frac{60}{16}=3.75$ | $\frac{3.75}{2.5}=1.5\times 2$ | 3 |
| Hydrogen | H | 5 | $\frac{5}{1}=5$ | $\frac{5}{2.5}=2\times 2$ | 4 |

Therefore, empirical formula $N_2H_4O_3$.
$N_2H_4O_3$ is nothing but $NH_4NO_3$.

**11.** 55 g $Ba(MnO_4)_2$ sample containing inert impurity is completely reacting with 100 mL of 56 volume strength of $H_2O_2$, then what will be the percentage purity of $Ba(MnO_4)_2$ in the sample? (Ba—137, Mn—55)

(A) 40% (B) 25% (C) 10% (D) 68.18%

**Solution**

(D) Milliequivalent of $Ba(MnO_4)_2$ reacted = Milliequivalent of $H_2O_2$ reacted

$$\frac{56\times 100}{5.6} = 1000$$

Mass equivalent of $H_2O_2$ = 1 equivalent of $H_2O_2$
Therefore, Moles of $Ba(MnO_4)_2$ ($n$-factor = 10) = 0.1 mol
and Weight of $Ba(MnO_4)_2$ = 0.1 × 375 g = 37.5 g

So, $\%\text{ purity of } Ba(MnO_4)_2 = \frac{37.5}{55}\times 100 = 68.18\%$

**12.** The simplest formula of a compound containing 50% of element X (atomic mass 10) and 50% of element Y (atomic mass 20) is

(A) XY (B) $X_2Y$ (C) $XY_3$ (D) $X_2Y_3$

**Solution**

(B) From the given data, we have

| Element | % ($a$) | Atomic weight ($b$) | $a/b$ | Ratio |
|---|---|---|---|---|
| X | 50 | 10 | 5 | 2 |
| Y | 50 | 20 | 2.5 | 1 |

So, the simplest formula = $X_2Y$.

**13.** Hydrochloric solutions A and B have concentration of 0.5 N and 0.1 N, respectively. The volume of solution A and B required to make 2 liters of 0.2 N hydrochloric are:

(A) 0.5 L of A + 1.5 L of B (C) 1.5 L of A + 0.5 L of B
(B) 1.0 L of A + 1.0 L of B (D) 0.75 L of A + 1.25 L of B

**Solution**

(A) $N_AV_A + N_BV_B = N_fV_f$. So,

$$0.5\times V_A + 0.1\times V_B = 0.2\times 2$$

Working out with each option, we get option (A) to the correct choice.

**14.** The density of 1 M solution of NaCl is 1.0585 g $mL^{-1}$. The molality of the solution is

(A) 1.0585 (B) 1.00 (C) 0.10 (D) 0.0585

**Solution**

(B) $m = \dfrac{w}{M \times W(\text{kg})}$ and $\rho = \dfrac{m}{V} \Rightarrow m = 1$ molal

**15.** $H_3PO_4$ is a acid and one of its salt in $NaH_2PO_4$. What volume of 1 M NaOH solution should be added to 12 g of $NaH_2PO_4$ to convert it into $Na_3PO_4$?

(A) 100 mL (B) 200 mL (C) 80 mL (D) 300 mL

**Solution**

(B) The reaction involved is

$$NaH_2PO_4 + 2NaOH \rightarrow Na_3PO_4 + 2H_2O$$

From the reaction, 1 mol $NaH_2PO_4$ reacts with 2 mol NaOH

Number of moles of $NaH_2PO_4$ = $\dfrac{12}{120} = 0.1$ mol

Therefore, number moles of NaOH = 2 × 0.1 = 0.2 mol

Given that molarity of NaOH = 1 M, so

$$1 = \frac{0.2}{V} \times 1000 \Rightarrow V = 200 \text{ mL}$$

**16.** A sample of hard water contains 1 mg $CaCl_2$ and 1 mg $MgCl_2$ per liter. Calculate the hardness of water in terms of $CaCO_3$ present in per $10^6$ parts of water.

(A) 215 ppm (B) 2.15 ppm (C) 105 ppm (D) None of these

**Solution**

(B) $55.5 \text{ g } CaCl_2 \equiv 50 \text{ g } CaCO_3 \Rightarrow 1 \text{ mg } CaCl_2 \equiv \dfrac{50}{55.5} \text{ mg } CaCO_3 \equiv 0.9\, CaCO_3 \text{ mg}$

$47.5 \text{ g } MgCl_2 \equiv 50 \text{ g } CaCO_3 \Rightarrow 1 \text{ mg } MgCl_2 = \dfrac{50}{47.5} \text{ mg } CaCO_3 \equiv 1.05\, CaCO_3$

So, Hardness in $CaCO_3$ (ppm) $\equiv \dfrac{(0.9 + 1.05) \times 10^{-3} \text{ gm}}{1/1000} \equiv 2.15$ ppm

**17.** A 10.0 g sample of a mixture of calcium and sodium chloride is treated with $Na_2CO_3$ to precipitate the calcium as calcium carbonate. This $CaCO_3$ is heated to convert all the calcium to CaO and final mass of CaO is 1.62 g. The % by mass of $CaCl_2$ in the original mixture is

(A) 15.2% (B) 32.1% (C) 21.8% (D) 11.7%

**Solution**

(B) $CaCl_2 + NaCl \xrightarrow{Ca_3^{2-}} CaCO_3 \downarrow \xrightarrow{\Delta} CaO + CO_2$

$$\text{Moles of CaO} = \frac{1.62}{56} \equiv \text{Moles of } CaCO_3 \equiv \text{Moles of } CaCl_2 = \frac{\text{Grams of } CaCl_2}{111}$$

So, Grams of $CaCl_2 = 3.21 \text{ g} \Rightarrow \% \; CaCl_2 = \dfrac{3.21}{10} \times 100 = 32.1\%$

**18.** A sample of hard water contains 244 ppm of $HCO_3^-$ ions. What is the minimum mass of CaO required to remove ions completely from 1 kg of such water sample?

(A) 56 mg (B) 112 mg (C) 168 mg (D) 244 mg

**Solution**

(B) From the given data, we have

$224 \text{ ppm } HCO_3^- \equiv 244 \text{gm } HCO_3^- \text{ in 100 L} \equiv 244 \text{mg } HCO_3^- \text{ in1.0 L}$

$\equiv 4 \text{ mmoles of } HCO_3^- \text{ in 1.0 L}$

$\equiv 2 \text{ mmoles of } Ca(HCO_3)_2 \text{ in 1.0 L}$

$$CaO + Ca(HCO_3)_2 \longrightarrow 2CaCO_3 + H_2O \equiv 2 \text{ mmoles CaO in 1.0 L}$$
$$\equiv 2\times 56 \equiv 112 \text{ mg CaO}$$

**19.** A metal oxide has the formula $Z_2O_3$. It can be reduced by hydrogen to give free metal and water. 0.16 g of the metal oxide required 6 mg of hydrogen for complete reduction. The atomic weight of the metal is

(A) 27.9 (B) 159.6 (C) 79.8 (D) 55.8

**Solution**

(D) $Z_2O_3 + 3H_2O \longrightarrow 2Z + 3H_2O$

$$\text{mmol } H_2 = \frac{6}{2} = 3$$

Now, $3 \text{ mmol } H_2 \equiv 1 \text{ mmol } Z_2O_3 = \frac{0.16}{(2Z+48)}\times 1000 \Rightarrow Z = 56 \text{ g}$

**20.** P and Q are two elements that form $P_2Q_3$ and $PQ_2$. If 0.15 mol of $P_2Q_3$ weights 15.9 g and 0.15 mol of $PQ_2$ weights 9.3 g, the atomic weight of P and Q is (respectively):

(A) 18, 26 (B) 26, 18 (C) 13, 9 (D) None of these

**Solution**

(B) $2P + 3Q \longrightarrow P_2Q_3$

$P + 2Q \longrightarrow PQ_2$

$$M_{P_2Q_3} = \frac{15.9}{0.15} = 2P + 3Q \text{ and } M_{PQ_2} = \frac{9.3}{1.15} = P + 2Q$$

Hence, atomic weight of P = 26 and Q = 18.

**Multiple Correct Choice Type**

**Multiple Correct Choice Type**

**1.** Which of the following solutions have the same concentration?

(A) 20 g of NaOH in 200 mL of solution. (B) 0.5 mol of KCl 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.

**Solution**

(A, B)

(A) 20 g of NaOH in 200 mL of solution

$$M = \frac{\text{Weight required}\times 1000}{\text{Molecular weight}\times\text{Volume required}} = \frac{20\times 1000}{40\times 200}\text{ M} = 2.5\text{ M}$$

(B) 0.5 mol of KCl in 200 mL of solution

$$M = \frac{\text{Number of moles}}{\text{Liters of solution}} = \frac{0.5}{0.2\text{ L}} = 2.5\text{ M}$$

(C) 40 g of NaOH in 100 mL of solution

$$M = \frac{\text{Weight required}\times 1000}{\text{Molecular weight}\times\text{Volume required}} = \frac{40\times 1000}{40\times 100} = 10\text{ M}$$

(D) 20 g KOH in 200 mL of solution

Molecular weight of KOH = 56 g

$$M = \frac{20\times 1000}{56\times 200} = 1.785\text{ M}$$

Hence, options (A) and (B) are the same concentration.

**2.** Which of the following quantities are dependent on temperature?

(A) Molarity (B) Normality (C) Molality (D) Mole fraction

**Solution**

(A, B) Molarity and normality are dependent on temperature, since molarity and normality involves the use of volume of solution. Molality and mole fraction do not depend on temperature.

**3.** $PF_3$ reacts with $XeF_4$ to give $PF_5$.

$$2PF_3(g) + XeF_4(s) \rightarrow 2PF_5(g) + Xe(g)$$

If 100.0 g of $PF_3$ and 50.0 of $XeF_4$ react, then which of the following statements is true?

(A) $XeF_4$ is the limiting reagent.
(B) $PF_3$ is the limiting reagent.
(C) 1.137 mol of $PF_5$ are produced.
(D) 0.482 mol of $PF_5$ are produced.

**Solution**

(A, D) $2\ PF_3(g) + XeF_4(s) \rightarrow 2\ PF_5(g) + Xe(g)$

$$100.0\text{ g PF}_3 \times \frac{1\text{ mol PF}_3}{87.968\text{ g}} \times \frac{2\text{ mol PF}_5}{2\text{ mol PF}_3} = 1.137\text{ mol PF}_5$$

$$50.0\text{ g XeF}_4 \times \frac{1\text{ mol XeF}_4}{207.28\text{ g}} \times \frac{2\text{ mol PF}_5}{1\text{ mol XeF}_4} = 0.482\text{ mol PF}_5$$

$XeF_4$ produces fewer moles of $PF_5$; therefore, it is the limiting reagent and 0.482 mol of $PF_5$ would be produced.

**4.** 1 mol of $H_2SO_4$ will exactly neutralize

(A) 2 mol of ammonia
(B) 1 mol of $Ca(OH)_2$
(C).0.5 mol of $Ba(OH)_2$
(D) 2 mol NaOH

**Solution**

(A, B, D)

Number of equivalent of $H_2SO_4$ = Moles × $n$-factor = 1 × 2 = 2

Number of equivalent of $Ca(OH)_2$ = 1 × 2 = 2 (neutralized)

Number of equivalent of NaOH = 2 × 1 = 2 (neutralized)

Number of equivalent of $NH_3$ = 2 × 1 = 2 (neutralized)

The required reactions are

$$H_2SO_4 + Ca(OH)_2 \rightarrow CaSO_4 + 2H_2O$$
$$2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$$
$$2NH_3 + H_2SO_4 \rightarrow 2NH_4^+ + SO_4^{2-}$$

**Assertion–Reasoning**

Choose the correct option from the following:

(A) Statements 1 and 2 are True and Statement 2 is the correct explanation of Statement 1.
(B) Statements 1 and 2 are True but Statement 2 is not the correct explanation of Statement 1.
(C) Statement 1 is True, Statement 2 is False.
(D) Statements 1 is False, Statement 2 is True.

**1. Statement 1:** Molecular mass = 2 × Vapor density.

**Statement 2:** Vapor density is the mass ratio of 1 mol of vapor to that of hydrogen.

**Solution**

(B) Under similar conditions,

$$\text{Vapor density} = \frac{\text{Mass of } V \text{ liter of gas}}{\text{Mass of } V \text{ liter of hydrogen}}$$

**2. Statement 1:** The molality of the solution does not change with temperature.

**Statement 2:** The molality is expressed in units of moles per 1000 g of solvent.

**Solution**

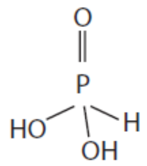
(B) Molality does not depend on volume; thus, it does not depend on temperature, as it does not involve volume term.

**3. Statement 1:** The percentage of nitrogen in urea is 46.6%.
**Statement 2:** Urea is a covalent compound.
**Solution**
(B) Urea is $H_2NCONH_2$

$$\% \text{ of N} = \frac{28}{60} \times 100 = 46.6\%$$

Urea is a covalent compound.

**4. Statement 1:** $H_3PO_3$ is a dibasic acid and its salt $Na_3PO_3$ does not exist.
**Statement 2:** Being dibasic in nature, only two hydrogen atoms are replaceable.
**Solution**
(A) The structure is




Here, the rightmost and leftmost H are replaceable; thus, only $Na_2HPO_3$ and $NaH_2PO_3$ salts are possible.



**Integer Answer Type**
The answer is a **non-negative integer.**

**1.** The number of moles and equivalents in 196 g of $H_3PO_4$ are ___.
**Solution**
(2,6)

$$n = \frac{196}{98} = 2$$

$$\text{Equivalents} = \text{Moles} \times n\text{-Factor} = 2 \times 3 = 6$$

**2.** The number of molecules in 16 g of $SO_2$ is ___.
**Solution**
$(1.5 \times 10^{23})$

$$n = \frac{16}{64} = 0.25$$

Therefore, $N = 0.25 \times 6.02 \times 10^{23} = 1.5 \times 10^{23}$

**3.** The number of gram atoms in 24 g of magnesium is ___.
**Solution**
(1) $n = \dfrac{w}{\text{Molecular weight}} = \dfrac{24\text{ g}}{24\text{ g/mol}} = 1\text{ mol}$

**4.** The residue obtained on strongly heating 2.48 g $Ag_2CO_3$ is ___. The reaction is

$$Ag_2CO_3 \xrightarrow{\Delta} Ag + CO_2 + O_2$$

Given that the atomic mass of Ag = 108 u.

**Solution**

(2) Given that 2.48 g residue is of Ag

$$Ag_2CO_3 \xrightarrow{\Delta} 2Ag\uparrow + CO_2\uparrow + O_2\uparrow$$

$$n_{Ag_2CO_3} = \frac{2.48}{276} = 0.009$$

Therefore, $n$Ag = 0.018

or $w_{Ag}$ = 0.018 × 108 = 1.944 ≈ 2 g

**5.** The atomicity of the molecule of $H_2SO_4$ is ___.

**Solution**

(7) Atomicity of an elementary substance is defined as the number of atoms in a molecule of the element. In $H_2SO_4$, there are 7 atoms: 2 × H + 1 × S + 4 × O = 7 atoms. Hence, atomicity is 7.

**6.** Two moles of 50% pure $Ca(HCO_3)_2$ on heating forms 1 mol of $CO_2$. The percentage yield of $CO_2$ is ____.

**Solution**

(100%) $\underset{1\ \text{mol}}{Ca(HCO_3)_2} \xrightarrow{1\ \text{mol}} CaCO_3 + H_2O + \underset{1\ \text{mol}}{CO_2}\uparrow$

Hence, the yield is 100%

**7.** 4.48 L of ammonia at STP is neutralized using 100 mL of a solution of $H_2SO_4$. The molarity of acid is ___.

**Solution**

(1) By equating equivalents, we get

$$\frac{4.48}{22.4} \times 1000 = 100 \times M \times 2 \Rightarrow M = 1\ \text{M}$$



**Matrix–Match Type**

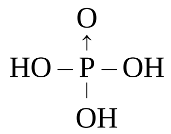
**1.** Match the acid with its nature based on its basicity.

| Column I | Column II |
|---|---|
| (A) $H_3PO_4$ | (p) Monobasic |
| (B) $H_3PO_3$ | (q) Pentabasic |
| (C) $H_3BO_3$ | (r) Tribasic |
| (D) EDTA | (s) Tetrabasic |
| | (t) Dibasic |

**Solution**

A → (r), B → (t), C → (p)

$H_3PO_4$ is tribasic acid ($n$ = 3); $H_3PO_3$ is dibasic acid ($n$ = 2); $H_3BO_3$ is monobasic acid ($n$ = 1); and EDTA is tetrabasic acid ($n$ = 4).



**2.** Match the number of moles with their amount.

| Column I | Column II |
|---|---|
| (A) 0.1 mol | (p) 4480 mL of $CO_2$ at STP |
| (B) 0.2 mol | (q) 0.1 g atom of iron |
| (C) 0 25 mol | (r) 1 5 × $10^{23}$ molecules of oxygen gas |
| (D) 0.5 mol | (s) 9 mL of water |
| | (t) 200 mg of hydrogen gas |

**Solution**

A → (q, t), B → (p), C → (r), D → (s)

(p) 4480 mL of $CO_2$ at STP

$$\frac{1\ \text{mol}}{22400\ \text{mL}} \times 4480\ \text{mL} = 0.2\ \text{mol}$$

(q) 0.1 g atom of iron = 0.1 mol

(r) $1.5 \times 10^{23}$ molecules of oxygen gas

$$\frac{1\ \text{mol}}{6 \times 10^{23}} \times 1.5 \times 10^{23} = 0.25\ \text{mol}$$

(s) 9 mL $H_2O$

$$\frac{1\ \text{mol}}{18\ \text{mL}\ H_2O} \times 9\ \text{mL} = 0.5\ \text{mol}$$

(t) 200 mg hydrogen gas

$$\frac{1\ \text{mol}}{2\ \text{g}} \times 0.2\ \text{g} = 0.1\ \text{mol}$$

**3.** Match the concentration terms with the factors affecting the concentrations.

| **Column I** | **Column II** |
|---|---|
| (A) Molarity (*M*) | (p) Temperature |
| (B) Molality (*m*) | (q) Pressure |
| (C) Mole fraction (*x*) | (r) Dilution |
| (D) Normality | (s) Volume |

**Solution**

A → (p, q, r, s), B → (q, r), C → (q, r), D → (r, s)

Molality, mole fraction, and normality are independent of temperature because all these involve weight, which does not depend on temperature.

On diluting a solution, all concentrations change.

Molarity, molality, and mole fraction are dependent on pressure.

Volume affects the concentration of molarity and normality.