Grade 11 Chemistry

CBC NOTES | Terms 1, 2 and 3 - notes for 6 topics

IN EVERY TOPIC

- What you will learn, the key words and key terms explained.
- Clear notes under headings and subheadings, with labelled pictures and step-by-step procedures.
- Comparison tables, and advantages and disadvantages where methods are compared.
- Key points to remember, then structured questions to practise.
- All the answers and a glossary at the back.

TOPICS IN THESE NOTES

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A TASTE OF THE FIRST TOPIC

The Mole

Atoms are far too small to count, yet chemists need to know exactly how much of each substance reacts. The mole solves this problem: it lets us count particles by weighing them. In this topic we learn what a mole is, how it links mass, number of particles, gas volume and concentration, how to find the formulae of compounds, how to prepare standard solutions, how to carry out acid-base and redox titrations, and how redox titrations are used in industry.

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Strand 1.0 Inorganic Chemistry | Term 1

1.1 The Mole

WHAT YOU WILL LEARN

By the end of this topic you should be able to:

- describe the mole as a unit of measurement
- determine the empirical and molecular formula of compounds
- prepare standard solutions from solid and liquid substances
- determine the concentration of solutions
- determine the amount of substances in chemical reactions

KEY WORDS

mole, Avogadro's constant, relative atomic mass, relative formula mass, molar mass, empirical formula, molecular formula, water of crystallisation, standard solution, molarity, concentration, dilution, titration, titre, end point, indicator, redox, molar gas volume.

Atoms are far too small to count, yet chemists need to know exactly how much of each substance reacts. The mole solves this problem: it lets us count particles by weighing them. In this topic we learn what a mole is, how it links mass, number of particles, gas volume and concentration, how to find the formulae of compounds, how to prepare standard solutions, how to carry out acid-base and redox titrations, and how redox titrations are used in industry.

1. The mole as a unit of measurement

Mole (mol)
The SI unit for the amount of substance: the amount that contains 6.0 × 10^23 particles (atoms, molecules or ions).

Avogadro's constant (L)
The number of particles in one mole: 6.0 × 10^23 (more exactly 6.02 × 10^23).

Relative atomic mass (Ar)
The average mass of an atom of an element compared with one-twelfth of the mass of a carbon-12 atom. It has no units.

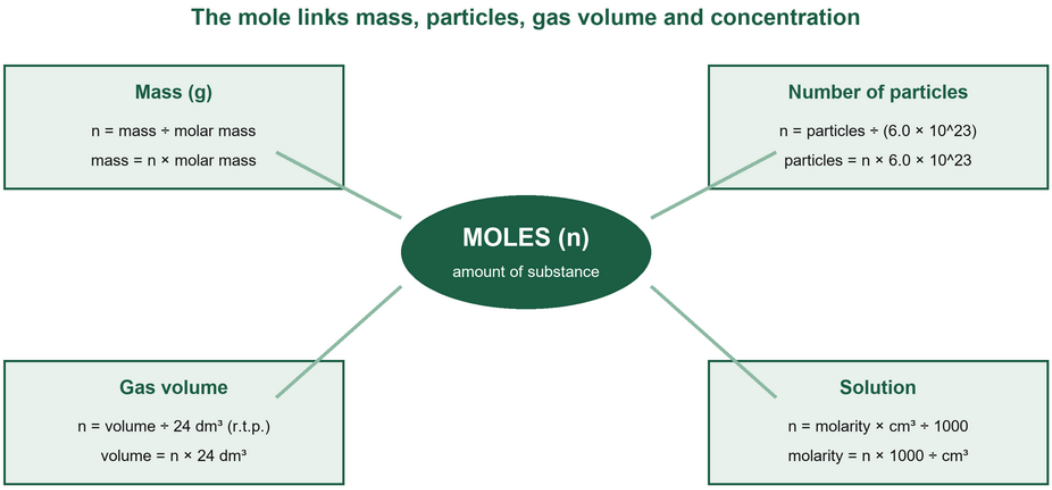
Relative formula mass (Mr)
The sum of the relative atomic masses of all the atoms in a formula. The MOLAR MASS is Mr in grams per mole (g/mol).

Just as a dozen eggs and a dozen melons contain the same number of items but have different masses, one mole of carbon (12 g) and one mole of magnesium (24 g) contain the same number of atoms.

Relative atomic mass from isotopes
Chlorine is 75% chlorine-35 and 25% chlorine-37.
Ar: (35 × 75 + 37 × 25) ÷ 100, which is (2625 + 925) ÷ 100 = 35.5.

Relative formula masses
Water, H2O: 1 × 2 + 16 = 18.
Calcium carbonate, CaCO3: 40 + 12 + 16 × 3 = 100.
Sodium carbonate, Na2CO3: 23 × 2 + 12 + 16 × 3 = 106.
Calcium hydroxide, Ca(OH)2: 40 + (16 + 1) × 2 = 74.

2. Converting between mass, moles, particles and gas volume



The mole links mass, particles, gas volume and concentration

Mole conversions

To find	Use	Example
Moles from mass	moles = mass ÷ molar mass	4 g of NaOH: 4 ÷ 40 = 0.1 mol
Mass from moles	mass = moles × molar mass	0.25 mol CaCO3: 0.25 × 100 = 25 g
Particles from moles	particles = moles × 6.0 × 10^23	0.5 mol CO2: 3.0 × 10^23 molecules
Moles of a gas	moles = volume ÷ molar gas volume	4.8 dm³ at r.t.p.: 4.8 ÷ 24 = 0.2 mol

Molar gas volume
The volume occupied by one mole of any gas: 24 dm³ at room temperature and pressure (r.t.p.) and 22.4 dm³ at standard temperature and pressure (s.t.p., 0 °C and 1 atmosphere).

Counting atoms in water
Moles of water in 9 g: 9 ÷ 18 = 0.5 mol of molecules.
Each molecule has 3 atoms: 0.5 × 3 = 1.5 mol of atoms.
Number of atoms: 1.5 × 6.0 × 10^23 gives 9.0 × 10^23 atoms.

EXAM METHOD
Always write the formula, work out the molar mass, convert the given quantity to moles, then convert moles to what is asked. Give units at every step.

3. Empirical and molecular formulae

Empirical formula
The simplest whole-number ratio of atoms of each element in a compound.

Molecular formula
The actual number of atoms of each element in one molecule; a whole-number multiple of the empirical formula.

Finding an empirical formula

1. Write the mass (or percentage) of each element.
2. Divide each by its relative atomic mass to get moles.
3. Divide every answer by the smallest number of moles.
4. If a ratio ends in .5, multiply all by 2; if it ends in .33 or .67, multiply by 3.

Magnesium oxide by experiment
Magnesium 0.24 g burnt to give 0.40 g of oxide; oxygen 0.40 - 0.24 = 0.16 g.
Mg: 0.24 ÷ 24 = 0.01 mol; O: 0.16 ÷ 16 = 0.01 mol. Ratio 1 : 1, so MgO.

From percentages, then the molecular formula
A compound is 40% C, 6.7% H and 53.3% O, with Mr 180.
C: 40 ÷ 12 = 3.33; H: 6.7 ÷ 1 = 6.7; O: 53.3 ÷ 16 = 3.33. Ratio 1 : 2 : 1, so CH_2O (mass 30).
n: 180 ÷ 30 = 6, so the molecular formula is $C_6H_{12}O_6$.

Water of crystallisation
Hydrated copper(II) sulphate, 2.50 g, gave 1.60 g of anhydrous salt on heating; water 2.50 - 1.60 = 0.90 g.
$CuSO_4$ (M 159.5): 1.60 ÷ 159.5 = 0.010 mol; water 0.90 ÷ 18 = 0.05 mol. Ratio 1 : 5, so $CuSO_4.5H_2O$.

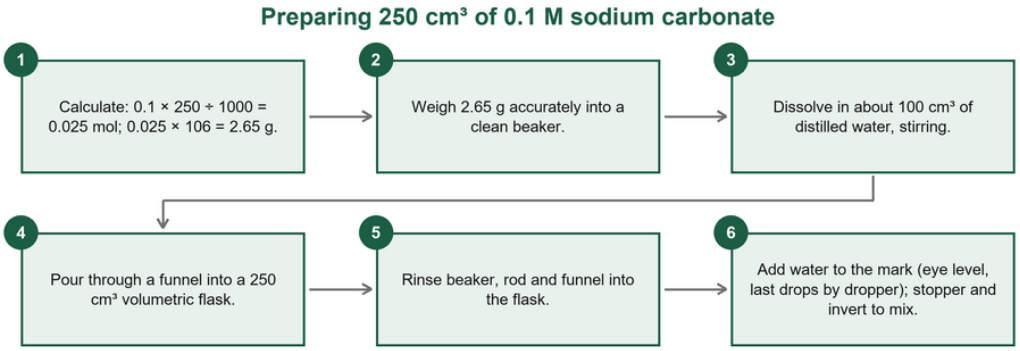
PRACTICAL POINTS
Clean magnesium ribbon with sandpaper, lift the crucible lid only slightly so smoke is not lost, and heat to constant mass so the reaction is complete.

4. Standard solutions, concentration and dilution

Standard solution
A solution whose concentration is accurately known.

Molarity (M)
Concentration in moles per cubic decimetre (mol/dm³). 1 dm³ is 1000 cm³.

- Moles in a solution = molarity × volume in cm³ ÷ 1000.
- Concentration in g/dm³ = molarity × molar mass.
- Dilution: moles stay the same, so M1 × V1 (before) equals M2 × V2 (after).



Preparing a standard solution from a solid

Dilution

50 cm³ of 2 M HCl is diluted to 500 cm³. New concentration: 2 × 50 ÷ 500 = 0.2 M.

Volume of 2 M HCl needed to make 250 cm³ of 0.1 M: 0.1 × 250 ÷ 2 = 12.5 cm³.

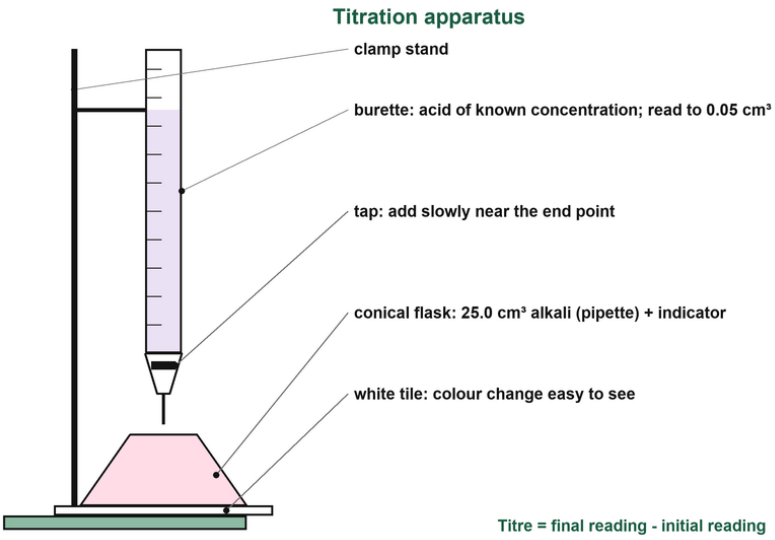
SAFETY

Always add acid to water, never water to concentrated acid. Use a pipette filler, never your mouth. Sodium hydroxide is not used as a primary standard because it absorbs water and carbon dioxide from the air.

5. Acid-base titrations

Titration

A method of finding the volume of one solution that reacts exactly with a known volume of another solution.



Titration apparatus

Carrying out a titration

1. Rinse the burette with the acid and the pipette with the alkali (not with water).
2. Pipette 25.0 cm³ of alkali into a conical flask; add 2-3 drops of indicator.
3. Run acid from the burette while swirling, drop by drop near the end point (colour just changes).
4. Record the initial and final burette readings to 0.05 cm³; titre = final - initial.
5. Repeat until titres agree within 0.1 cm³ and average them.

Finding the concentration of sodium hydroxide

NaOH + HCl -> NaCl + H2O (ratio 1 : 1). 25.0 cm³ of NaOH needed 20.0 cm³ of 0.1 M HCl.
Moles of HCl: 0.1 × 20 ÷ 1000 = 0.002, so moles of NaOH: 0.002.
Concentration of NaOH: 0.002 × 1000 ÷ 25 = 0.08 M.

A 1 : 2 ratio

H2SO4 + 2NaOH -> Na2SO4 + 2H2O. 25.0 cm³ of 0.1 M NaOH needed 12.5 cm³ of H2SO4.
Moles NaOH: 0.1 × 25 ÷ 1000 = 0.0025; moles H2SO4: 0.0025 ÷ 2 = 0.00125.
Concentration of H2SO4: 0.00125 × 1000 ÷ 12.5 = 0.1 M.

6. Redox titrations, mole ratios and their applications

Oxidation and reduction

Oxidation is loss of electrons (or gain of oxygen); reduction is gain of electrons (or loss of oxygen). A redox reaction has both.

Acidified potassium manganate(VII), purple MnO4-, oxidises iron(II) and is itself reduced to almost colourless Mn2+: MnO4- + 5Fe2+ + 8H+ -> Mn2+ + 5Fe3+ + 4H2O. It is self-indicating: the end point is the first permanent pale pink colour.

Redox titration

25.0 cm³ of Fe2+ solution needed 20.0 cm³ of 0.02 M KMnO4.
Moles MnO4-: 0.02 × 20 ÷ 1000 = 0.0004; moles Fe2+: 0.0004 × 5 = 0.002.
Concentration of Fe2+: 0.002 × 1000 ÷ 25 = 0.08 M, which is 0.08 × 56 = 4.48 g/dm³ of iron.

Reacting masses and gas volumes

CaCO3 -> CaO + CO2. From 50 g of CaCO3: 50 ÷ 100 = 0.5 mol, giving 0.5 × 56 = 28 g of CaO.
and 0.5 × 24 = 12 dm³ of CO2 at r.t.p.
Mg + 2HCl -> MgCl2 + H2. From 0.12 g of Mg: 0.12 ÷ 24 = 0.005 mol of H2, which is 0.005 × 24000 = 120 cm³.

Applications of redox titrations

Industry	Use of redox titrations
Pharmaceuticals	Checking iron in iron tablets and vitamin C in tablets and juices
Mining	Finding the iron content of ores
Water treatment	Measuring dissolved oxygen and chemical oxygen demand (pollution)
Bleach manufacture	Checking chlorine content (iodine and sodium thiosulphate)
Food and drinks	Measuring sulphur dioxide preservative

WASTE AND SAFETY

Neutralised solutions can be diluted and washed away; collect excess reagents and coloured manganate waste in labelled containers for the teacher. KMnO4 stains skin, so wear gloves and goggles.

KEY POINTS TO REMEMBER

- One mole contains 6.0 × 10^23 particles; its mass in grams equals the relative formula mass.
- Moles = mass ÷ molar mass; particles = moles × 6.0 × 10^23; one mole of gas occupies 24 dm³ at r.t.p.
- Empirical formulae come from mass or percentage data; the molecular formula is a whole-number multiple found using Mr.
- Standard solutions are made accurately in volumetric flasks; moles = molarity × cm³ ÷ 1000; dilution keeps moles constant.
- Titrations use the balanced equation's mole ratio to find unknown concentrations.
- In redox titrations MnO4- reacts with Fe2+ in a 1 : 5 ratio and is self-indicating; such titrations are used for quality control in industry.

ANSWER THESE

1. Calculate the number of moles in 11.7 g of sodium chloride (Na 23, Cl 35.5). (2 marks)

2. How many molecules are present in 64 g of oxygen gas (O 16, L = 6.0 × 10^23)? (2 marks)

3. A hydrocarbon contains 80% carbon and has a relative molecular mass of 30. Find (a) its empirical formula (b) its molecular formula. (4 marks)

4. Describe how you would prepare 250 cm³ of 0.1 M sodium carbonate solution (Na2CO3, Mr 106). (4 marks)

5. 20.0 cm³ of 0.1 M sulphuric acid neutralised 25.0 cm³ of sodium hydroxide solution. Calculate the concentration of the sodium hydroxide. (3 marks)

6. Why is no indicator needed in titrations with acidified potassium manganate(VII)? (2 marks)

7. What volume of carbon dioxide at r.t.p. is produced when 10 g of calcium carbonate is heated strongly? (Ca 40, C 12, O 16; molar gas volume 24 dm³) (3 marks)

8. Give two industrial applications of redox titrations. (2 marks)

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