Chapter notes

Class: XI

Chapter 1: Some Basic Concepts of Chemistry

Top concepts

1. The SI system has seven base units which pertain to the 7 fundamental scientific quantities

Base Physical Quantity	Symbol for quantity	Name of SI Unit	Symbol for SI Unit
Length	l	metre	m
Mass	m	kilogram	kg
Time	t	second	s
Electric current	I	ampere	A
Thermodynamic temperature	T	Kelvin	K
Amount of substance	n	mole	mol
Luminous intensity	I_v	candela	cd

2. The unit is written on the right of the quantity with a space between them.

3. The SI system allows the use of prefixes to indicate the multiples or submultiples of a unit.

Multiple	Prefix	Symbol
10^{-1}	deci	d

10	deca	da
10^2	hector	h
10^3	kilo	k
10^6	mega	M

4. To indicate very small numbers, we use negative exponents.
5. To indicate large numbers, we use positive exponents.
6. Scientific notation is the proper representation of a number in exponential form.
7. Precision indicates how closely repeated measurements match each other.
8. Accuracy indicates how closely a measurement matches the correct or expected value.
9. A result is valid only if it is both accurate and precise.
10. Significant figures are meaningful digits which are known with certainty.
11. There are certain rules for determining the number of significant figures:

    i) All non-zero digits are significant

    ii) Zeros preceding the first non-zero digit are not significant

    iii) Zeros between two non-zero digits are significant.

iv) Zeros at the end or right of the number are significant provided they are on the right side of the decimal point. But, if otherwise, the zeros are not significant.

14. During addition and subtraction, the result cannot have more digits to the right of the decimal point than either of the original numbers.

15. In multiplication and division with significant figures, the answer cannot have more significant figures than either of the original numbers.

16. There are 5 basic laws of chemical combinations that govern every reaction: Law of conservation of mass, law of definite proportions, law of multiple proportions, Gay Lussac's law of gaseous volumes, and lastly, Avogadro law.

17. **Law of Conservation of Mass:** Antoine Lavoisier established the Law of Conservation of Mass. It states that matter can neither be created nor destroyed. In other words, we can say that during any physical or chemical change, the total mass of reactants is equal to the total mass of products.

18. **Law of definite proportions:** Joseph Proust showed that a given compound always contains exactly the same proportion of elements by weight.

19. **Law of multiple proportions:** Dalton proposed the law of multiple proportions. According to this law if two elements can combine to form more than one compound, the mass of one element that combines with the fixed mass of the other element is in the ratio of small whole numbers

20. **Gay Lussac's Law of gaseous volumes:** When gases combine or are produced in a chemical reaction they do so in a simple ratio by volume, provided all the gases are at same temperature and pressure.

21. **Avogadro law:** At the same temperature and pressure, equal volumes of gases contain equal number of molecules.

22. **Dalton's atomic theory:** In 1808, Dalton published 'A New System of Chemical Philosophy' in which he proposed the following :
    - Matter consists of indivisible atoms.
    - All the atoms of a given element have identical properties including identical mass. Atoms of different elements differ in mass.
    - Compounds are formed when atoms of different elements combine in a fixed ratio.
    - Chemical reactions involve reorganisation of atoms. These are neither created nor destroyed in a chemical reaction.

23. Dalton's theory could explain the laws of chemical combination.
24. The number 6.022×10^{23} is called Avogadro's constant or Avogadro's number.
25. A mole is a collection of 6.022×10^{23} particles.
26. One mole is the amount of a substance that contains as many particles or entities as there are atoms in exactly 12 g (or 0.012 kg) of the ^{12}C
27. The mass of one mole of a substance in grams is called its molar mass.
28. The molar mass in grams is numerically equal to the atomic/molecular/formula mass in u.(u is the unified mass)

29. Molarity is the number of moles of solute in per liter of solution. Unit is moles per liter.

30. Molality is the number of solute present in 1kg of solvent.

31. Atomic Mass: Average relative mass of an atom of an element as compared with the mass of a carbon atom taken as 12 amu

32. Atomic mass expressed in grams is called gram atomic mass

33. Molecular Mass: Sum of the atomic masses of elements present in a molecule

34. Molecular mass expressed in grams is called gram molecular mass

35. Formula Mass: Sum of atomic masses of all atoms in a formula unit of the compound

36. Following relations given below can be summarized

- One mole of atoms $= 6.022 \times 10^{23}$ atoms=Gram atomic mass of element
- One mole of molecules= 6.022×10^{23} molecules= Gram molecular mass of substance

37. An empirical formula represents the simplest whole number ratio of various atoms present in a compound.

38. Molecular formula shows the exact number of different types of atoms present in a molecule of a compound.

39. If the mass per cent of various elements present in a compound is known, its empirical formula can be determined.

40. Molecular formula = n (Empirical formula), where n is a simple number and may have values 1, 2, 3….

41. Following steps should be followed to determine empirical formula of the compound

- Step 1: Conversion of mass per cent of various elements into grams.
- Step 2: Convert mass obtained in step1 into number of moles

- Step 3: Divide the mole value obtained in step 2 by the smallest mole value (out of the mole value of various elements calculated)
- Step 4: In case the ratios are not whole numbers, then they may be converted into whole number by multiplying by the suitable coefficient.
- Step 5: Write empirical formula by mentioning the numbers after writing the symbols of respective elements.

42. Anything that has mass and occupies space is called matter

43. Matter can either be a mixture or be a pure substance

44. Homogenous mixtures are those whose components completely mix with each other to make a uniform composition

45. Heterogeneous mixtures are not uniform, and their components are separable through physical methods

46. Pure substances can be elements or compounds

47. An element consists of only one type of particles

48. Two or more atoms of different elements combine to form a molecule of a compound

49. The constituents of a compound can be separated only by chemical methods.

50. A compound has properties different from its constituent elements

51. Isotopes are elements with same atomic number but different mass number.

52. Atomic mass is donated by *"u"* – unified mass.

53. One mole is the amount of a substance that contains as many particles as there are atoms in exactly12 g of the ^{12}C isotope

54. The mass of one mole of a substance in grams is called its molar mass

55. Out of various reactants in a reaction ,a reactant that is completely consumed in a chemical reaction is called limiting reagent

56. Stoichiometry gives a quantitative relation between reactant and product in a reaction. It also helps us in identifying limiting reagents

Top Formulae

1. Mass % of an element $= \dfrac{\text{Mass of that element in the compound}}{\text{Molar mass of compound}} \times 100$

2. $\text{Mass per cent} = \dfrac{\text{Mass of solute}}{\text{Mass of solution}} \times 100$

3. $\text{Mole Fraction} = \dfrac{\text{No. of mole of a particular component}}{\text{Total No. of moles of solution}}$

4. $\text{Molarity} = \dfrac{\text{No. of moles of solute}}{\text{Volume of solution in litres}}$

5. $\text{Molality} = \dfrac{\text{No. of moles of solute}}{\text{Mass of solvent in kg}}$

6. $\text{Moles of an element} = \dfrac{\text{Mass of element}}{\text{Atomic mass}}$

7. $\text{Mass of one atom} = \dfrac{\text{Atomic mass}}{6.022 \times 10^{23}}$

8. $\text{Moles of a compound} = \dfrac{\text{Mass of compound}}{\text{Molecular mass}}$

9. $\text{Mass of one molecule} = \dfrac{\text{Molecular mass}}{6.022 \times 10^{23}}$