

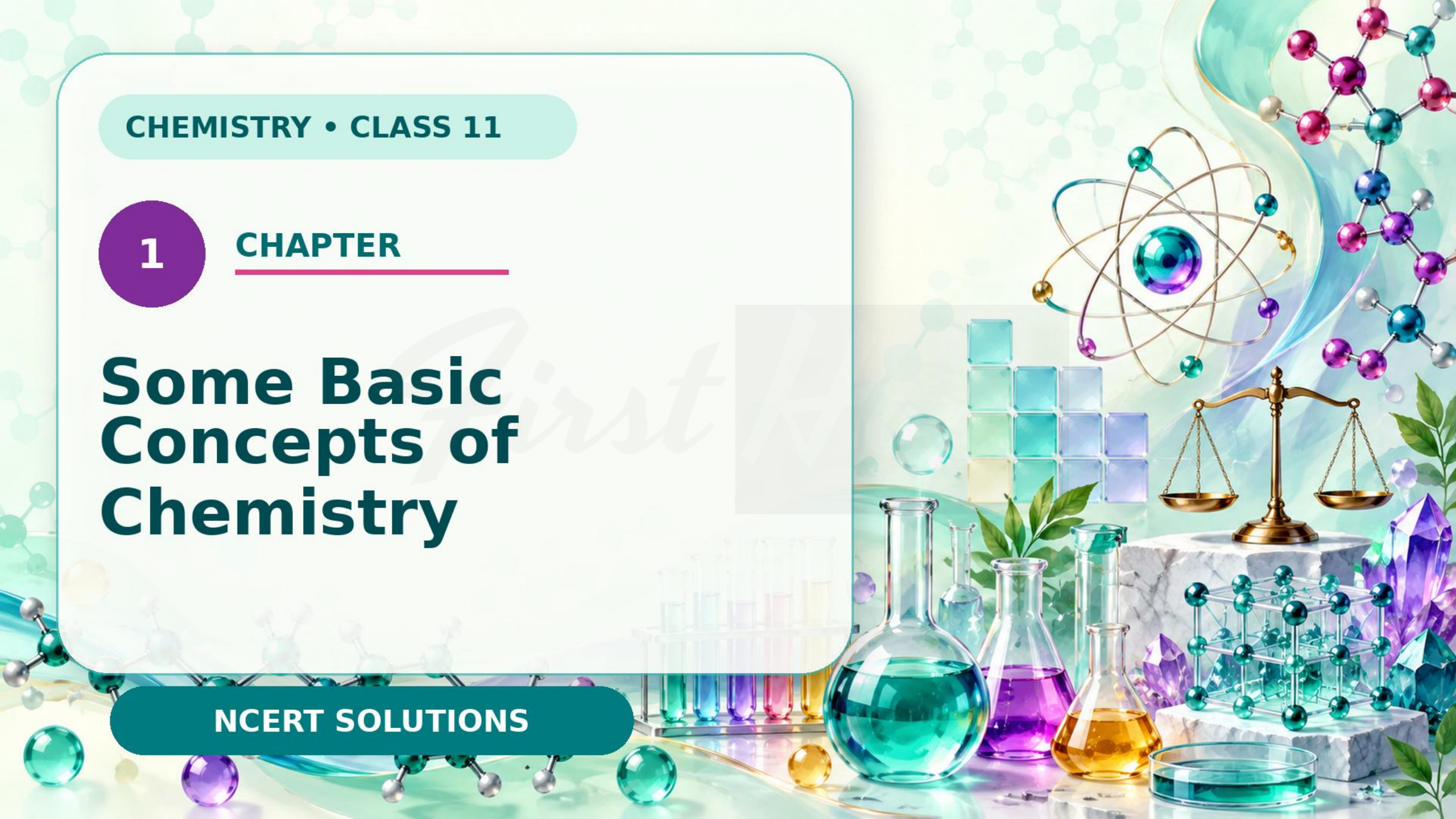
CHEMISTRY • CLASS 11

1

CHAPTER

Some Basic Concepts of Chemistry

NCERT SOLUTIONS



EXERCISE SOLUTIONS: 1.1-1.2

1.1 Calculate the molar mass of
(i) H_2O (ii) CO_2 (iii) CH_4 .

Answer:

Add the atomic masses of all atoms in one formula.

$$\text{H} = 1.008, \text{C} = 12.01, \text{O} = 16.00 \text{ g mol}^{-1}$$

$$\text{(i) } \text{H}_2\text{O}: 2(1.008) + 16.00 = 18.02 \text{ g mol}^{-1}$$

$$\text{(ii) } \text{CO}_2: 12.01 + 2(16.00) = 44.01 \text{ g mol}^{-1}$$

$$\text{(iii) } \text{CH}_4: 12.01 + 4(1.008) = 16.04 \text{ g mol}^{-1}$$

1.2 Calculate the mass per cent of each element in sodium sulphate, Na_2SO_4 .

Answer:

$$M = 2(22.99) + 32.06 + 4(16.00) = 142.04 \text{ g mol}^{-1}$$

$$\text{Mass percentage} = \text{mass contribution} \div \text{molar mass} \times 100.$$

$$\text{Na} = \frac{45.98}{142.04} \times 100 = 32.37\%$$

$$\text{S} = \frac{32.06}{142.04} \times 100 = 22.57\%$$

$$\text{O} = \frac{64.00}{142.04} \times 100 = 45.06\%$$

EXERCISE SOLUTIONS: 1.3

1.3 Determine the empirical formula of an oxide of iron containing 69.9% iron and 30.1% oxygen by mass.

Answer: Take a 100 g sample. Convert each elemental mass to moles.

Element	Mass / g	Moles	Ratio
Fe	69.9	$69.9 / 55.85 = 1.252$	$1.252 / 1.252 = 1$
O	30.1	$30.1 / 16.00 = 1.881$	$1.881 / 1.252 \approx 1.5$

Multiply both ratios by two to get whole numbers.

$$\text{Fe} : \text{O} = 1 : 1.5 = 2 : 3$$

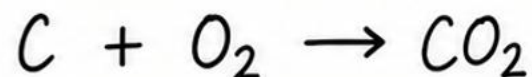
$$\text{Empirical formula} = \text{Fe}_2\text{O}_3$$

Use the atomic mass of oxygen because the ratio counts atoms in the oxide.

EXERCISE SOLUTIONS: 1.4

- 1.4 Calculate the amount of CO_2 produced when (i) 1 mol carbon burns in air, (ii) 1 mol carbon burns in 16 g dioxygen, and (iii) 2 mol carbon burn in 16 g dioxygen.

Answer:



The mole ratio is one to one to one. The limiting reactant fixes the amount of product.

(i) Air supplies excess O_2 : $n(\text{CO}_2) = 1 \text{ mol}$; $m(\text{CO}_2) = 1 \times 44 = 44 \text{ g}$.

(ii) $n(\text{O}_2) = \frac{16}{32} = 0.50 \text{ mol}$. O_2 is limiting. $n(\text{CO}_2) = 0.50 \text{ mol}$; $m(\text{CO}_2) = 0.50 \times 44 = 22 \text{ g}$.

(iii) Again $n(\text{O}_2) = 0.50 \text{ mol}$, so O_2 is limiting. $n(\text{CO}_2) = 0.50 \text{ mol}$; $m(\text{CO}_2) = 22 \text{ g}$.

Adding more carbon cannot increase CO_2 when the oxygen amount stays fixed.

EXERCISE SOLUTIONS: 1.5-1.6

- 1.5 Calculate the mass of CH_3COONa needed for 500 mL of 0.375 M solution.
Molar mass = $82.0245 \text{ g mol}^{-1}$.

Answer: First find the required moles.

$$V = 500 \text{ mL} = 0.500 \text{ L}$$

$$n = M \times V = 0.375 \times 0.500 = 0.1875 \text{ mol}$$

$$m = n \times \text{molar mass}$$

$$m = 0.1875 \times 82.0245 = 15.3796 \text{ g}$$

$$\text{Required mass} \approx 15.4 \text{ g}$$

- 1.6 Find the molarity of HNO_3 with density 1.41 g mL^{-1} and 69% HNO_3 by mass.

Answer: Take one litre of the solution.

$$\text{Mass of solution} = 1.41 \times 1000 = 1410 \text{ g}$$

$$\text{Mass of } \text{HNO}_3 = \frac{69}{100} \times 1410 = 972.9 \text{ g}$$

$$\text{Molar mass of } \text{HNO}_3 = 63.0 \text{ g mol}^{-1}$$

$$M = \frac{\frac{972.9}{63.0}}{1.00} \approx 15.4 \text{ mol L}^{-1}$$

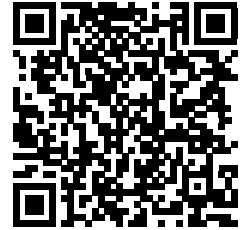


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