

Introduction to the Fundamentals of Chemistry

An Interactive Textbook for the Sciences

From atoms and the mole to energy, equilibrium and the nucleus, built around worked examples and hands-on simulations.

1	Eighteen chapters, one for each group of the periodic table.																18			
1 H	2											13	14	15	16	17	18 He			
3 Li	4 Be	Unit I The language of chemistry					Unit II Reactions and energy					5 B	6 C	7 N	8 O	9 F	10 Ne			
11 Na	12 Mg	Unit III Periodic table and bonding					Unit IV States of matter					13 Al	14 Si	15 P	16 S	17 Cl	18 Ar			
19 K	20 Ca	Unit V Dynamics and equilibrium					Unit VI Advanced topics					28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe			
55 Cs	56 Ba	57–71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn			
87 Fr	88 Ra	89–103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og			
57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu						
89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr						

Nazeem Edwards, PhD

Print edition with worked solutions to all 119 exercises

Interactive edition: chemtextbook.neetapps.com



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2026



Introduction to the Fundamentals of Chemistry

An Interactive Textbook for the Sciences — Print Edition

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This print edition accompanies the interactive online edition of Introduction to the Fundamentals of Chemistry, a NEET-apps textbook with simulations, calculators, practice quizzes, a searchable glossary and a periodic table. Simulation screenshots in this book show the online edition, which is available at <https://chemtextbook.neetapps.com>.

Cover design, figures and interactive simulations: NEET-apps.

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Introduction to the Fundamentals of Chemistry

An Interactive Textbook for the Sciences — Print Edition

Nazeem Edwards, PhD

18 chapters · 44 sections · 81 additional worked examples · 119 end-of-chapter exercises

Answers and worked solutions are in the appendix.

2026 Edition

Interactive online edition: chemtextbook.neetapps.com



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Chapter 1: Matter, Measurement, and the Scientific Method

Unit I: The Language of Chemistry

1.1 The Nature of Chemistry

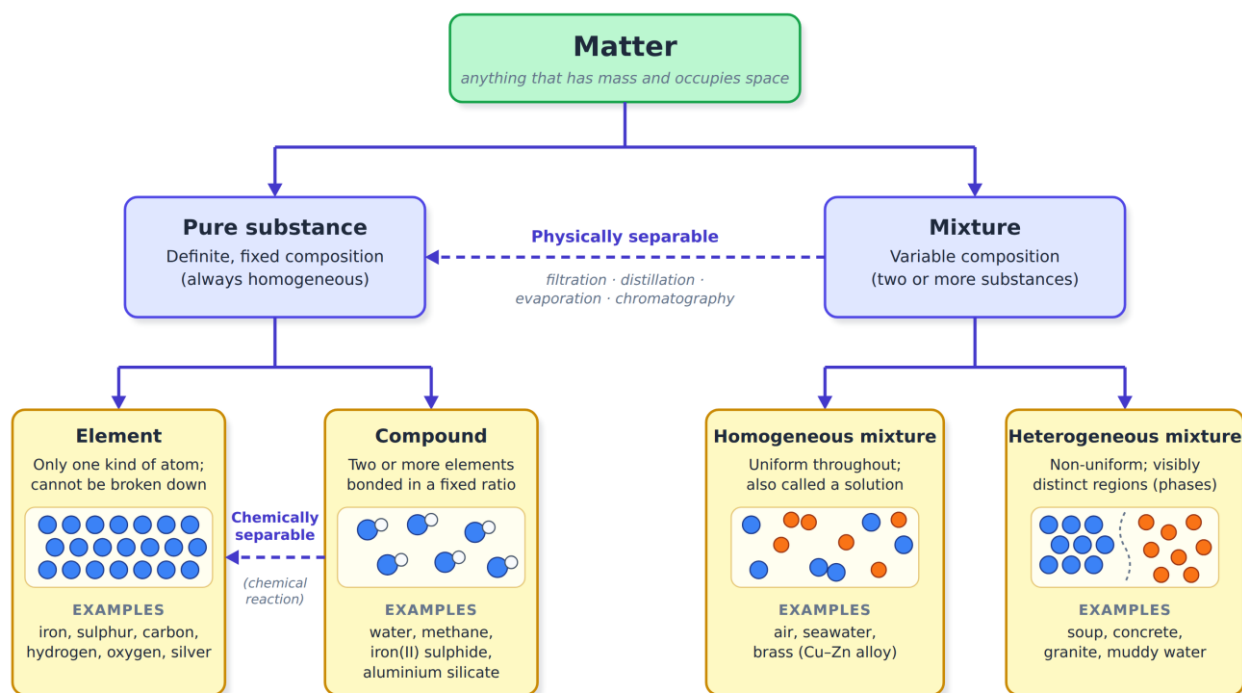
Chemistry is the study of matter—its composition, structure, properties, and the changes it undergoes. From pharmaceuticals to polymers, chemistry underpins modern civilisation.

Definition: *Matter* is anything that has mass and occupies space.

The Three Levels of Chemical Description

Level	Description	Example
Macroscopic	What we observe and measure	Water boils at 100°C
Particulate	Atoms, molecules, ions	H ₂ O molecules gain energy
Symbolic	Equations and notation	H ₂ O(l) → H ₂ O(g)

1.2 Classification of Matter



Classification of matter. The small panels show the particle level: identical atoms (element), identical molecules of bonded atoms (compound), two kinds of particles evenly mixed (homogeneous mixture) and separate regions of each kind (heterogeneous mixture).

WORKED EXAMPLE 1.1

Classify: (a) Seawater (b) Helium gas (c) Table salt (NaCl) (d) Concrete

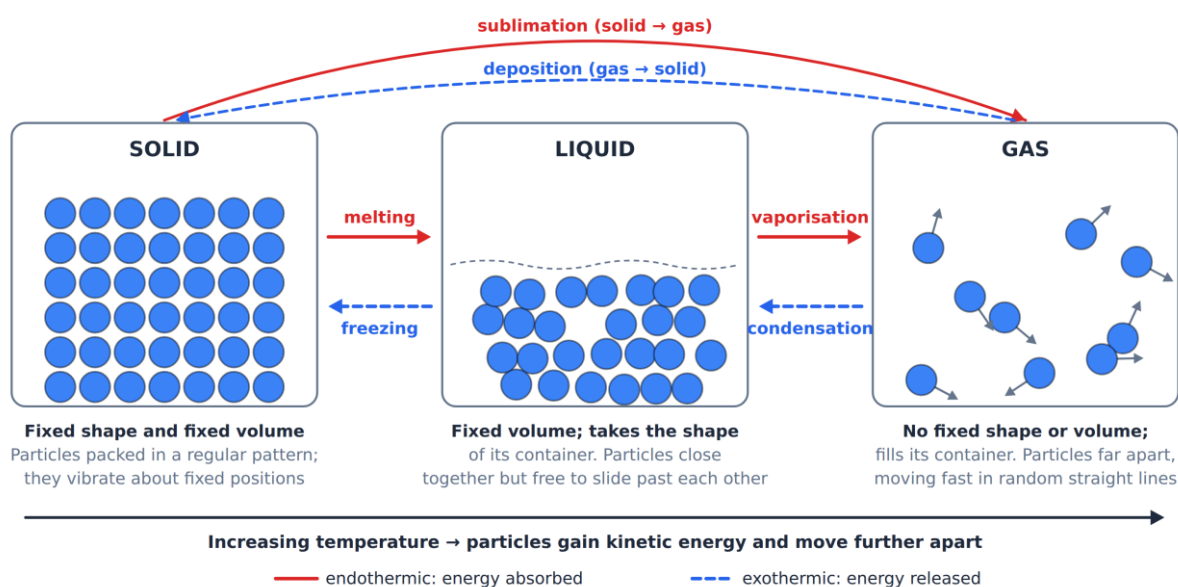
SOLUTION

(a) Seawater → Homogeneous mixture (a solution: salts are dissolved uniformly in the water; filtered seawater looks the same throughout)

(b) Helium → Element (only He atoms)

(c) NaCl → Compound (Na and Cl chemically bonded, fixed 1:1 ratio)

(d) Concrete → Heterogeneous mixture (visibly different components)

1.3 States of Matter

The three states at the particle level. Solid → liquid → gas changes absorb energy (red); the reverse changes release it (blue).

Phase Change	Direction	Energy
Melting	Solid → Liquid	Endothermic
Freezing	Liquid → Solid	Exothermic
Vaporisation	Liquid → Gas	Endothermic
Condensation	Gas → Liquid	Exothermic
Sublimation	Solid → Gas	Endothermic
Deposition	Gas → Solid	Exothermic

1.4 Significant Figures & Measurement

Rules for Significant Figures

#	Rule	Example
1	All nonzero digits are significant	345 → 3 sig figs
2	Zeros between nonzero digits are significant	304 → 3 sig figs
3	Leading zeros are NOT significant	0.0045 → 2 sig figs
4	Trailing zeros after decimal ARE significant	2.300 → 4 sig figs
5	Mult/Div: fewest sig figs wins	$3.45 \times 2.1 = 7.2$ (2 sf)
6	Add/Sub: fewest decimal places wins	$12.3 + 6.70 = 19.0$

WORKED EXAMPLE 1.2

Calculate density: mass = 23.456 g, volume = 8.7 cm³

SOLUTION

$\rho = 23.456 / 8.7 = 2.6960... \rightarrow$ Volume has 2 sig figs $\rightarrow \rho = 2.7 \text{ g/cm}^3$

1.5 Temperature Conversions

$$T_K = T_C + 273.15$$

$$T_F = (9/5)T_C + 32$$

$$T_C = (5/9)(T_F - 32)$$

Interactive tool in the online version: Temperature Converter — chemtextbook.neetapps.com.

Additional Explanation: Measurement Is a Model, Not Just a Number

A measured value is meaningful only when the quantity, unit, and precision are clear. Density links mass to volume, while significant figures communicate the resolution justified by the measurements. In a subtraction used to find a volume difference, decimal-place precision matters before the final division is rounded. Keep guard digits during the calculation and round only the final result.

Worked Example: Density by Water Displacement

A 44.5 g metal sample raises the water level from 25.0 mL to 30.0 mL. Find its density.

Displaced volume = 30.0 mL – 25.0 mL = 5.0 mL.

Density = mass ÷ volume = 44.5 g ÷ 5.0 mL = 8.90 g/mL (calculator value).

The volume difference has two significant figures, so report 8.9 g/mL (equivalently 8.9 g/cm³).

Check: the sample's density is greater than water's density, so a compact sample of this metal would generally sink in water.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 1.A: Sig figs in a multi-step calculation

Calculate $(12.03 + 4.1) \times 2.00$ to the correct number of significant figures.

SOLUTION

Step 1 (addition): $12.03 + 4.1 = 16.13$. The answer is limited to *one decimal place* (from 4.1), so it is really 16.1 — 3 sig figs. Keep the extra digit for now.

Step 2 (multiplication): $16.13 \times 2.00 = 32.26$. Both factors have 3 sig figs → round to 3 sig figs.

Answer: 32.3

Rounding too early ($16.1 \times 2.00 = 32.2$) gives a different answer. Round only at the end.

Worked Example 1.B: Density and unit conversion

A block of aluminium measures $2.00 \text{ cm} \times 3.00 \text{ cm} \times 5.00 \text{ cm}$ and has a mass of 81.0 g. Find its density in g/cm^3 and in kg/m^3 .

SOLUTION

$$V = 2.00 \times 3.00 \times 5.00 = 30.0 \text{ cm}^3$$

$$\rho = m/V = 81.0 \text{ g} \div 30.0 \text{ cm}^3 = 2.70 \text{ g/cm}^3$$

$$\text{Convert: } 2.70 \text{ g/cm}^3 \times (1 \text{ kg} / 1000 \text{ g}) \times (100 \text{ cm} / 1 \text{ m})^3 = 2.70 \times 10^6 / 10^3 = 2.70 \times 10^3 \text{ kg/m}^3$$

$$\text{Shortcut: } 1 \text{ g/cm}^3 = 1000 \text{ kg/m}^3.$$

Worked Example 1.C: Temperature conversion with very cold liquids

Liquid nitrogen boils at 77 K. Express this in $^{\circ}\text{C}$ and $^{\circ}\text{F}$.

SOLUTION

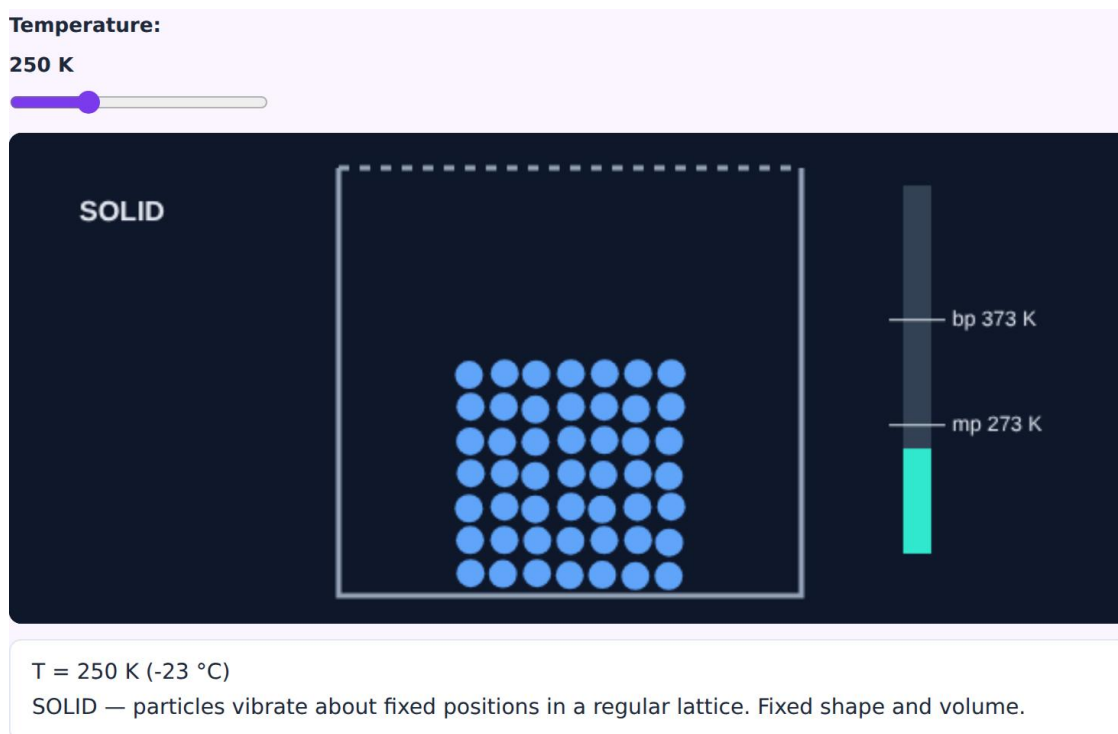
$$T_{\text{C}} = 77 - 273.15 = -196 \text{ }^{\circ}\text{C}$$

$$T_{\text{F}} = (9/5)(-196.15) + 32 = -321 \text{ }^{\circ}\text{F}$$

Note that a *change* of 1 K equals a change of $1 \text{ }^{\circ}\text{C}$, but a change of $1 \text{ }^{\circ}\text{C}$ equals a change of $1.8 \text{ }^{\circ}\text{F}$.

Interactive Simulation: Particles in the Three States

Change the temperature of a sample of water and watch the particles. Notice how the motion and spacing change at the melting point (273 K) and boiling point (373 K).



Companion-app task — Particles in the Three States: change temperature slowly and record what changes in particle motion versus particle spacing. Distinguish observations from explanations.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 1 Quiz

Q1. Which of the following is a compound?

- A. Oxygen gas (O_2)
- B. Salt water
- C. Sodium chloride (NaCl)
- D. Air

Q2. How many significant figures are in 0.00450?

- A. 2
- B. 3
- C. 4
- D. 5

Q3. Convert 72°F to Celsius:

- A. 22.2°C
- B. 25.0°C
- C. 40.0°C
- D. 90.0°C

End-of-Chapter Exercises

Level 1 — Recall and understanding

1.1 Classify each as an element, compound, homogeneous mixture or heterogeneous mixture: (a) brass (b) distilled water (c) oxygen gas (d) muddy water.

1.2 How many significant figures are in: (a) 0.02050 (b) 1.200×10^3 (c) 5000?

Level 2 — Application

1.3 Calculate $4.52 \times 0.0130 \div 2.1$ and give the answer to the correct number of significant figures.

Answer: _____

1.4 A 25.4 g metal sample is dropped into a measuring cylinder. The water level rises from 20.0 mL to 23.2 mL. Calculate the density in g/cm^3 .

Answer: _____ g/cm^3

1.5 Convert -40°C to $^\circ\text{F}$.

Answer: _____ $^\circ\text{F}$

1.6 Normal body temperature is 98.6°F . Convert this to kelvin.

Answer: _____ K

Level 3 — Challenge

1.7 Explain, in terms of particles and energy, why sweating cools the body.

Chapter 2: Atoms, Molecules, and Ions

Unit I: The Language of Chemistry

2.1 Atomic Structure

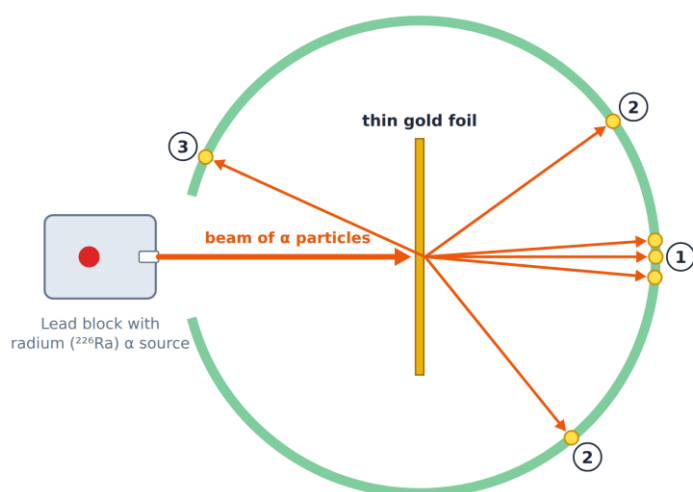
Particle	Symbol	Charge	Mass (amu)	Location
Proton	p^+	+1	1.0073	Nucleus
Neutron	n^0	0	1.0087	Nucleus
Electron	e^-	-1	0.000549	Electron cloud

Key Definitions:

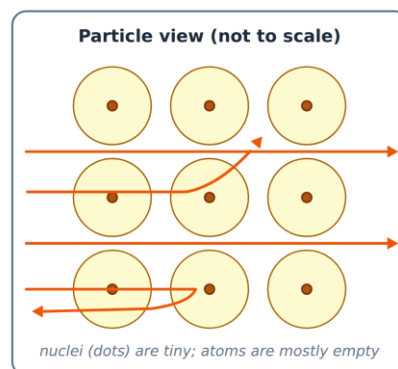
- **Atomic number (Z)** = number of protons (defines the element)
- **Mass number (A)** = protons + neutrons
- **Isotopes** = same Z, different number of neutrons

${}^A_Z X$ (e.g., ${}^{35}_{17}\text{Cl}$ has 17 protons, 18 neutrons)

2.2 Rutherford's Gold Foil Experiment



- ① most α particles pass straight through · ② a few are deflected
③ about 1 in 8000 bounces back — a flash on the ZnS screen marks each hit



Conclusion: an atom is mostly empty space.
Its mass and positive charge are concentrated in a tiny, dense nucleus (about 10^{-15} m across).

Geiger and Marsden fired α particles at gold foil. The rare large deflections could only be caused by a tiny, massive, positive nucleus.

Key Finding: ~1 in 8000 alpha particles bounced back, proving a concentrated positive charge exists in the centre.

2.3 Ions and Naming Compounds

Ion Formation

Type	Process	Examples
Cation (+)	Loses electrons	Na^+ , Ca^{2+} , Al^{3+} , Fe^{2+}

Type	Process	Examples
Anion (-)	Gains electrons	Cl^- , O^{2-} , N^{3-} , S^{2-}

Naming Rules

Rule	Example
Cation first, then anion	NaCl = sodium chloride
Monatomic anions end in -ide	Cl^- = chloride, O^{2-} = oxide
Transition metals: Roman numerals	FeCl_3 = iron(III) chloride
Polyatomic ions keep their names	SO_4^{2-} = sulphate, NH_4^+ = ammonium

WORKED EXAMPLE

Name: (a) Na_2O (b) FeCl_3 (c) $(\text{NH}_4)_2\text{SO}_4$

SOLUTION

(a) sodium oxide (Group 1, no Roman numeral needed)

(b) $3 \text{Cl}^- \rightarrow \text{Fe}$ must be +3 $\rightarrow$ iron(III) chloride

(c) ammonium sulphate

Additional Explanation: Keep Proton Identity Separate from Ion Charge

The proton number defines the element and does not change when an ordinary atom becomes an ion. The neutron number changes the isotope, while the electron number determines the net charge. For a neutral atom, electrons equal protons; for a positive ion, electrons have been removed; for a negative ion, electrons have been added. This bookkeeping prevents a common error: changing the element symbol when only the charge changes.

Worked Example: Build a Calcium-40 Ion

Write the particle counts and nuclide notation for a calcium-40 ion with a 2+ charge.

Calcium has atomic number 20, so it has 20 protons.

Neutrons = mass number – atomic number = $40 - 20 = 20$.

A 2+ ion has lost two electrons: $20 - 2 = 18$ electrons. Write the symbol as ${}^{40}_{20}\text{Ca}^{2+}$.

Check: 20 protons and 18 electrons give a net charge of +2; the element remains calcium because its proton number is still 20.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 2.A: Average atomic mass from isotopes

Chlorine is 75.78% ^{35}Cl (34.969 u) and 24.22% ^{37}Cl (36.966 u). Calculate the relative atomic mass of chlorine.

SOLUTION

$$A_r = \Sigma(\text{fraction} \times \text{isotope mass})$$

$$= (0.7578)(34.969) + (0.2422)(36.966)$$

$$= 26.500 + 8.953 = 35.45 \text{ u}$$

The value is closer to 35 than to 37 because ^{35}Cl is more abundant.

Worked Example 2.B: Counting particles in an ion

How many protons, neutrons and electrons are in $^{56}\text{Fe}^{3+}$?

SOLUTION

Z for Fe = 26 $\rightarrow$ 26 protons

Neutrons = $A - Z = 56 - 26 = 30$ neutrons

Charge +3 means 3 electrons lost: $26 - 3 = 23$ electrons

Worked Example 2.C: Writing formulas from names

Write formulas for: (a) calcium nitrate (b) aluminium sulphate (c) copper(II) hydroxide.

SOLUTION

Balance the charges so the total is zero ("criss-cross" the charge numbers):

(a) Ca^{2+} and $\text{NO}_3^- \rightarrow$ need 2 $\text{NO}_3^- \rightarrow \text{Ca}(\text{NO}_3)_2$

(b) Al^{3+} and $\text{SO}_4^{2-} \rightarrow 2 \times (+3) = 3 \times (-2) \rightarrow \text{Al}_2(\text{SO}_4)_3$

(c) Cu^{2+} and $\text{OH}^- \rightarrow \text{Cu}(\text{OH})_2$

Use brackets around a polyatomic ion when you need more than one of it.

Interactive Simulation: Build an Atom

Add or remove protons, neutrons and electrons. The element, mass number, charge and nuclide symbol update as you build.

Protons $-$ 6 $+$ Neutrons $-$ 6 $+$ Electrons $-$ 6 $+$

$^{12}_6\text{C}$
Carbon ($Z = 6$)
 Mass number $A = 6 + 6 = 12$
 Charge $= 6 - 6 = 0 \rightarrow$ neutral atom
 Nucleus: ☒ a stable isotope
 Shells drawn with a simple 2, 8, 8 model.

Companion-app task — *Build an Atom*: add 20 protons and 20 neutrons, then remove two electrons. Compare the displayed element, mass number, charge, and nuclide symbol with the worked example.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 2 Quiz

Q1. How many neutrons are in $^{40}_{20}\text{Ca}$?

- A. 20
- B. 40
- C. 60
- D. 10

Q2. What is the correct formula for iron(III) oxide?

- A. FeO
- B. Fe_2O_3
- C. Fe_3O_2
- D. FeO_3

Q3. Isotopes of an element differ in:

- A. Number of protons
- B. Number of electrons
- C. Number of neutrons
- D. Atomic number

End-of-Chapter Exercises

Level 1 — Recall and understanding

2.1 Give the number of protons, neutrons and electrons in $^{39}\text{K}^+$.

2.2 Name: (a) MgCl_2 (b) CuO (c) NaHCO_3 (d) KMnO_4 .

2.3 Write formulas for: (a) ammonium phosphate (b) iron(II) sulphate (c) calcium carbonate.

Level 2 — Application

2.4 An ion has 16 protons, 16 neutrons and 18 electrons. Write its full nuclide symbol.

2.5 Gallium is 60.11% ^{69}Ga (68.926 u) and 39.89% ^{71}Ga (70.925 u). Calculate its relative atomic mass.

Answer: _____ u

Level 3 — Challenge

2.6 Boron has two isotopes: ^{10}B (10.013 u) and ^{11}B (11.009 u). Its relative atomic mass is 10.81. Calculate the percentage abundance of ^{11}B .

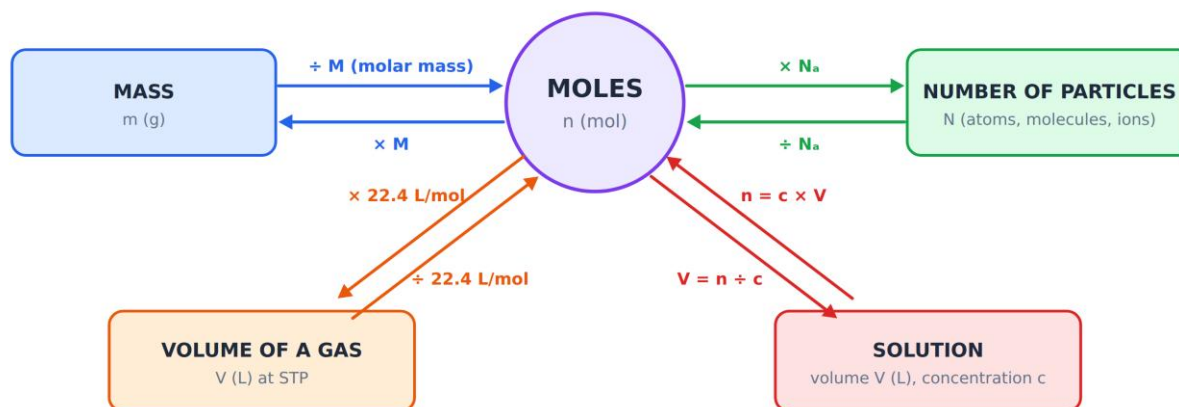
Answer: _____ %

Chapter 3: Stoichiometry and the Mole Concept

Unit I: The Language of Chemistry

3.1 The Mole and Avogadro's Number

Definition: One mole (mol) = 6.022×10^{23} entities (Avogadro's number, N_a)



$$n = m/M \cdot N = n \times N_a \quad (N_a = 6.022 \times 10^{23} \text{ mol}^{-1}) \cdot V = n \times 22.4 \text{ L/mol (gas at STP: } 0^\circ\text{C, } 1 \text{ atm)} \cdot n = cV$$

Every conversion passes through moles. Choose the arrow that matches the quantity you have and the one you need.

$$n = \text{mass} / M \quad | \quad N = n \times N_a \quad | \quad M = \text{g/mol}$$

WORKED EXAMPLE 3.1

Calculate the molar mass of $\text{Ca}_3(\text{PO}_4)_2$

SOLUTION

Ca: $3 \times 40.08 = 120.24$

P: $2 \times 30.97 = 61.94$

O: $8 \times 16.00 = 128.00$

Total = 310.18 g/mol

3.2 Balancing Chemical Equations

WORKED EXAMPLE 3.2

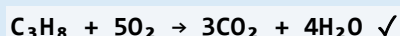
Balance: $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$

SOLUTION (STEP BY STEP)

1. Balance C: $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow 3\text{CO}_2 + \text{H}_2\text{O}$

2. Balance H: $\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}$

3. Balance O: Right has $3(2)+4(1)=10 \rightarrow$ need 5O_2



INTERACTIVE: EQUATION BALANCER

Try balancing these equations:

**3.3 Limiting Reactant & Percent Yield**

Key Concept: The limiting reactant is consumed first and determines maximum product. The excess reactant remains after reaction.

WORKED EXAMPLE 3.3

10.0 g Al reacts with 25.0 g O₂: $4\text{Al} + 3\text{O}_2 \rightarrow 2\text{Al}_2\text{O}_3$

SOLUTION

$$\text{mol Al} = 10.0/26.98 = 0.3706 \text{ mol}$$

$$\text{mol O}_2 = 25.0/32.00 = 0.7813 \text{ mol}$$

$$\text{O}_2 \text{ needed for all Al: } 0.3706 \times (3/4) = 0.2780 \text{ mol} < 0.7813 \rightarrow \text{Al is limiting}$$

$$\text{mol Al}_2\text{O}_3 = 0.3706 \times (2/4) = 0.1853 \text{ mol}$$

$$\text{Mass Al}_2\text{O}_3 = 0.1853 \times 101.96 = 18.9 \text{ g}$$

$$\% \text{ yield} = (\text{actual yield} / \text{theoretical yield}) \times 100\%$$

Additional Explanation: Stoichiometry Is a Mole-Ratio Map

A balanced chemical equation gives ratios of particles and moles, not direct ratios of masses. Convert each measured reactant to moles, compare the available amounts with the equation coefficients, and identify the limiting reactant before calculating product. The limiting reactant fixes the theoretical yield; any excess reactant remains after the limiting reactant is consumed.

Worked Example: Identify the Limiting Reactant

Aluminium reacts with chlorine: $2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$. If 5.40 g Al reacts with 10.6 g Cl₂, which reactant limits the product?

$$n(\text{Al}) = 5.40 \text{ g} \div 26.98 \text{ g mol}^{-1} = 0.200 \text{ mol. } n(\text{Cl}_2) = 10.6 \text{ g} \div 70.90 \text{ g mol}^{-1} = 0.149 \text{ mol.}$$

The equation requires $3/2 = 1.5 \text{ mol Cl}_2$ per mole Al. For 0.200 mol Al, 0.300 mol Cl₂ would be required, but only 0.149 mol is available.

Cl₂ is limiting. Product moles = $0.1495 \times (2/3) = 0.0997 \text{ mol AlCl}_3$ (using the unrounded amount of Cl₂); theoretical mass $\approx 0.0997 \times 133.33 = 13.3 \text{ g}$.

Check: the limiting reactant is the one that produces the smaller possible amount of product. Use unrounded intermediate values where possible.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 3.A: Empirical and molecular formula

A compound is 40.0% C, 6.71% H and 53.3% O by mass. Its molar mass is 180 g/mol. Find the empirical and molecular formulas.

SOLUTION

Assume 100 g, so the percentages become grams:

Element	Mass (g)	Moles	÷ smallest
C	40.0	$40.0/12.01 = 3.33$	1
H	6.71	$6.71/1.008 = 6.66$	2
O	53.3	$53.3/16.00 = 3.33$	1

Empirical formula: CH_2O ($M = 30.03 \text{ g/mol}$)

$n = 180/30.03 = 6 \rightarrow$ molecular formula $\text{C}_6\text{H}_{12}\text{O}_6$ (glucose)

Worked Example 3.B: Percent yield

15.0 g of N_2 reacts with excess H_2 : $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. Only 13.5 g of NH_3 is collected. Calculate the percent yield.

SOLUTION

$$n(\text{N}_2) = 15.0/28.02 = 0.5354 \text{ mol}$$

$$n(\text{NH}_3) = 2 \times 0.5354 = 1.071 \text{ mol (mole ratio 1 : 2)}$$

$$\text{Theoretical mass} = 1.071 \times 17.03 = 18.24 \text{ g}$$

$$\% \text{ yield} = 13.5/18.24 \times 100\% = 74.0\%$$

Worked Example 3.C: Preparing a standard solution

What mass of NaOH is needed to make 250 mL of a 0.150 mol/L solution?

SOLUTION

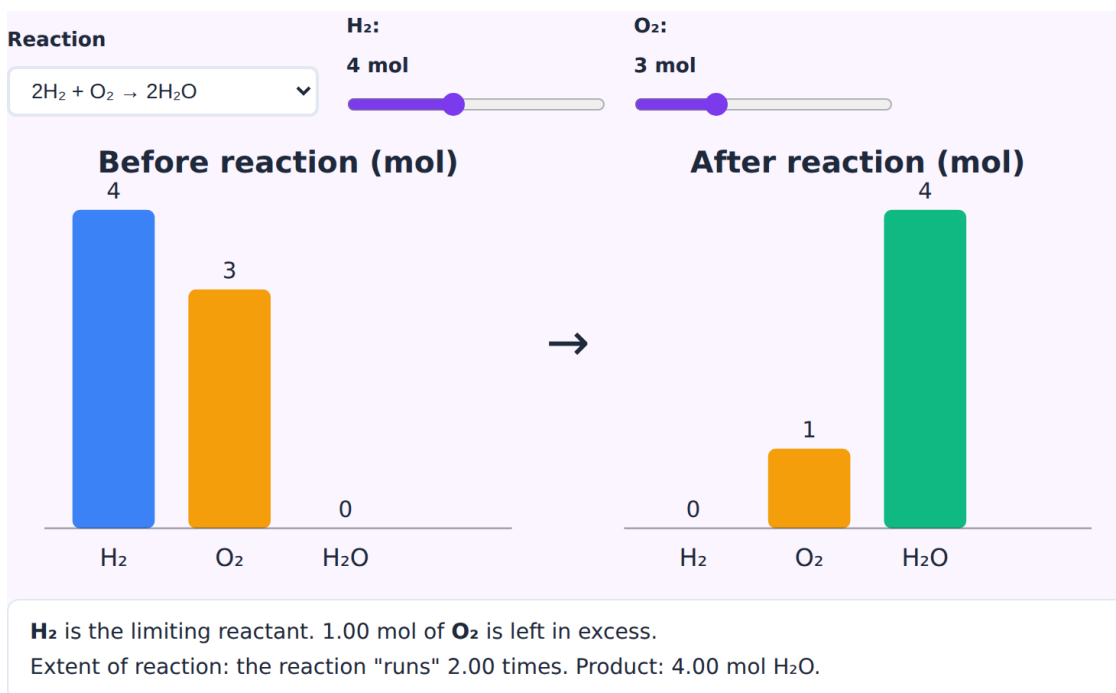
$$n = cV = 0.150 \text{ mol/L} \times 0.250 \text{ L} = 0.0375 \text{ mol}$$

$$m = nM = 0.0375 \times 40.00 = 1.50 \text{ g}$$

Dissolve in some water in a 250 mL volumetric flask, then make up to the mark.

Interactive Simulation: Limiting Reactant Lab

Choose a reaction and set the moles of each reactant. The bars show what is present before and after the reaction goes to completion.



Companion-app task — Limiting Reactant Lab: change one reactant amount at a time. Predict the limiting reactant before running the activity, then compare the predicted product amount with the result.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 3 Quiz

Q1. How many molecules are in 3.50 g of water?

- A. 1.17×10^{23}
- B. 6.02×10^{23}
- C. 3.50×10^{23}
- D. 1.94×10^{23}

Q2. In $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$, if you have 4 mol H₂ and 1 mol O₂, the limiting reactant is:

- A. H₂
- B. O₂
- C. H₂O
- D. Neither

Q3. The empirical formula of C₆H₁₂O₆ is:

- A. C₆H₁₂O₆
- B. C₃H₆O₃
- C. CH₂O
- D. C₂H₄O₂

End-of-Chapter Exercises

Level 1 — Recall and understanding

3.1 How many moles are in 50.0 g of CaCO₃ (M = 100.09 g/mol)?

Answer: _____ mol

3.2 Balance: $\text{Al} + \text{HCl} \rightarrow \text{AlCl}_3 + \text{H}_2$

Level 2 — Application

3.3 How many oxygen atoms are in 0.250 mol of H_2SO_4 ?

Answer: _____ atoms

3.4 Find the empirical formula of a hydrocarbon that is 85.6% C and 14.4% H by mass.

3.5 5.85 g of NaCl is dissolved to make 500 mL of solution. Calculate the concentration.

Answer: _____ mol/L

Level 3 — Challenge

3.6 5.00 g Al reacts with 10.0 g Cl_2 : $2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3$. What mass of AlCl_3 forms?

Answer: _____ g

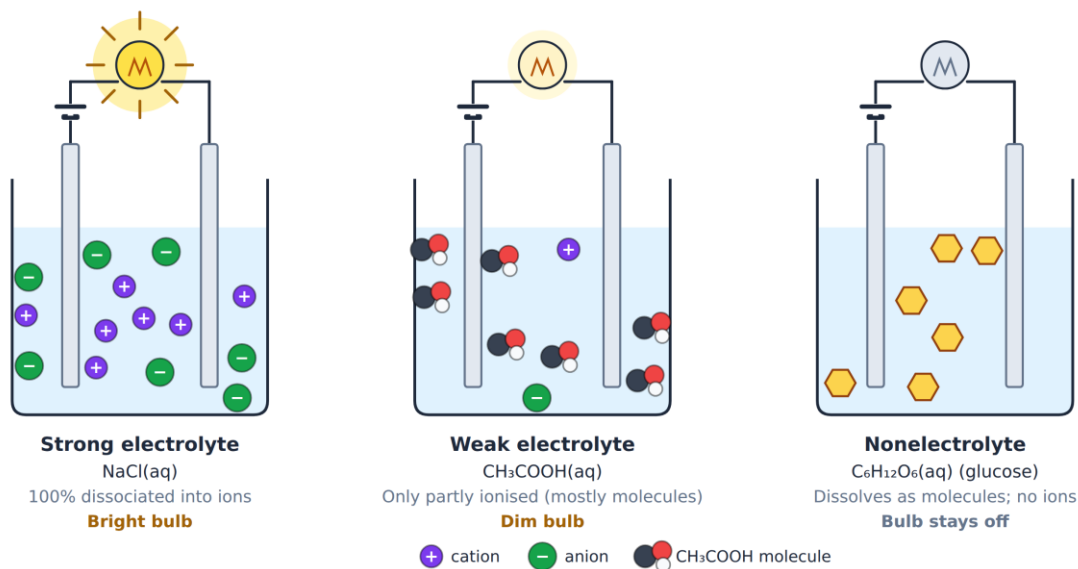
3.7 In the previous question, 10.9 g of AlCl_3 was actually collected. Calculate the percent yield.

Answer: _____ %

Chapter 4: Chemical Reactions in Aqueous Solution

Unit II: Reactions & Energy

4.1 Electrolytes



Only mobile ions carry current through a solution, so the brightness of the bulb shows how many ions are present.

Strong Acids to Memorise

HCl, HBr, HI, HNO₃, HClO₄, H₂SO₄ (first proton)

Strong Bases to Memorise

NaOH, KOH, LiOH, Ca(OH)₂, Sr(OH)₂, Ba(OH)₂

4.2 Precipitation Reactions

Solubility Rules

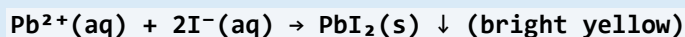
Always Soluble	Usually Insoluble (Exceptions)
Group 1, NH ₄ ⁺	OH ⁻ (except Group 1, Ba ²⁺)
NO ₃ ⁻ , ClO ₄ ⁻ , CH ₃ COO ⁻	CO ₃ ²⁻ , PO ₄ ³⁻ , S ²⁻ (except Grp 1, NH ₄ ⁺)
Cl ⁻ , Br ⁻ , I ⁻ (except Ag ⁺ , Pb ²⁺)	SO ₄ ²⁻ (except Ba ²⁺ , Sr ²⁺ , Pb ²⁺)

WORKED EXAMPLE

Predict: Pb(NO₃)₂(aq) + KI(aq) → ?

SOLUTION

PbI₂ is insoluble (Pb²⁺ exception with I⁻) → Yellow precipitate forms!



4.3 Redox Reactions

OIL RIG: Oxidation Is Loss, Reduction Is Gain (of electrons)

Oxidation Number Rules

Rule	Example
Free elements = 0	O ₂ , Fe, Na → 0
Monatomic ions = charge	Na ⁺ = +1, Cl ⁻ = -1
O is usually -2	H ₂ O, CO ₂ (except peroxides: -1)
H is usually +1	HCl, H ₂ O (except hydrides: -1)
Sum = 0 (compound) or charge (ion)	KMnO ₄ : Mn = +7

Additional Explanation: Molecular Equations and Net Ionic Equations Answer Different Questions

A molecular equation records compounds as formula units, while a complete ionic equation separates strong soluble electrolytes into ions. Ions that appear unchanged on both sides are spectator ions. Removing them produces the net ionic equation, which highlights the chemical change. State symbols matter: a precipitate is an insoluble solid formed from ions in solution.

Worked Example: Precipitation of Silver Chloride

Mix aqueous silver nitrate and sodium chloride. Write the balanced molecular and net ionic equations.

Molecular equation: $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$.

Complete ionic equation: $\text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$.

Cancel Na⁺ and NO₃⁻, which are unchanged spectators. Net ionic equation: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$.

Check: atoms and total charge balance in the net ionic equation (+1 on each side). The solid state symbol identifies the precipitate.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 4.A: Molecular, complete ionic and net ionic equations

Write all three equations for the reaction of AgNO₃(aq) with NaCl(aq).

SOLUTION

Molecular: $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$

Complete ionic: $\text{Ag}^+(\text{aq}) + \text{NO}_3^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{Na}^+(\text{aq}) + \text{NO}_3^-(\text{aq})$

Cancel the spectator ions (Na⁺ and NO₃⁻):

Net ionic: $\text{Ag}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \rightarrow \text{AgCl}(\text{s})$ (white precipitate)

Worked Example 4.B: Identifying oxidising and reducing agents

In $\text{Zn(s)} + \text{Cu}^{2+}(\text{aq}) \rightarrow \text{Zn}^{2+}(\text{aq}) + \text{Cu(s)}$, identify what is oxidised and reduced, and the oxidising and reducing agents.

SOLUTION

Oxidation numbers: $\text{Zn } 0 \rightarrow +2$ (increase) $\rightarrow$ Zn is oxidised: $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$

$\text{Cu } +2 \rightarrow 0$ (decrease) $\rightarrow \text{Cu}^{2+}$ is reduced: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$

The substance that is reduced is the oxidising agent (Cu^{2+}); the substance oxidised is the reducing agent (Zn).

Worked Example 4.C: Acid–base titration

25.0 mL of NaOH solution is neutralised by 18.6 mL of 0.100 mol/L HCl. Find the NaOH concentration.

SOLUTION

$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ (ratio 1 : 1)

$n(\text{HCl}) = 0.100 \times 0.0186 = 1.86 \times 10^{-3} \text{ mol} = n(\text{NaOH})$

$c(\text{NaOH}) = 1.86 \times 10^{-3} / 0.0250 = 0.0744 \text{ mol/L}$

Worked Example 4.D: Oxidation number in a polyatomic ion

Find the oxidation number of Cr in $\text{Cr}_2\text{O}_7^{2-}$.

SOLUTION

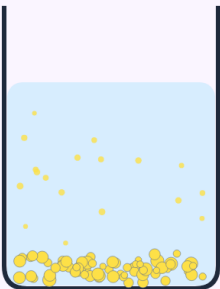
$2x + 7(-2) = -2 \rightarrow 2x = 12 \rightarrow x = +6$

Interactive Simulation: Precipitation Lab

Mix a solution containing the chosen cation with one containing the chosen anion. Predict first, then click Mix to see whether a precipitate forms.

Cation solution	Anion solution	Your prediction
Pb ²⁺ (as nitrate) ▼	I ⁻ (as sodium salt) ▼	— choose — ▼





A **bright yellow precipitate of PbI₂** forms.

Net ionic equation:
 $\text{Pb}^{2+}(\text{aq}) + 2\text{I}^{-}(\text{aq}) \rightarrow \text{PbI}_2(\text{s})$

Spectator ions: NO₃⁻ and Na⁺.

Companion-app task — Precipitation Lab: predict whether a solid forms, identify the ions that combine, and then write the net ionic equation for the observed result.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 4 Quiz

Q1. Which is a strong electrolyte?

- A. CH_3COOH
- B. $\text{C}_6\text{H}_{12}\text{O}_6$
- C. NaCl
- D. CH_3OH

Q2. The oxidation number of Mn in KMnO_4 is:

- A. +2
- B. +5
- C. +7
- D. +4

Q3. A precipitate forms when mixing:

- A. $\text{NaCl} + \text{KNO}_3$
- B. $\text{BaCl}_2 + \text{Na}_2\text{SO}_4$
- C. $\text{NaOH} + \text{HCl}$
- D. $\text{KBr} + \text{NaNO}_3$

End-of-Chapter Exercises

Level 1 — Recall and understanding

- 4.1** Classify as strong, weak or non-electrolyte: HNO_3 , NH_3 , $\text{C}_2\text{H}_5\text{OH}$, KBr .
- 4.2** Give the oxidation number of S in: H_2S , SO_2 , H_2SO_4 , S_8 .

Level 2 — Application

- 4.3** Write the net ionic equation for $\text{BaCl}_2(\text{aq}) + \text{K}_2\text{SO}_4(\text{aq})$.
- 4.4** In $2\text{Mg} + \text{O}_2 \rightarrow 2\text{MgO}$, identify the oxidising agent and the reducing agent.
- 4.5** 50.0 mL of 0.200 mol/L AgNO_3 is mixed with excess NaCl . What mass of AgCl precipitates?

Answer: _____ g

Level 3 — Challenge

- 4.6** 20.0 mL of H_2SO_4 is neutralised by 32.4 mL of 0.150 mol/L NaOH . Calculate the H_2SO_4 concentration.
- Answer: _____ mol/L
- 4.7** Balance the half-reaction $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$ in acidic solution.

Chapter 5: Thermochemistry

Unit II: Reactions & Energy

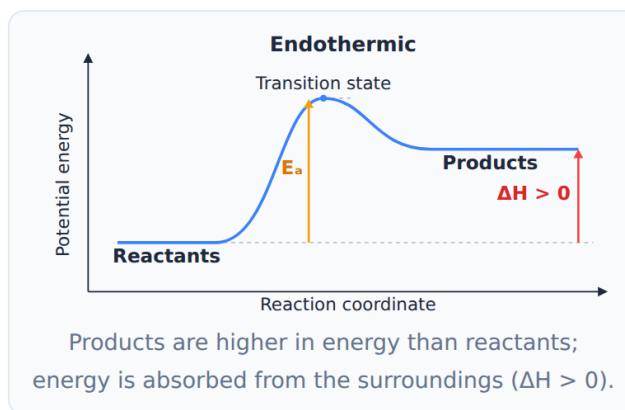
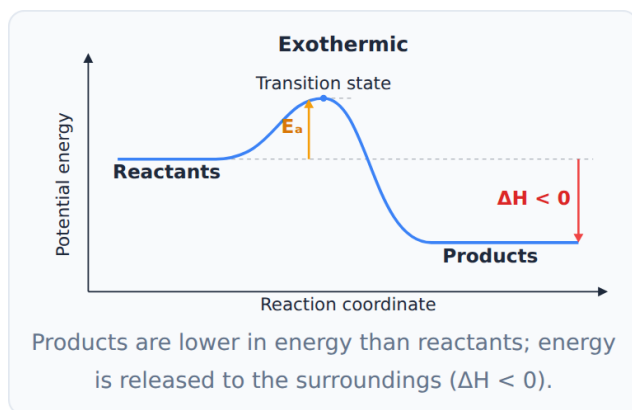
5.1 Energy and Enthalpy

$\Delta E = q + w$ (First Law: energy is conserved)

$q = mc\Delta T$ (heat transfer)

$\Delta H < 0$: Exothermic (releases heat)

$\Delta H > 0$: Endothermic (absorbs heat)



5.2 Calorimetry

WORKED EXAMPLE

5.00 g NaOH dissolved in 100.0 g water. T rises 20.0°C → 33.5°C. Find ΔH .

SOLUTION

$$q_{\text{soln}} = (105.0 \text{ g})(4.18 \text{ J/g}\cdot^{\circ}\text{C})(13.5^{\circ}\text{C}) = 5926 \text{ J} = 5.93 \text{ kJ}$$

$$q_{\text{rxn}} = -5.93 \text{ kJ (exothermic)}$$

$$\text{mol NaOH} = 5.00/40.00 = 0.125 \text{ mol}$$

$$\Delta H = -5.93/0.125 = -47.4 \text{ kJ/mol}$$

Interactive tool in the online version: Calorimetry Calculator — chemtextbook.neetapps.com.

5.3 Hess's Law

Hess's Law: ΔH for a reaction is the same whether it occurs in one step or multiple steps. ΔH is a state function.

WORKED EXAMPLEFind ΔH for: $\text{C(s)} + \frac{1}{2}\text{O}_2(\text{g}) \rightarrow \text{CO(g)}$ Given: $\text{C} + \text{O}_2 \rightarrow \text{CO}_2$, $\Delta H_1 = -393.5 \text{ kJ}$ Given: $\text{CO} + \frac{1}{2}\text{O}_2 \rightarrow \text{CO}_2$, $\Delta H_2 = -283.0 \text{ kJ}$ **SOLUTION**Reverse eq. 2: $\text{CO}_2 \rightarrow \text{CO} + \frac{1}{2}\text{O}_2$, $\Delta H = +283.0 \text{ kJ}$ Add: $(-393.5) + (+283.0) = -110.5 \text{ kJ}$ **Additional Explanation: Track the System and the Surroundings**

In calorimetry, define the system before assigning a sign to heat. If the reaction releases heat to the water or calorimeter, the surroundings gain energy and the reaction's heat is negative under the usual sign convention. The calculation $q = mc\Delta T$ gives the heat absorbed by the measured material; relate that value to the reaction with the opposite sign when heat exchange with the surroundings is the only relevant transfer.

Worked Example: Heat Absorbed by Water

A reaction warms 100.0 g of water from 20.0 °C to 25.0 °C. Use $c = 4.184 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1}$ to estimate the heat change.

$$\Delta T = 25.0 - 20.0 = +5.0 \text{ }^\circ\text{C}.$$

$$q_{\text{water}} = mc\Delta T = (100.0 \text{ g})(4.184 \text{ J g}^{-1} \text{ }^\circ\text{C}^{-1})(5.0 \text{ }^\circ\text{C}) = 2092 \text{ J} \approx 2.1 \text{ kJ}.$$

If heat exchange with the water is the only significant transfer, $q_{\text{reaction}} \approx -2.1 \text{ kJ}$: the reaction is exothermic.

Check: the water temperature rose, so the water gained heat; the reacting system must have supplied it. In real experiments, include calorimeter heat capacity if it is not negligible.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 5.A: Enthalpy change from enthalpies of formation

Use ΔH°_f values (kJ/mol: CH_4 -74.8, CO_2 -393.5, $\text{H}_2\text{O(l)}$ -285.8) to find ΔH° for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O(l)}$.

SOLUTION

$$\Delta H^\circ_{\text{rxn}} = \sum n\Delta H^\circ_f(\text{products}) - \sum n\Delta H^\circ_f(\text{reactants})$$

$$\text{Products: } (-393.5) + 2(-285.8) = -965.1 \text{ kJ}$$

$$\text{Reactants: } (-74.8) + 2(0) = -74.8 \text{ kJ } (\Delta H^\circ_f \text{ of an element in its standard state is 0})$$

$$\Delta H^\circ = -965.1 - (-74.8) = -890.3 \text{ kJ}$$

Worked Example 5.B: Finding the specific heat of a metal

50.0 g of a metal at 100.0 °C is placed in 100.0 g of water at 22.0 °C. The final temperature is 25.4 °C. Find c for the metal.

SOLUTION

Heat lost by metal = heat gained by water

$$q_{\text{water}} = (100.0)(4.18)(25.4 - 22.0) = 1421 \text{ J}$$

$$q_{\text{metal}} = -1421 \text{ J} = (50.0)(c)(25.4 - 100.0)$$

$$c = 1421 / (50.0 \times 74.6) = 0.381 \text{ J/(g}\cdot\text{°C)} \text{ — close to copper (0.385).}$$

Worked Example 5.C: Using bond enthalpies

Estimate ΔH for $\text{H}_2 + \text{Cl}_2 \rightarrow 2\text{HCl}$. Bond enthalpies (kJ/mol): H–H 436, Cl–Cl 242, H–Cl 431.

SOLUTION

$$\Delta H \approx \Sigma(\text{bonds broken}) - \Sigma(\text{bonds formed})$$

Broken: $436 + 242 = 678 \text{ kJ}$ (energy absorbed)

Formed: $2 \times 431 = 862 \text{ kJ}$ (energy released)

$$\Delta H \approx 678 - 862 = -184 \text{ kJ (exothermic)}$$

Interactive Simulation: Thermal Equilibrium Mixer

Bring two samples into contact in an insulated container. Heat flows from hot to cold until both reach the same final temperature.

Sample 1

Substance
Water (c = 4.18) ▼

Mass (g)
100

Initial T (°C)
20

Sample 2

Substance
Copper (c = 0.385) ▼

Mass (g)
50

Initial T (°C)
95

🔥 Bring into contact

Heat lost by the hotter sample = heat gained by the cooler sample

$$m_1c_1(T_f - T_1) + m_2c_2(T_f - T_2) = 0 \rightarrow T_f = (m_1c_1T_1 + m_2c_2T_2)/(m_1c_1 + m_2c_2)$$

$$T_f = (418.0 \times 20 + 19.3 \times 95) / (418.0 + 19.3) = \mathbf{23.3\text{ }^{\circ}\text{C}}$$

Heat transferred: **1380 J** from sample 2 to sample 1.

Notice: the sample with the larger heat capacity (m × c) changes temperature less.

Companion-app task — Thermal Equilibrium Mixer: change the starting temperatures or masses, predict the direction of heat transfer, and explain why the final temperature lies between the starting temperatures for an isolated two-object system.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 5 Quiz

Q1. An exothermic reaction has:

- A. $\Delta H > 0$
- B. $\Delta H < 0$
- C. $\Delta H = 0$
- D. $q = 0$

Q2. $q = mc\Delta T$. If $m=50\text{g}$, $c=4.18\text{ J/g}^\circ\text{C}$, $\Delta T=20^\circ\text{C}$, then $q =$

- A. 418 J
- B. 4180 J
- C. 41.8 kJ
- D. 4180 kJ

Q3. Hess's Law works because enthalpy is a:

- A. Path function
- B. State function
- C. Rate function
- D. Equilibrium function

End-of-Chapter Exercises

Level 1 — Recall and understanding

5.1 How much heat is needed to warm 250 g of water from 20.0°C to 95.0°C ?

Answer: _____ kJ

Level 2 — Application

5.2 $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{l})$, $\Delta H = -571.6\text{ kJ}$. How much heat is released when 10.0 g of H_2 burns?

Answer: _____ kJ

5.3 Find ΔH° for $\text{C}_2\text{H}_4(\text{g}) + 3\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$. ΔH°_f (kJ/mol): C_2H_4 +52.3, CO_2 -393.5, $\text{H}_2\text{O}(\text{l})$ -285.8.

Answer: _____ kJ

5.4 Given $\text{S} + \text{O}_2 \rightarrow \text{SO}_2$ ($\Delta H = -296.8\text{ kJ}$) and $2\text{SO}_2 + \text{O}_2 \rightarrow 2\text{SO}_3$ ($\Delta H = -197.8\text{ kJ}$), find ΔH for $\text{S} + 3/2\text{O}_2 \rightarrow \text{SO}_3$.

Answer: _____ kJ

Level 3 — Challenge

5.5 Estimate ΔH for $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}(\text{g})$ using bond enthalpies (kJ/mol): C-H 413, O=O 498, C=O 799, O-H 463.

Answer: _____ kJ

5.6 100 mL of 1.0 mol/L HCl and 100 mL of 1.0 mol/L NaOH, both at 21.0°C , are mixed. The temperature rises to 27.8°C . Find ΔH of neutralisation per mole of water formed (assume density 1.00 g/mL , $c = 4.18\text{ J/(g}^\circ\text{C)}$).

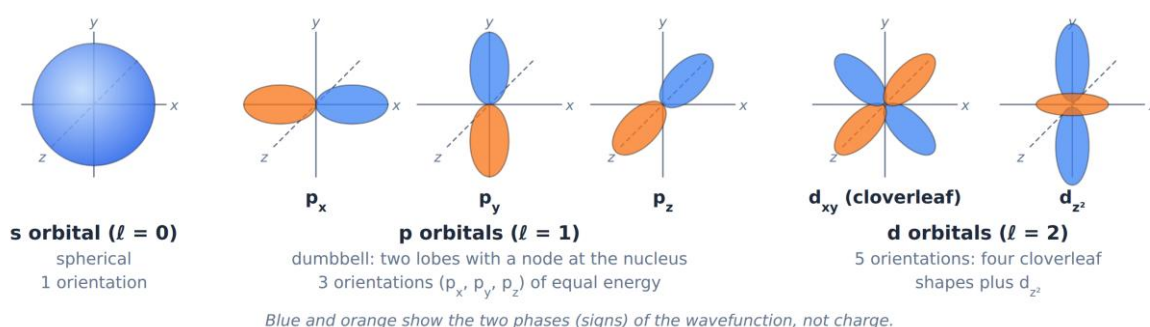
Answer: _____ kJ/mol

Chapter 6: Electronic Structure of Atoms

Unit II: Reactions & Energy

6.1 Quantum Numbers

Quantum Number	Symbol	Values	Describes
Principal	n	1, 2, 3, ...	Energy level, size
Angular momentum	ℓ	0 to ($n-1$)	Subshell shape (s,p,d,f)
Magnetic	m_ℓ	$-\ell$ to $+\ell$	Orbital orientation
Spin	m_s	$+\frac{1}{2}$, $-\frac{1}{2}$	Electron spin



Boundary-surface sketches enclosing most of the electron probability. Shape is set by ℓ ; the number of orientations is $2\ell + 1$.

6.2 Electron Configuration

Filling Order (Aufbau)

$1s \rightarrow 2s \rightarrow 2p \rightarrow 3s \rightarrow 3p \rightarrow 4s \rightarrow 3d \rightarrow 4p \rightarrow 5s \rightarrow 4d \rightarrow 5p \rightarrow 6s \rightarrow 4f \rightarrow 5d \dots$

Three Rules

Rule	Statement
Aufbau	Fill lowest energy orbitals first
Pauli Exclusion	Max 2 electrons per orbital (opposite spins)
Hund's Rule	In degenerate orbitals, fill singly before pairing

WORKED EXAMPLE

Write electron configuration for Fe ($Z=26$) and Fe^{3+}

SOLUTION

Fe: $[\text{Ar}] 4s^2 3d^6$

Fe^{3+} (remove $3e^-$): remove 4s first $\rightarrow [\text{Ar}] 3d^5$

$3d^5$: $[\uparrow][\uparrow][\uparrow][\uparrow][\uparrow]$ — five unpaired electrons

Additional Explanation: Electron Configurations Describe Occupancy, Not Fixed Orbits

Electron configurations summarise how electrons occupy atomic orbitals. The Aufbau principle gives a useful filling guide, the Pauli exclusion principle limits an orbital to two electrons with opposite spins, and Hund's rule favours singly occupying equal-energy orbitals before pairing. Orbital diagrams are especially useful when counting unpaired electrons and explaining magnetic behaviour.

Worked Example: Sulphur's Valence Electrons

Write the ground-state electron configuration of sulphur ($Z = 16$) and determine the number of unpaired 3p electrons.

Place 16 electrons in order: $1s^2 2s^2 2p^6 3s^2 3p^4$, or $[\text{Ne}] 3s^2 3p^4$.

The three 3p orbitals receive one electron each before pairing begins. The fourth 3p electron pairs in one orbital.

The 3p occupancy is therefore one pair and two singly occupied orbitals: two unpaired electrons.

Check: the total electron count is $2 + 2 + 6 + 2 + 4 = 16$. This is a ground-state configuration, not a picture of electrons travelling in circular paths.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 6.A: Energy of a photon and the hydrogen spectrum

The red line in the hydrogen spectrum has $\lambda = 656 \text{ nm}$. Find the energy of one photon and of one mole of photons. Show that it matches the $n = 3 \rightarrow n = 2$ transition.

SOLUTION

$$c = \lambda \nu \quad E = h\nu = hc/\lambda$$

$$E = (6.626 \times 10^{-34})(2.998 \times 10^8)/(656 \times 10^{-9}) = 3.03 \times 10^{-19} \text{ J}$$

$$\text{Per mole: } 3.03 \times 10^{-19} \times 6.022 \times 10^{23} = 182 \text{ kJ/mol}$$

$$\text{Bohr model: } \Delta E = 2.18 \times 10^{-18} (1/2^2 - 1/3^2) = 2.18 \times 10^{-18} \times 0.1389 = 3.03 \times 10^{-19} \text{ J}$$

Worked Example 6.B: Allowed quantum numbers

List all possible sets of quantum numbers for an electron in a 3p orbital. Is ($n = 2$, $\ell = 2$) allowed?

SOLUTION

3p $\rightarrow n = 3$, $\ell = 1$, so $m_\ell = -1, 0, +1$ and $m_s = +\frac{1}{2}$ or $-\frac{1}{2}$: 6 possible sets (3 orbitals $\times$ 2 electrons).

($n = 2$, $\ell = 2$) is not allowed because ℓ can only go up to $n - 1 = 1$.

Worked Example 6.C: Orbital diagram and magnetism

Draw the orbital diagram for N. How many unpaired electrons does it have?

SOLUTION

N ($Z = 7$): $1s^2 2s^2 2p^3$

1s $\uparrow\downarrow$ 2s $\uparrow\downarrow$ 2p $\uparrow \uparrow \uparrow$

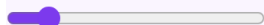
Hund's rule: the three 2p electrons occupy separate orbitals with parallel spins $\rightarrow$ 3 unpaired electrons, so N atoms are paramagnetic.

Interactive Simulation: Electron Configuration Builder

Slide through the elements and watch the orbitals fill in Aufbau order. Look out for the exceptions such as Cr and Cu.

Atomic number Z:

8



Oxygen (O), $Z = 8$

Configuration: $1s^2 2s^2 2p^4$

Noble-gas shorthand: $[\text{He}]2s^2 2p^4$

Unpaired electrons: 2 $\rightarrow$ paramagnetic (attracted by a magnet)

1s $\uparrow\downarrow$ 2s $\uparrow\downarrow$ 2p $\uparrow\downarrow \uparrow \uparrow$

Companion-app task — Electron Configuration Builder: construct sulphur's configuration and compare the orbital filling with the written notation. Try a nearby element to see how the 3p subshell fills.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 6 Quiz

Q1. How many electrons can the $n=3$ shell hold?

- A. 8
- B. 18
- C. 32
- D. 6

Q2. The electron configuration of Cu ($Z=29$) is:

- A. $[\text{Ar}]4s^23d^7$
- B. $[\text{Ar}]4s^13d^{10}$
- C. $[\text{Ar}]4s^23d^9$
- D. $[\text{Ar}]4s^24p^9$

Q3. Which quantum number determines orbital shape?

- A. n
- B. ℓ
- C. m_ℓ
- D. m_s

End-of-Chapter Exercises

Level 1 — Recall and understanding

6.1 Write the full electron configuration of S and of the S^{2-} ion.

6.2 What is the maximum number of electrons in (a) the $n = 4$ shell (b) a 4f subshell?

Level 2 — Application

6.3 How many unpaired electrons are in Mn^{2+} ?

Answer: _____

6.4 Which sets of quantum numbers (n , ℓ , m_ℓ , m_s) are NOT allowed? (a) 3, 2, -2, $+\frac{1}{2}$ (b) 2, 2, 0, $-\frac{1}{2}$ (c) 4, 0, 0, $+\frac{1}{2}$ (d) 3, 1, -2, $+\frac{1}{2}$

6.5 Calculate the frequency of light with wavelength 500 nm.

Answer: _____ Hz

6.6 Calculate the energy of one photon of violet light with $\lambda = 400$ nm.

Answer: _____ J

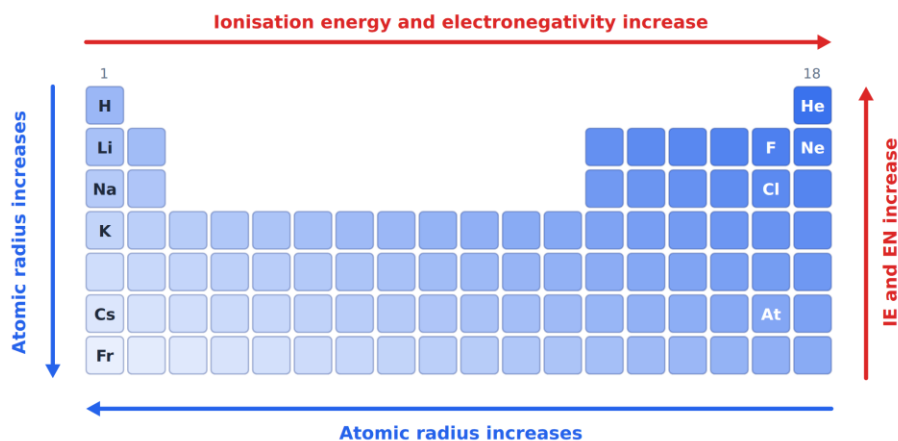
Level 3 — Challenge

6.7 Write the configuration of Cr and explain why it does not follow the Aufbau order.

Chapter 7: Periodic Properties of the Elements

Unit III: Periodic Table & Bonding

7.1 Periodic Trends



Across a period: same outer shell, but Z_{eff} increases $\rightarrow$ electrons held closer and more tightly.
Down a group: more shells and more shielding $\rightarrow$ outer electrons further away and easier to remove.

Top right: F is the most electronegative element; He has the highest first IE (noble gases are usually not given EN values).
Bottom left: Cs and Fr have the largest atoms and the lowest IE and EN. Shading shows the general trend in first IE.

General trends only — see the exceptions below ($\text{Be} > \text{B}$ and $\text{N} > \text{O}$ for first ionisation energy).

Exceptions to Ionisation Energy Trend

B < Be: Be has filled $2s^2$ (extra stability). B's electron in $2p$ is easier to remove.

O < N: N has half-filled $2p^3$ (Hund's stability). O's paired $2p^4$ has e^- repulsion.

7.2 Effective Nuclear Charge

$$Z_{\text{eff}} = Z - S \quad (Z = \text{atomic number}, S = \text{shielding})$$

Across a period: Z increases, S stays roughly constant $\rightarrow Z_{\text{eff}}$ increases $\rightarrow$ electrons pulled closer $\rightarrow$ radius decreases.

Down a group: new shells added $\rightarrow$ more shielding $\rightarrow$ radius increases despite more protons.

Additional Explanation: Periodic Trends Reflect Competing Effects

Across a period, nuclear charge increases while added electrons enter the same principal energy level; the effective attraction to the nucleus generally increases and atomic radius tends to decrease. Down a group, added occupied shells and shielding generally increase atomic size. Ionisation energy follows related but not perfectly smooth trends because subshell structure and electron pairing can create exceptions.

Worked Example: Compare Sodium and Magnesium

Predict which atom is smaller and which generally has the higher first ionisation energy: Na or Mg.

Na and Mg are in the same period, so their valence electrons occupy the third principal shell.

Mg has one more proton than Na, while shielding does not increase enough to cancel the extra nuclear attraction. Mg's valence electrons are held more tightly, so Mg is generally smaller.

Removing a valence electron from Mg generally requires more energy; Mg has the higher first ionisation energy.

Check: use the trend to make a prediction, then consult actual values when a question asks for precise ordering because known exceptions exist.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 7.A: Ranking atomic radius

Arrange Mg, Na, Cl and S in order of decreasing atomic radius.

SOLUTION

All are in Period 3. Across a period Z_{eff} increases, so the radius decreases.

$\text{Na} > \text{Mg} > \text{S} > \text{Cl}$

Worked Example 7.B: Isoelectronic species

Rank O^{2-} , F^- , Na^+ and Mg^{2+} by radius.

SOLUTION

All have 10 electrons ([Ne]). With the same number of electrons, more protons pull the electrons in more strongly.

Z : O 8, F 9, Na 11, Mg 12 $\rightarrow \text{O}^{2-} > \text{F}^- > \text{Na}^+ > \text{Mg}^{2+}$

Worked Example 7.C: Using successive ionisation energies

Element X has $\text{IE}_1 = 738$, $\text{IE}_2 = 1451$, $\text{IE}_3 = 7733$ kJ/mol. Which group is X in?

SOLUTION

The huge jump between IE_2 and IE_3 (more than $5\times$) means the third electron comes from a filled inner shell.

X has 2 valence electrons $\rightarrow$ Group 2 (it is Mg).

Worked Example 7.D: Estimating effective nuclear charge

Use $Z_{\text{eff}} \approx Z - (\text{number of core electrons})$ to compare Na and Cl.

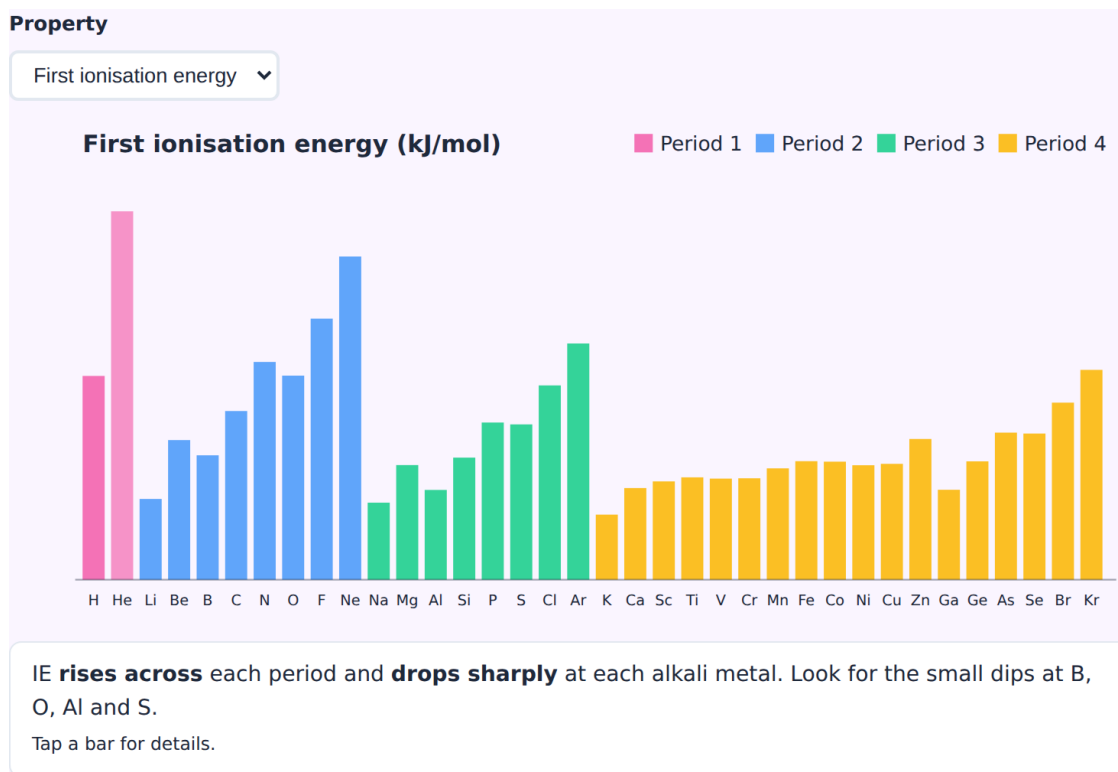
SOLUTION

Na: $11 - 10 = +1$ Cl: $17 - 10 = +7$

Cl's valence electrons feel a much larger pull, so Cl is smaller, has a higher IE and is more electronegative.

Interactive Simulation: Periodic Trend Explorer

Choose a property to plot against atomic number for elements 1–36. Tap a bar to read its value. Notice the repeating pattern in every period.



Companion-app task — Periodic Trend Explorer: compare Na and Mg, then compare two elements down a group. State which trend is observed and explain it using nuclear charge, shells, and shielding.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 7 Quiz

Q1. Which has the largest atomic radius?

- A. F
- B. Cl
- C. Br
- D. I

Q2. Which has the highest first ionisation energy?

- A. Na
- B. Mg
- C. Al
- D. Si

Q3. Na^+ is ___ than Na, and Cl^- is ___ than Cl.

- A. larger, larger
- B. smaller, smaller
- C. smaller, larger
- D. larger, smaller

End-of-Chapter Exercises

Level 1 — Recall and understanding

7.1 Which is the most electronegative: C, N, P or Si?

7.2 Which is larger: (a) S or S^{2-} (b) Fe^{2+} or Fe^{3+} ?

Level 2 — Application

7.3 Arrange K, Ca, Br, Kr in order of increasing first ionisation energy.

7.4 Estimate Z_{eff} (Z – core electrons) for the valence electrons of Si.

Answer: _____

7.5 Successive IEs of an element are 577, 1817, 2745 and 11 577 kJ/mol. Identify its group.

Level 3 — Challenge

7.6 Arrange S^{2-} , Cl^- , K^+ and Ca^{2+} in order of decreasing radius and explain.

7.7 Explain why the first ionisation energy of O is lower than that of N.

Chapter 8: Chemical Bonding I: Lewis Structures

Unit III: Periodic Table & Bonding

8.1 Drawing Lewis Structures

Procedure

1. Count total valence electrons
2. Place least electronegative atom in centre (H always terminal)
3. Connect with single bonds ($2e^-$ each)
4. Distribute remaining electrons as lone pairs
5. Form multiple bonds if central atom lacks octet

WORKED EXAMPLE: CO_2

SOLUTION

Valence e^- : $4 + 2(6) = 16$

C in centre: $O-C-O$ uses $4e^-$, remaining $12e^-$

Place $6e^-$ on each O $\rightarrow$ C only has $4e^- \rightarrow$ form double bonds

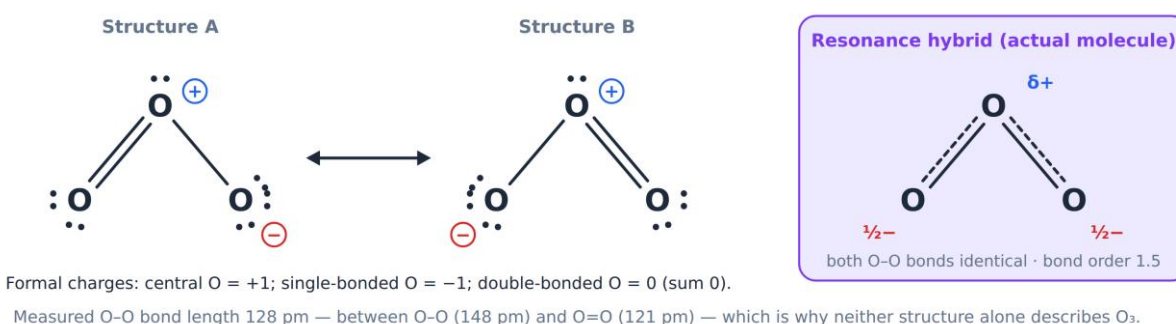
$:O = C = O:$ (all atoms have octet, FC = 0)

8.2 Formal Charge and Resonance

$$FC = V - N_{\text{nonbonding}} - \frac{1}{2}(N_{\text{bonding}})$$

Best structure: smallest formal charges, negative charge on most electronegative atom.

Resonance: Ozone (O_3)



The double-headed arrow does not mean the molecule flips between A and B: the real structure is a single blend of the two.

Additional Explanation: Lewis Structures Are Electron-Counting Models

A Lewis structure accounts for valence electrons as bonding pairs and lone pairs. A reliable method is to count total valence electrons, choose a skeleton, complete terminal-atom octets, place remaining electrons on the central atom, and form multiple bonds if needed. Formal charge is a bookkeeping tool for evaluating alternatives; the sum of formal charges must equal the overall charge of the species.

Worked Example: Lewis Structure of Carbon Dioxide

Draw the Lewis structure for CO_2 and check formal charges.

Total valence electrons = 4 for C + 2(6) for O = 16. Use O–C–O as the skeleton.

Complete the oxygen octets, then convert one lone pair from each oxygen into a bonding pair to form $\text{O}=\text{C}=\text{O}$.

Each O has two lone pairs and two bonds; C has four bonds. Formal charge on each atom is zero.

Check: all atoms have octets, the structure uses 16 valence electrons, and formal charges sum to zero.

Lewis structures do not by themselves show molecular shape or bond polarity.

More Worked Examples

Try each problem yourself before reading the solution.

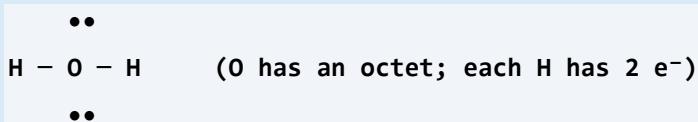
Worked Example 8.A: Lewis structure of water

Draw the Lewis structure of H_2O .

SOLUTION

Valence electrons: $2(1) + 6 = 8$

O in centre; two O–H bonds use 4 e^- ; remaining 4 e^- go on O as 2 lone pairs.

**Worked Example 8.B: Resonance in the carbonate ion**

Draw the resonance structures of CO_3^{2-} and find the C–O bond order.

SOLUTION

Valence electrons: $4 + 3(6) + 2 = 24$

C needs one C=O double bond to complete its octet. The double bond can be to any of the three O atoms → 3 resonance structures.

Formal charges: double-bonded O = 0; each single-bonded O = -1 (total -2)

Bond order = 4 bonds shared over 3 positions = $1\frac{1}{3}$; all three C–O bonds are identical.

Worked Example 8.C: Exceptions to the octet rule

How many electrons surround the central atom in BF_3 and in SF_6 ?

SOLUTION

BF_3 : $3 + 3(7) = 24\text{ e}^-$. Three B–F bonds; B has only 6 electrons (electron-deficient).

SF_6 : $6 + 6(7) = 48\text{ e}^-$. Six S–F bonds; S has 12 electrons (expanded octet — possible for Period 3 and below).

Worked Example 8.D: Choosing between structures with formal charge

For CO_2 , compare $\text{O}=\text{C}=\text{O}$ with $\text{O}\equiv\text{C}-\text{O}$.

SOLUTION

$\text{O}=\text{C}=\text{O}$: FC = 0 on every atom.

$\text{O}\equiv\text{C}-\text{O}$: FC($\text{O}\equiv$) = $6 - 2 - 3 = +1$; FC(C) = 0; FC($-\text{O}$) = $6 - 6 - 1 = -1$.

The structure with the smallest formal charges, $\text{O}=\text{C}=\text{O}$, is preferred.

Interactive Simulation: Valence Electron & Formal Charge Helper

Type a formula and charge to count the valence electrons for a Lewis structure. Then use the formal charge calculator to check each atom in your drawing.

1. Count valence electrons

Formula

Charge

Count

C: $1 \times 4 = 4$

O: $3 \times 6 = 18$

charge -2: add 2

Total = 24 valence electrons (12 pairs)

Likely central atom: **C** (least electronegative, H is never central)

2. Formal charge of one atom

Valence electrons (V)

Non-bonding electrons (dots)

Bonds (lines) to this atom

FC = V – (non-bonding e^-) – (bonds)

= $6 - 6 - 1 = -1$

Electrons around this atom: 8 (octet ✓)

Companion-app task — Valence Electron & Formal Charge Helper: enter CO_2 , compare electron counts and formal charges, and test an alternative structure to see why formal charge is useful.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 8 Quiz

Q1. Total valence electrons in SO_4^{2-} :

- A. 30
- B. 32
- C. 34
- D. 28

Q2. The formal charge on N in NH_4^+ is:

- A. 0
- B. +1
- C. -1
- D. +2

Q3. Which molecule has resonance structures?

- A. CH_4
- B. CO_2
- C. O_3
- D. H_2O

End-of-Chapter Exercises

Level 1 — Recall and understanding

8.1 Count the total valence electrons in: PCl_3 , NO_3^- , NH_4^+ , XeF_4 .

8.2 Which of these do not obey the octet rule: BF_3 , PCl_5 , NO , CH_4 ?

Level 2 — Application

8.3 Assign formal charges to each atom in carbon monoxide, $:\text{C}\equiv\text{O}:$.

8.4 Draw the resonance structures for NO_3^- and give the N–O bond order.

Level 3 — Challenge

8.5 In a Lewis structure of SO_4^{2-} with four S–O single bonds, what is the formal charge on S?

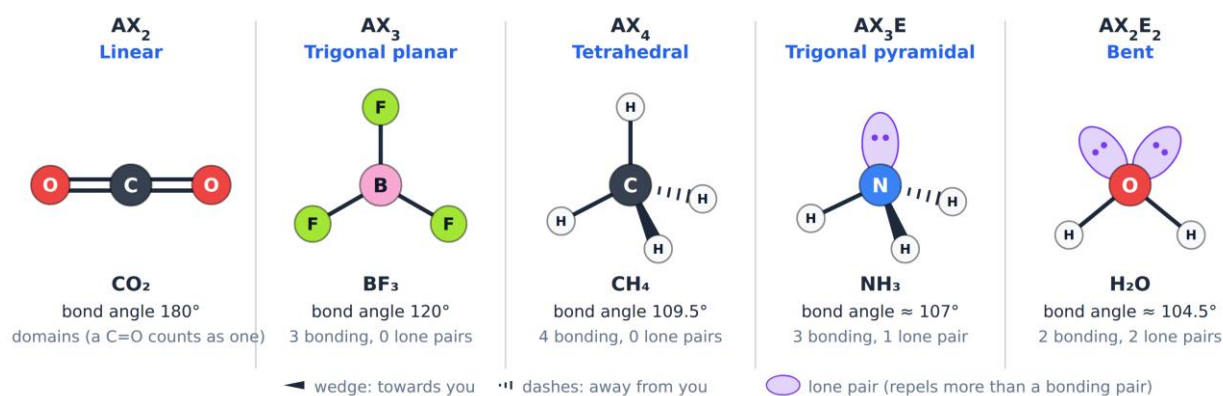
Answer: _____

Chapter 9: Molecular Geometry & Hybridisation

Unit III: Periodic Table & Bonding

9.1 VSEPR Theory

Domains	Electron Geometry	Molecular Shape	Angle	Example
2	Linear	Linear	180°	CO ₂
3	Trigonal planar	Trig planar / Bent	120°	BF ₃ / SO ₂
4	Tetrahedral	Tet / Pyr / Bent	109.5°	CH ₄ / NH ₃ / H ₂ O
5	Trig bipyramidal	Various	90°, 120°	PCl ₅ / SF ₄
6	Octahedral	Various	90°	SF ₆ / XeF ₄



Electron-domain geometry for NH₃ and H₂O is tetrahedral; the shape name describes only the positions of the atoms.

9.2 Hybridisation

Hybridisation	Geometry	Angle	Example
sp	Linear	180°	BeCl ₂ , CO ₂
sp ²	Trigonal planar	120°	BF ₃ , C ₂ H ₄
sp ³	Tetrahedral	109.5°	CH ₄ , NH ₃ , H ₂ O
sp ³ d	Trig bipyramidal	90°, 120°	PCl ₅
sp ³ d ²	Octahedral	90°	SF ₆

Polarity Rule: A molecule is polar if it has polar bonds AND asymmetric geometry (dipoles don't cancel).

Additional Explanation: Electron-Domain Geometry and Molecular Shape Are Not Identical

VSEPR theory treats regions of electron density around a central atom as repelling one another. A single, double, or triple bond counts as one electron domain; a lone pair also counts as one domain but often repels more strongly than a bonding pair. Molecular shape names describe the positions of atoms, not the positions of lone pairs.

Worked Example: Predict the Shape of Ammonia

Predict the electron-domain geometry and molecular shape of NH_3 .

Nitrogen has five valence electrons and forms three N–H single bonds, leaving one lone pair on N.

There are four electron domains, so the electron-domain geometry is tetrahedral.

Because one domain is a lone pair, the arrangement of atoms is trigonal pyramidal; the H–N–H angle is about 107° , smaller than the ideal tetrahedral angle.

Check: count electron domains first, then name the molecular shape by considering only atom positions.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 9.A: Seesaw shape: SF_4

Predict the shape and hybridisation of SF_4 .

SOLUTION

Valence electrons: $6 + 4(7) = 34$. Four S–F bonds and 12 lone pairs on F use 32; 2 e^- remain $\rightarrow$ 1 lone pair on S.

5 domains (AX_4E) $\rightarrow$ trigonal bipyramidal electron geometry; the lone pair sits in an equatorial position.

Shape: seesaw; hybridisation sp^3d .

Worked Example 9.B: Square planar: XeF_4

Predict the shape of XeF_4 .

SOLUTION

$8 + 4(7) = 36\text{ e}^- \rightarrow 4\text{ bonds} + 2\text{ lone pairs on Xe (AX}_4\text{E}_2\text{)}$.

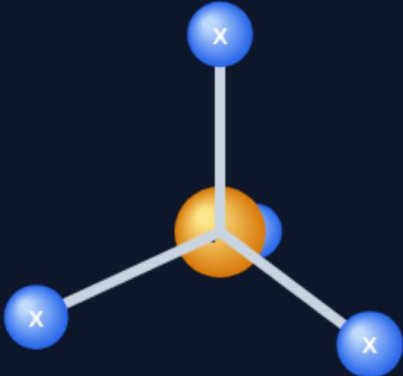
6 domains $\rightarrow$ octahedral; the two lone pairs go opposite each other to minimise repulsion.

Shape: square planar, 90° bonds, sp^3d^2 , nonpolar.

Worked Example 9.C: Predicting polarityWhich are polar: CO_2 , SO_2 , CH_2Cl_2 ?**SOLUTION** CO_2 : linear, two equal $\text{C}=\text{O}$ dipoles in opposite directions cancel $\rightarrow$ nonpolar. SO_2 : bent (lone pair on S), dipoles don't cancel $\rightarrow$ polar. CH_2Cl_2 : tetrahedral, but $\text{C}-\text{H}$ and $\text{C}-\text{Cl}$ dipoles are unequal $\rightarrow$ polar.**Interactive Simulation: VSEPR Shape Explorer**

Choose the number of bonding pairs and lone pairs on the central atom. The 3D model rotates — drag the slider to change the speed.

Bonding pairs (atoms) **Lone pairs on central atom** **Rotation speed:** 0.6



A = central atom X = bonded atom pink lobe = lone pair

AX_4 — 4 electron domains
 Electron geometry: **Tetrahedral** · Molecular shape: **Tetrahedral**
 Bond angle: **109.5°** · Hybridisation: **sp^3** · Examples: CH_4 , CCl_4
 Polarity: symmetric — **nonpolar** if all terminal atoms are the same

Companion-app task — VSEPR Shape Explorer: build NH_3 and compare it with CH_4 . Record both electron-domain geometry and molecular shape, and note the effect of a lone pair.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 9 Quiz

Q1. The molecular geometry of NH_3 is:

- A. Tetrahedral
- B. Trigonal pyramidal
- C. Bent
- D. Trigonal planar

Q2. The hybridisation of C in CH_4 is:

- A. sp
- B. sp^2
- C. sp^3
- D. sp^3d

Q3. Which molecule is nonpolar?

- A. H_2O
- B. NH_3
- C. CCl_4
- D. CHCl_3

End-of-Chapter Exercises

Level 1 — Recall and understanding

9.1 Give the hybridisation of the central atom in: BeCl_2 , SO_3 , NH_4^+ , PF_5 , SF_6 .

9.2 How many σ and π bonds are in (a) C_2H_4 (b) HCN ?

Level 2 — Application

9.3 Predict the shape and approximate bond angle of PCl_3 .

9.4 Polar or nonpolar? BF_3 , NF_3 , CH_2Cl_2 , XeF_4 .

Level 3 — Challenge

9.5 Predict the shape of ClF_3 .

9.6 Predict the shape of the triiodide ion, I_3^- .

Chapter 10: Gases and Kinetic Molecular Theory

Unit IV: States of Matter

10.1 The Ideal Gas Law

$$PV = nRT \quad \text{where } R = 0.08206 \text{ L} \cdot \text{atm}/(\text{mol} \cdot \text{K}) = 8.314 \text{ J}/(\text{mol} \cdot \text{K})$$

Individual Gas Laws

Law	Relationship	Constant
Boyle's	$P \propto 1/V$	T, n
Charles's	$V \propto T$	P, n
Avogadro's	$V \propto n$	T, P
Gay-Lussac's	$P \propto T$	V, n

Interactive tool in the online version: Ideal Gas Law Calculator — chemtextbook.neetapps.com.

10.2 Kinetic Molecular Theory

Five Postulates

1. Gas particles are point masses (negligible volume)
2. No intermolecular forces
3. Constant, random, straight-line motion
4. Collisions are perfectly elastic
5. Average KE depends only on T: $\overline{KE} = 3/2 k_B T$

$$u_{\text{rms}} = \sqrt{3RT/M} \quad \text{Graham's Law: } r_1/r_2 = \sqrt{M_2/M_1}$$

WORKED EXAMPLE

Calculate u_{rms} of N_2 at 25°C ($M = 0.028 \text{ kg/mol}$)

SOLUTION

$$u_{\text{rms}} = \sqrt{3 \times 8.314 \times 298.15 / 0.028} = \sqrt{265,590} = 515 \text{ m/s}$$

Additional Explanation: Use Absolute Temperature in Gas-Law Calculations

The ideal gas equation $PV = nRT$ connects pressure, volume, amount, and absolute temperature. Temperature must be in kelvin because gas-law proportionalities use a scale whose zero corresponds to the limiting zero of thermal motion. Keep units consistent with the chosen gas constant, and convert pressure or volume before substituting when necessary.

Worked Example: Amount of Gas in a Container

A gas occupies 2.00 L at 1.00 atm and 25.0°C . Estimate the amount in moles using $R = 0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1}$.

Convert temperature: $T = 25.0 + 273.15 = 298.15 \text{ K}$.

Rearrange $PV = nRT$: $n = PV/(RT)$.

$n = (1.00 \text{ atm})(2.00 \text{ L}) / [(0.08206 \text{ L atm mol}^{-1} \text{ K}^{-1})(298.15 \text{ K})] = 0.0817 \text{ mol}$.

Check: L and atm cancel with the corresponding units in R, leaving moles. The ideal-gas model is most reliable at relatively low pressure and high temperature.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 10.A: Combined gas law

A gas occupies 2.50 L at 25 °C and 101.3 kPa. What volume will it occupy at 75 °C and 150 kPa?

SOLUTION

$$P_1 V_1 / T_1 = P_2 V_2 / T_2$$

$T_1 = 298.15 \text{ K}$, $T_2 = 348.15 \text{ K}$ (always use kelvin)

$$V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1) = 2.50 \times (101.3/150) \times (348.15/298.15)$$

$$V_2 = 1.97 \text{ L}$$

Worked Example 10.B: Molar mass from gas density

A gas has a density of 1.25 g/L at STP (0 °C, 1 atm). Find its molar mass.

SOLUTION

From $PV = nRT$ and $n = m/M$: $M = dRT/P$

$$M = (1.25)(0.08206)(273.15)/1.00 = 28.0 \text{ g/mol} \text{ — it could be } \text{N}_2 \text{ or } \text{CO}.$$

Worked Example 10.C: Dalton's law of partial pressures

A mixture contains 0.20 mol O_2 , 0.50 mol N_2 and 0.30 mol Ar at a total pressure of 2.00 atm. Find each partial pressure.

SOLUTION

$$P_i = X_i \times P_{\text{total}}$$

Total moles = 1.00. $X(\text{O}_2) = 0.20$, $X(\text{N}_2) = 0.50$, $X(\text{Ar}) = 0.30$

$$P(\text{O}_2) = 0.40 \text{ atm}, P(\text{N}_2) = 1.00 \text{ atm}, P(\text{Ar}) = 0.60 \text{ atm}$$

Worked Example 10.D: Gas stoichiometry

What volume of CO_2 at STP forms when 10.0 g of CaCO_3 decomposes? $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$

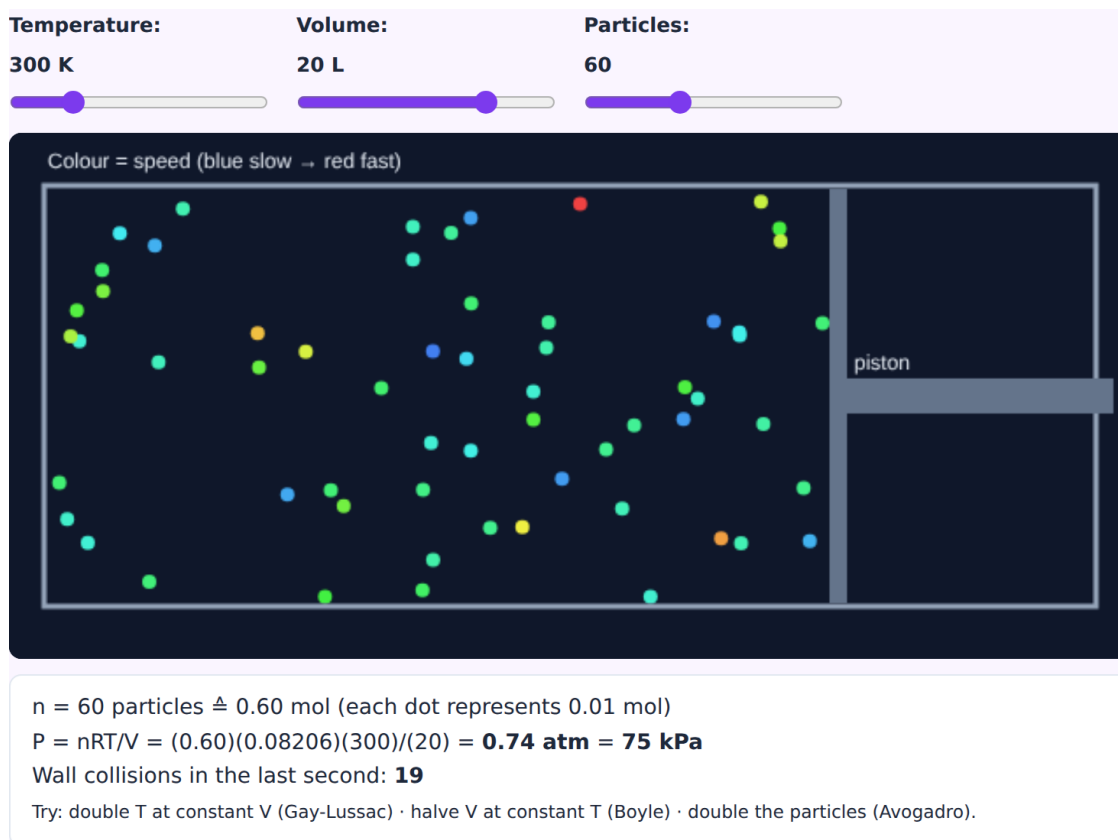
SOLUTION

$$n(\text{CaCO}_3) = 10.0/100.09 = 0.0999 \text{ mol} = n(\text{CO}_2)$$

$$V = n \times V_m = 0.0999 \times 22.4 = 2.24 \text{ L}$$

Interactive Simulation: Gas in a Piston

Change temperature, volume and the number of particles. Watch how the particles move and how the pressure responds. The collision counter shows how often particles hit the walls.



Companion-app task — Gas in a Piston: hold two variables fixed while changing a third. Predict the direction of the response before moving the piston, then connect the observation to $PV = nRT$.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 10 Quiz

Q1. At STP, 1 mole of ideal gas occupies:

- A. 11.2 L
- B. 22.4 L
- C. 44.8 L
- D. 1.0 L

Q2. If temperature doubles at constant volume, pressure:

- A. Halves
- B. Doubles
- C. Stays same
- D. Quadruples

Q3. H_2 effuses how many times faster than O_2 ?

- A. 2×
- B. 4×
- C. 8×
- D. 16×

End-of-Chapter Exercises

Level 1 — Recall and understanding

10.1 A gas at 100 kPa occupies 4.00 L. At constant temperature it is compressed to 250 kPa. What is the new volume?

Answer: _____ L

10.2 500 mL of gas at 20 °C is heated to 100 °C at constant pressure. Find the new volume.

Answer: _____ mL

Level 2 — Application

10.3 Calculate the pressure of 0.500 mol of gas in a 10.0 L container at 27 °C.

Answer: _____ atm

10.4 Air is 78% N₂ by moles. Find the partial pressure of N₂ when the air pressure is 101.3 kPa.

Answer: _____ kPa

10.5 6.54 g of Zn reacts with excess HCl: $\text{Zn} + 2\text{HCl} \rightarrow \text{ZnCl}_2 + \text{H}_2$. What volume of H₂ forms at STP?

Answer: _____ L

Level 3 — Challenge

10.6 0.856 g of a gas occupies 0.500 L at 25 °C and 1.00 atm. Find its molar mass.

Answer: _____ g/mol

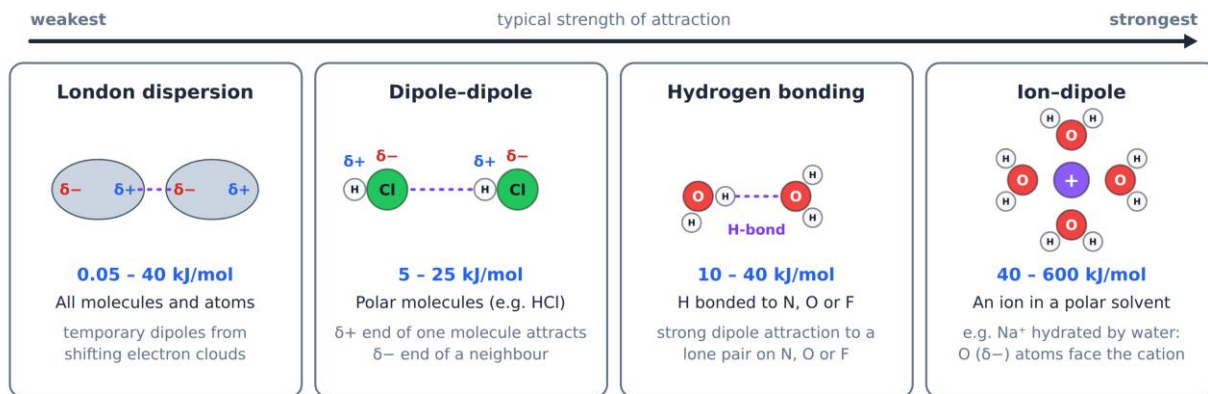
10.7 How many times faster does He effuse than CH₄?

Answer: _____ times

Chapter 11: Liquids, Solids, and Intermolecular Forces

Unit IV: States of Matter

11.1 Intermolecular Forces



The ranges overlap: London forces in large molecules (e.g. I_2 or long hydrocarbons) can outweigh dipole-dipole forces in small polar molecules.

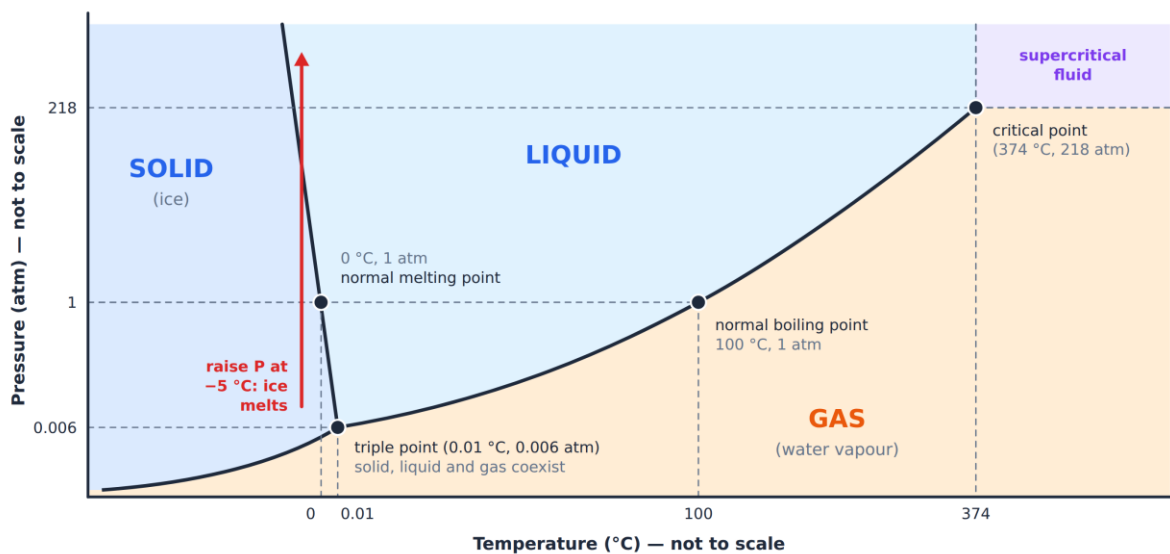
Dotted purple lines show the attraction between particles; $\delta+$ and $\delta-$ mark partial charges.

Key: H-bonding requires H directly bonded to N, O, or F. This is why H_2O (bp 100°C) $\gg$ H_2S (bp -60°C).

11.2 Types of Crystalline Solids

Type	Particles	Forces	Example	Properties
Ionic	Ions	Ion-ion	NaCl	High mp, brittle
Covalent network	Atoms	Covalent	Diamond	Very high mp, hard
Molecular	Molecules	IMF	Ice, I_2	Low mp, soft
Metallic	Cations + e^-	Metallic	Fe, Cu	Conduct, malleable

11.3 Phase Diagrams



Water's solid–liquid line slopes slightly to the left (negative slope), so raising the pressure lowers the melting point of ice — unusual, because ice is less dense than liquid water.

Additional Explanation: Temperature Changes and Phase Changes Differ

When a substance warms within a single phase, added energy increases average particle kinetic energy and temperature rises. During a phase change at constant pressure, energy is used to change intermolecular arrangements, so the temperature can remain nearly constant while the phase changes. The energy per amount for this process is represented by a latent heat or enthalpy of phase change.

Worked Example: Energy to Vaporise Water

Estimate the energy needed to vaporise 18.0 g of water at its boiling point. Use $\Delta H_{\text{vap}} = 40.7 \text{ kJ mol}^{-1}$.

Moles of water = $18.0 \text{ g} \div 18.02 \text{ g mol}^{-1} = 0.999 \text{ mol}$ (approximately 1.00 mol).

At the boiling point, $q = n\Delta H_{\text{vap}} = (0.999 \text{ mol})(40.7 \text{ kJ mol}^{-1})$.

$q \approx 40.7 \text{ kJ}$. This energy is needed for vaporisation without a temperature rise during the phase change.

Check: vaporisation is endothermic for the water. If the starting water were below its boiling point, additional energy would be required to warm it first.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 11.A: Energy along a heating curve

How much energy is needed to turn 50.0 g of ice at $-10.0\text{ }^{\circ}\text{C}$ into steam at $110.0\text{ }^{\circ}\text{C}$? ($c_{\text{ice}} = 2.09$, $c_{\text{water}} = 4.18$, $c_{\text{steam}} = 2.01\text{ J/(g}\cdot^{\circ}\text{C)}$; $\Delta H_{\text{fus}} = 6.01\text{ kJ/mol}$; $\Delta H_{\text{vap}} = 40.7\text{ kJ/mol}$)

SOLUTION

$$n(\text{H}_2\text{O}) = 50.0/18.02 = 2.775\text{ mol}$$

Step	Calculation	q (kJ)
1. Warm ice $-10 \rightarrow 0\text{ }^{\circ}\text{C}$	$50.0 \times 2.09 \times 10.0$	1.05
2. Melt at $0\text{ }^{\circ}\text{C}$	2.775×6.01	16.7
3. Warm water $0 \rightarrow 100\text{ }^{\circ}\text{C}$	$50.0 \times 4.18 \times 100$	20.9
4. Boil at $100\text{ }^{\circ}\text{C}$	2.775×40.7	112.9
5. Warm steam $100 \rightarrow 110\text{ }^{\circ}\text{C}$	$50.0 \times 2.01 \times 10.0$	1.01

Total = 153 kJ. Notice that boiling takes far more energy than any other step.

Worked Example 11.B: Identifying intermolecular forces

Identify the IMFs in CH_4 , HCl , CH_3OH and NaCl(aq) , and rank the first three by boiling point.

SOLUTION

CH_4 : London only. HCl : London + dipole–dipole. CH_3OH : London + dipole–dipole + hydrogen bonding (O–H). NaCl(aq) : ion–dipole.

Boiling point: $\text{CH}_4 (-162\text{ }^{\circ}\text{C}) < \text{HCl } (-85\text{ }^{\circ}\text{C}) < \text{CH}_3\text{OH } (65\text{ }^{\circ}\text{C})$

Worked Example 11.C: Classifying solids

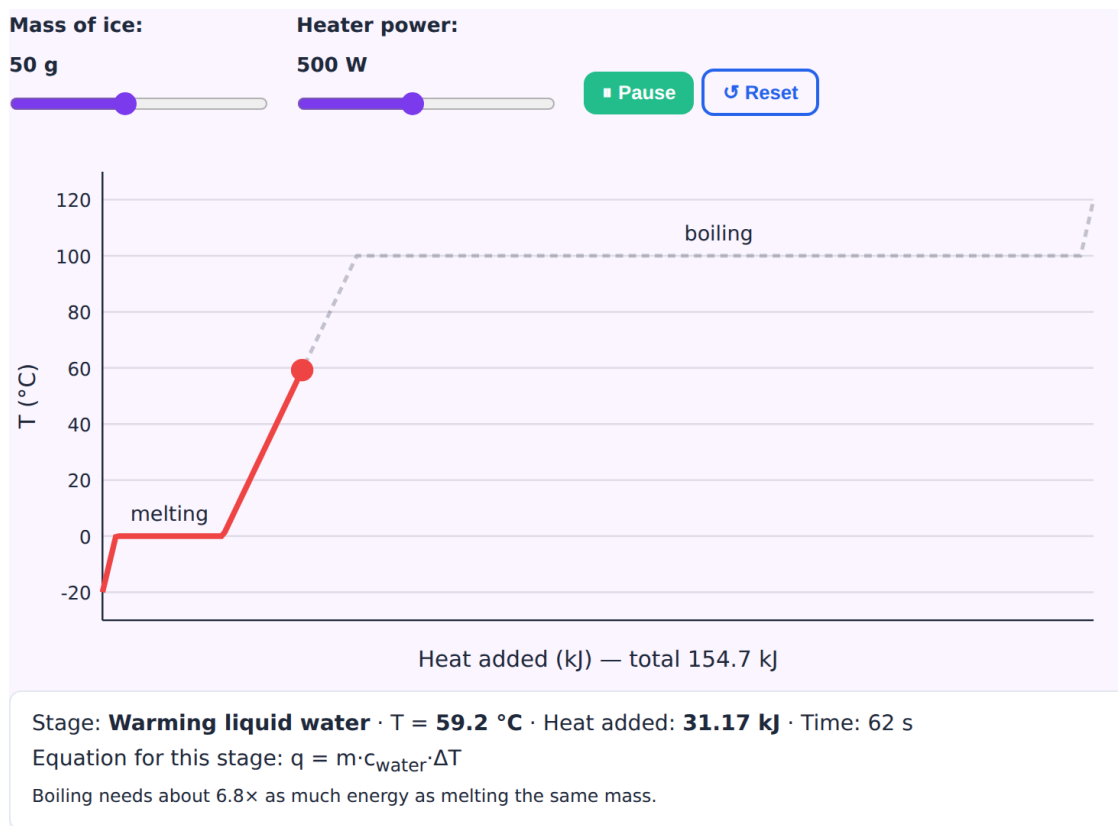
Classify SiO_2 , dry ice (CO_2), Mg and KBr .

SOLUTION

SiO_2 : covalent network (very hard, high mp). $\text{CO}_2(\text{s})$: molecular (sublimes at $-78\text{ }^{\circ}\text{C}$). Mg : metallic (conducts, malleable). KBr : ionic (high mp, brittle, conducts when molten).

Interactive Simulation: Heating Curve of Water

Supply heat to a sample of ice and watch the temperature. The graph shows temperature against heat added; flat sections are phase changes.



Companion-app task — Heating Curve of Water: identify sloped segments and plateaus. Explain which segments show temperature change and which show a phase change.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 11 Quiz

Q1. Which IMF is present in ALL molecules?

- A. H-bonding
- B. Dipole-dipole
- C. London dispersion
- D. Ion-dipole

Q2. Why does H_2O have a higher boiling point than H_2S ?

- A. H_2O is larger
- B. H-bonding in H_2O
- C. H_2S is nonpolar
- D. Ionic bonding in H_2O

Q3. The triple point is where:

- A. Only solid exists
- B. Solid and liquid coexist
- C. All three phases coexist
- D. Critical point occurs

End-of-Chapter Exercises

Level 1 — Recall and understanding

11.1 Identify the strongest IMF in: (a) Br_2 (b) NH_3 (c) H_2S (d) KCl dissolved in water.

11.2 Rank by increasing boiling point: C_4H_{10} , CH_4 , C_3H_8 , C_2H_6 .

Level 2 — Application

11.3 How much heat is needed to melt 25.0 g of ice at 0°C ? ($\Delta H_{\text{fus}} = 6.01 \text{ kJ/mol}$)

Answer: _____ kJ

11.4 How much heat is needed to boil 10.0 g of water at 100°C ? ($\Delta H_{\text{vap}} = 40.7 \text{ kJ/mol}$)

Answer: _____ kJ

11.5 Classify and give one property of: diamond, Cu , NaCl , I_2 .

Level 3 — Challenge

11.6 Ice at -5°C is put under a very large pressure. Use the phase diagram of water to predict what happens.

11.7 Explain why ice floats on water.

Chapter 12: Solutions and Their Properties

Unit IV: States of Matter

12.1 Concentration Units

Unit	Formula
Molarity (M)	mol solute / L solution
Molality (m)	mol solute / kg solvent
Mass percent	(g solute / g solution) × 100%
Mole fraction (X)	n_solute / n_total

12.2 Colligative Properties

$$\Delta T_b = i \cdot K_b \cdot m \quad (\text{boiling point elevation})$$

$$\Delta T_f = i \cdot K_f \cdot m \quad (\text{freezing point depression})$$

$$\Pi = iMRT \quad (\text{osmotic pressure})$$

$$P_{\text{solution}} = X_{\text{solvent}} \times P^{\circ}_{\text{solvent}} \quad (\text{Raoult's Law})$$

Solute	i (van't Hoff, ideal)
NaCl	2
CaCl ₂	3
AlCl ₃	4
Glucose	1

WORKED EXAMPLE

15.0 g ethylene glycol (M=62.07) in 250 g water. Find ΔT_f and ΔT_b .

SOLUTION

$$m = (15.0/62.07)/0.250 = 0.967 \text{ m}; i = 1 \text{ (nonelectrolyte)}$$

$$\Delta T_f = (1)(1.86)(0.967) = 1.80^{\circ}\text{C} \rightarrow T_f = -1.80^{\circ}\text{C}$$

$$\Delta T_b = (1)(0.512)(0.967) = 0.495^{\circ}\text{C} \rightarrow T_b = 100.50^{\circ}\text{C}$$

Additional Explanation: Concentration Depends on What Is in the Denominator

Molarity is moles of solute per litre of final solution; molality is moles of solute per kilogram of solvent. These are not interchangeable. Molarity can vary with temperature because solution volume changes, whereas molality is based on mass. For colligative-property calculations, identify the appropriate concentration unit and the number of dissolved particles represented by the van 't Hoff factor.

Worked Example: Calculate Molarity

A solution contains 0.500 mol of solute in a final solution volume of 250.0 mL. Find its molarity.

Convert the final solution volume to litres: 250.0 mL = 0.2500 L.

Molarity = moles solute ÷ litres solution = 0.500 mol ÷ 0.2500 L.

$M = 2.00 \text{ mol L}^{-1}$, or 2.00 M.

Check: use final solution volume, not the initial volume of solvent. State the concentration unit with the numerical answer.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 12.A: From mass percent to molarity

Concentrated hydrochloric acid is 37.0% HCl by mass and has a density of 1.19 g/mL. Find its molarity.

SOLUTION

Take 1.00 L = 1000 mL → mass = 1000 × 1.19 = 1190 g of solution

Mass of HCl = 0.370 × 1190 = 440 g → $n = 440/36.46 = 12.1 \text{ mol}$

$c = 12.1 \text{ mol/L}$

Worked Example 12.B: Osmotic pressure

Find the osmotic pressure of 0.100 mol/L NaCl at 25 °C.

SOLUTION

$\Pi = iMRT = (2)(0.100)(0.08206)(298.15) = 4.89 \text{ atm}$

Even a dilute salt solution gives a large osmotic pressure — about 50 m of water.

Worked Example 12.C: Raoult's law

1.00 mol glucose is dissolved in 9.00 mol water at 25 °C ($P^\circ_{\text{water}} = 3.17 \text{ kPa}$). Find the vapour pressure of the solution.

SOLUTION

$X_{\text{water}} = 9.00/10.00 = 0.900$

$P = 0.900 \times 3.17 = 2.85 \text{ kPa}$

Worked Example 12.D: Dilution

What volume of 2.00 mol/L stock solution is needed to make 250 mL of 0.500 mol/L?

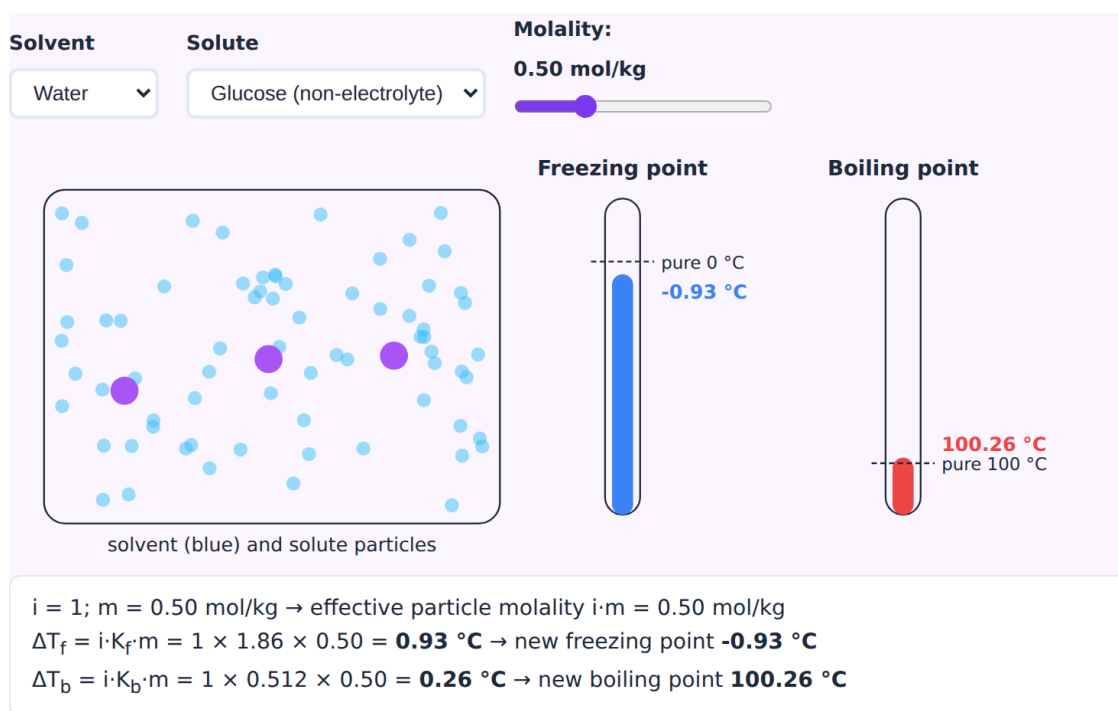
SOLUTION

$c_1V_1 = c_2V_2 \rightarrow V_1 = (0.500)(250)/2.00 = 62.5 \text{ mL}$

Pipette 62.5 mL of stock into a 250 mL volumetric flask and make up to the mark.

Interactive Simulation: Colligative Properties Explorer

Choose a solvent and solute and set the molality. Notice that the number of particles (i) matters, not what the particles are.



Companion-app task — Colligative Properties Explorer: compare a nonelectrolyte with an electrolyte at the same nominal concentration. Explain how the particle count affects boiling-point elevation or freezing-point depression.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 12 Quiz

Q1. Which solution has the lowest freezing point?

- A. 0.1m NaCl
- B. 0.1m CaCl_2
- C. 0.1m glucose
- D. 0.1m AlCl_3

Q2. Osmosis is the flow of:

- A. Solute through membrane
- B. Solvent from dilute to concentrated
- C. Solvent from concentrated to dilute
- D. Ions through solution

End-of-Chapter Exercises

Level 1 — Recall and understanding

12.1 10.0 g of NaOH is dissolved to make 250 mL of solution. Find its concentration.

Answer: _____ mol/L

12.2 What volume of 6.0 mol/L HCl is needed to make 500 mL of 0.30 mol/L HCl?

Answer: _____ mL

Level 2 — Application

12.3 20.0 g of urea ($M = 60.06 \text{ g/mol}$) is dissolved in 500 g of water. Find the molality.

Answer: _____ mol/kg

12.4 Find the mole fraction of ethanol in a mixture of 46.07 g ethanol and 72.06 g water.

Answer: _____

12.5 Find the freezing point of a 0.250 mol/kg CaCl_2 solution in water (ideal $i = 3$, $K_f = 1.86 \text{ }^\circ\text{C}\cdot\text{kg/mol}$).

Answer: _____ $^\circ\text{C}$

Level 3 — Challenge

12.6 2.50 g of an unknown nonelectrolyte in 50.0 g of benzene lowers the freezing point by $1.28 \text{ }^\circ\text{C}$ ($K_f = 5.12 \text{ }^\circ\text{C}\cdot\text{kg/mol}$). Find the molar mass.

Answer: _____ g/mol

12.7 Calculate the osmotic pressure of 0.0500 mol/L glucose at body temperature ($37 \text{ }^\circ\text{C}$).

Answer: _____ atm

Chapter 13: Chemical Kinetics

Unit V: Dynamics & Equilibrium

13.1 Rate Laws and Order

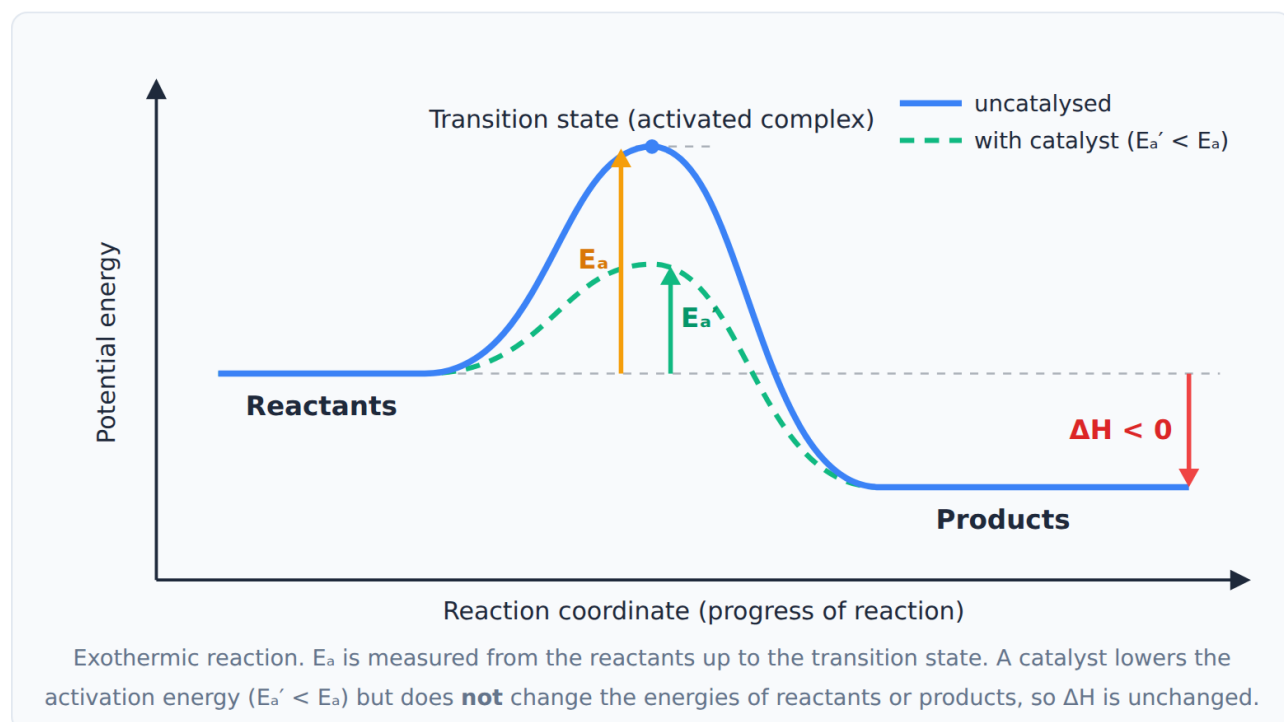
Rate = $k[A]^m[B]^n$ (m, n determined experimentally, NOT from coefficients)

Order	Rate Law	Integrated Form	Half-life
0	Rate = k	$[A] = [A]_0 - kt$	$[A]_0/(2k)$
1	Rate = $k[A]$	$\ln[A] = \ln[A]_0 - kt$	$0.693/k$
2	Rate = $k[A]^2$	$1/[A] = 1/[A]_0 + kt$	$1/(k[A]_0)$

13.2 Arrhenius Equation

$$k = Ae^{(-E_a/RT)}$$

$$\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$$



Reading the diagram: E_a (forward) = $E(\text{transition state}) - E(\text{reactants})$. E_a (reverse) = $E(\text{transition state}) - E(\text{products}) = E_a(\text{forward}) - \Delta H$.

Additional Explanation: Determine Reaction Order from Evidence

The rate law is determined experimentally; it should not be inferred from the balanced overall equation unless a reaction is explicitly elementary. For a first-order process, concentration decays exponentially and the half-life is independent of initial concentration. Integrated rate laws allow concentration, time, and the rate constant to be connected and can help identify reaction order from data.

Worked Example: First-Order Half-Life

A first-order reaction has $k = 0.0462 \text{ min}^{-1}$. Calculate its half-life.

For a first-order reaction, $t_{1/2} = \ln 2/k = 0.693/k$.

Substitute: $t_{1/2} = 0.693 \div 0.0462 \text{ min}^{-1}$.

$t_{1/2} = 15.0 \text{ min}$ (to three significant figures).

Check: the units are minutes because the inverse units of k are min. For a first-order process, each successive half-life halves the amount remaining.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 13.A: Method of initial rates

For $2\text{NO} + \text{O}_2 \rightarrow 2\text{NO}_2$:

Exp	[NO] (M)	[O ₂] (M)	Rate (M/s)
1	0.010	0.010	2.5×10^{-5}
2	0.020	0.010	1.0×10^{-4}
3	0.010	0.020	5.0×10^{-5}

Find the rate law and k .

SOLUTION

Exp 1 $\rightarrow$ 2: [NO] doubles, rate $\times 4 = 2^2 \rightarrow$ second order in NO

Exp 1 $\rightarrow$ 3: [O₂] doubles, rate $\times 2 = 2^1 \rightarrow$ first order in O₂

Rate = $k[\text{NO}]^2[\text{O}_2]$; $k = 2.5 \times 10^{-5} / ((0.010)^2(0.010)) = 25 \text{ M}^{-2}\text{s}^{-1}$

Worked Example 13.B: First-order integrated rate law

N_2O_5 decomposes with $k = 5.0 \times 10^{-4} \text{ s}^{-1}$. If $[\text{N}_2\text{O}_5]_0 = 0.100 \text{ M}$, find $[\text{N}_2\text{O}_5]$ after 30.0 min and the half-life.

SOLUTION

$t = 1800 \text{ s}$; $\ln[A] = \ln[A]_0 - kt \rightarrow [A] = 0.100 \times e^{-(5.0 \times 10^{-4})(1800)} = 0.100 \times e^{-0.90}$

$[A] = 0.0407 \text{ M}$

$t_{1/2} = 0.693/k = 0.693/5.0 \times 10^{-4} = 1.39 \times 10^3 \text{ s (23 min)}$

Worked Example 13.C: Activation energy from two temperatures

A rate constant doubles when the temperature rises from 25 °C to 35 °C. Find E_a .

SOLUTION

$$\ln(k_2/k_1) = (E_a/R)(1/T_1 - 1/T_2)$$

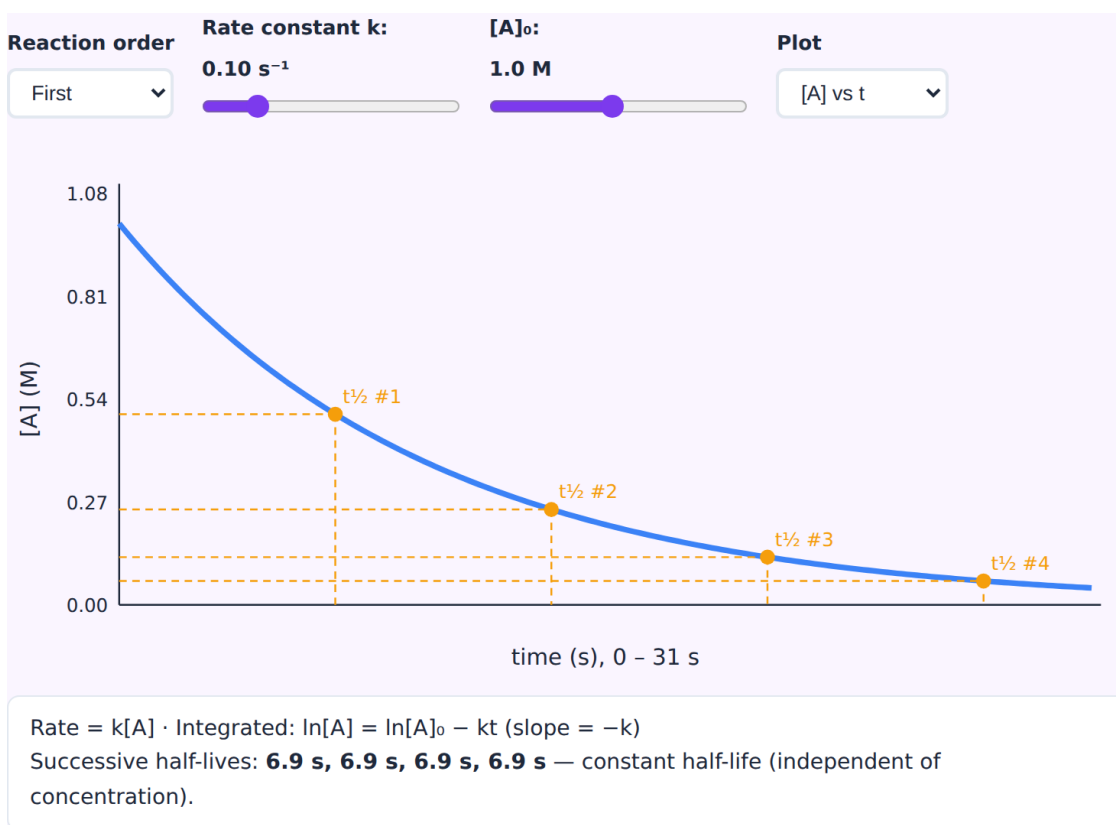
$$\ln 2 = (E_a/8.314)(1/298.15 - 1/308.15) = (E_a/8.314)(1.088 \times 10^{-4})$$

$$E_a = (0.693 \times 8.314)/(1.088 \times 10^{-4}) = 53 \text{ kJ/mol}$$

This is the origin of the rule of thumb "rates double for every 10 °C rise".

Interactive Simulation: Concentration–Time Plotter

Choose a reaction order and adjust k and $[A]_0$. Dotted lines mark successive half-lives. Switch to the linear plot to see which graph gives a straight line for each order.



Companion-app task — Concentration–Time Plotter: compare the concentration–time curve with a plot of $\ln[A]$ against time. Identify which representation is linear for a first-order reaction.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 13 Quiz

Q1. A first-order reaction has $t_{1/2} = 10 \text{ min}$. After 30 min, fraction remaining:

- A. $1/2$
- B. $1/4$
- C. $1/8$
- D. $1/3$

Q2. A catalyst works by:

- A. Increasing ΔH
- B. Lowering E_a
- C. Increasing temperature
- D. Shifting equilibrium

End-of-Chapter Exercises

Level 1 — Recall and understanding

13.1 A first-order reaction has $k = 0.0231 \text{ min}^{-1}$. What is its half-life?

Answer: _____ min

13.2 A zero-order reaction has $k = 0.010 \text{ M/s}$ and $[A]_0 = 0.50 \text{ M}$. Find the half-life.

Answer: _____ s

Level 2 — Application

13.3 A first-order reaction has $[A]_0 = 0.80 \text{ M}$ and $k = 0.10 \text{ s}^{-1}$. Find $[A]$ after 15 s.

Answer: _____ M

13.4 A second-order reaction has $k = 0.50 \text{ M}^{-1}\text{s}^{-1}$ and $[A]_0 = 0.20 \text{ M}$. How long until $[A] = 0.10 \text{ M}$?

Answer: _____ s

13.5 For $A + B \rightarrow C$: Exp 1 $[A]=0.10$, $[B]=0.10$, rate 4.0×10^{-3} ; Exp 2 $[A]=0.20$, $[B]=0.10$, rate 8.0×10^{-3} ; Exp 3 $[A]=0.10$, $[B]=0.20$, rate 1.6×10^{-2} . Find the rate law and k (in $\text{M}^{-2}\text{s}^{-1}$).

Answer: _____ $\text{M}^{-2}\text{s}^{-1}$

Level 3 — Challenge

13.6 $k = 1.0 \times 10^{-3} \text{ s}^{-1}$ at 300 K and $4.0 \times 10^{-3} \text{ s}^{-1}$ at 320 K. Calculate E_a in kJ/mol.

Answer: _____ kJ/mol

13.7 Sketch and describe how a catalyst changes the energy profile of a reaction.

Chapter 14: Chemical Equilibrium

Unit V: Dynamics & Equilibrium

14.1 Equilibrium Constant

$$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

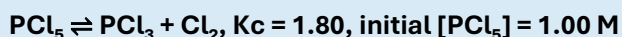
$$K_p = K_c(RT)^{\Delta n} \quad \text{where } \Delta n = \text{mol gas products} - \text{mol gas reactants}$$

Rules: Pure solids/liquids omitted. $K \gg 1$: products favoured. $K \ll 1$: reactants favoured.

14.2 Le Châtelier's Principle

Disturbance	Shift
Add reactant	→ toward products
Remove product	→ toward products
Increase P (decrease V)	Toward fewer gas moles
Increase T (exothermic rxn)	← toward reactants
Add catalyst	No shift (faster equilibrium)

ICE TABLE EXAMPLE



SOLUTION

$$x^2/(1.00-x) = 1.80 \rightarrow x^2 + 1.80x - 1.80 = 0 \rightarrow x = 0.716$$

$$[\text{PCl}_5] = 0.284 \text{ M}, [\text{PCl}_3] = [\text{Cl}_2] = 0.716 \text{ M}$$

Additional Explanation: Compare Q with K to Predict Direction

The equilibrium constant K describes the equilibrium ratio at a specified temperature. The reaction quotient Q has the same mathematical form but uses the current concentrations or partial pressures. Comparing Q with K predicts the direction of net change: the system shifts to reduce the difference. A catalyst changes how quickly equilibrium is reached, not the value of K or the equilibrium composition.

Worked Example: Predict the Direction of Shift

For a reaction at a fixed temperature, $Q = 0.20$ and $K = 4.0$. Which direction will the reaction proceed to approach equilibrium?

Compare the values: $Q < K$ ($0.20 < 4.0$).

The current product-to-reactant ratio is too small relative to equilibrium.

The net reaction proceeds in the forward direction, forming more products until Q reaches K .

Check: $Q = K$ means the system is at equilibrium; $Q > K$ predicts net change toward reactants; $Q < K$ predicts net change toward products.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 14.A: Heterogeneous equilibrium

Write K_c for $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$.

SOLUTION

Pure solids are left out: $K_c = [\text{CO}_2]$

So at a given temperature the CO_2 concentration at equilibrium is fixed, no matter how much CaCO_3 or CaO is present.

Worked Example 14.B: Calculating K from equilibrium data

1.00 mol H_2 and 1.00 mol I_2 are placed in a 1.00 L flask. At equilibrium there are 1.56 mol HI . Find K_c for $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$.

SOLUTION

	H_2	I_2	HI
Initial	1.00	1.00	0
Change	-0.78	-0.78	+1.56
Equilibrium	0.22	0.22	1.56

$$K_c = (1.56)^2 / ((0.22)(0.22)) = 50$$

Worked Example 14.C: Using Q to predict the direction

For $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, $K_c = 0.50$. A mixture has $[\text{N}_2] = [\text{H}_2] = [\text{NH}_3] = 1.0 \text{ M}$. Which way will it shift?

SOLUTION

$$Q = (1.0)^2 / ((1.0)(1.0)^3) = 1.0$$

$Q > K \rightarrow$ too much product $\rightarrow$ shifts left (towards N_2 and H_2).

Worked Example 14.D: Converting K_c to K_p

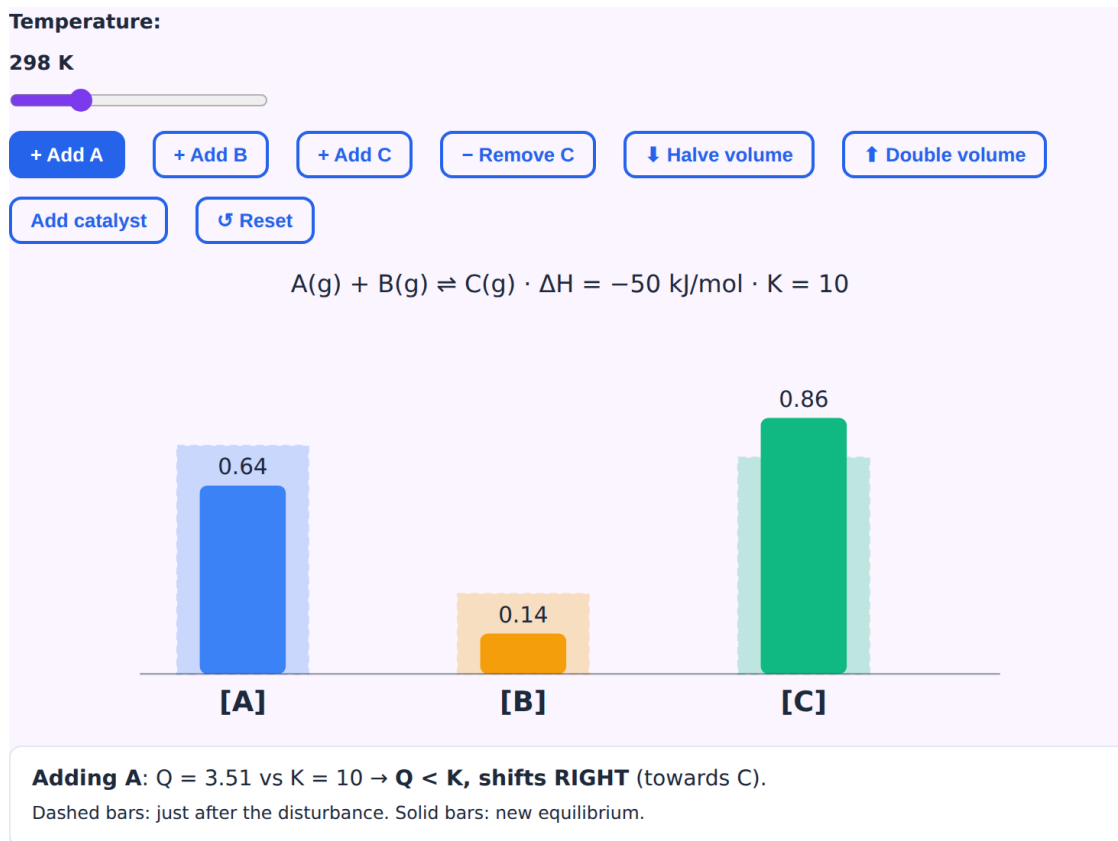
For $\text{N}_2\text{O}_4(\text{g}) \rightleftharpoons 2\text{NO}_2(\text{g})$, $K_c = 4.6 \times 10^{-3}$ at 298 K. Find K_p .

SOLUTION

$$\Delta n = 2 - 1 = 1; K_p = K_c(RT)^{\Delta n} = 4.6 \times 10^{-3} \times (0.08206 \times 298)^1 = 0.11$$

Interactive Simulation: Le Châtelier's Principle Lab

The system $A(g) + B(g) \rightleftharpoons C(g)$ is exothermic in the forward direction. Disturb it, predict the shift, then compare with the faded bars (before) and solid bars (new equilibrium).



Companion-app task — Le Châtelier's Principle Lab: change concentration, pressure, or temperature one at a time. Separate a change in equilibrium position from a change in the equilibrium constant.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 14 Quiz

Q1. If $Q > K$, the reaction will:

- A. Shift right
- B. Shift left
- C. Stay at equilibrium
- D. Speed up

Q2. Adding a catalyst to a system at equilibrium:

- A. Shifts right
- B. Shifts left
- C. Increases K
- D. No shift, faster equilibrium

End-of-Chapter Exercises

Level 1 — Recall and understanding

14.1 Write the K_c expression for $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$.

14.2 If K_c for $A \rightleftharpoons 2B$ is 4.0, what is K_c for (a) $2B \rightleftharpoons A$ (b) $2A \rightleftharpoons 4B$?

Level 2 — Application

14.3 For $2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3$, equilibrium concentrations are $[\text{SO}_2] = 0.20 \text{ M}$, $[\text{O}_2] = 0.10 \text{ M}$, $[\text{SO}_3] = 0.60 \text{ M}$. Calculate K_c .

Answer: _____

14.4 $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$, $\Delta H = -92 \text{ kJ}$. Predict the shift when you: (a) add H_2 (b) raise T (c) increase P by compressing (d) add Ne at constant volume.

Level 3 — Challenge

14.5 $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$ has $K_c = 49.0$. If 0.500 M H_2 and 0.500 M I_2 are mixed, find $[\text{HI}]$ at equilibrium.

Answer: _____ M

14.6 For $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ at 500 K , $K_c = 0.060$. Calculate K_p .

Answer: _____

Chapter 15: Acids and Bases

Unit V: Dynamics & Equilibrium

15.1 pH and Ka

$$\text{pH} = -\log[\text{H}^+] \quad \text{pOH} = -\log[\text{OH}^-] \quad \text{pH} + \text{pOH} = 14$$

$$K_a \times K_b = K_w = 1.0 \times 10^{-14} \text{ (at } 25^\circ\text{C)}$$

Interactive tool in the online version: pH Calculator — chemtextbook.neetapps.com.

15.2 Buffers (Henderson-Hasselbalch)

$$\text{pH} = \text{pK}_a + \log\left(\frac{[\text{A}^-]}{[\text{HA}]}\right)$$

WORKED EXAMPLE

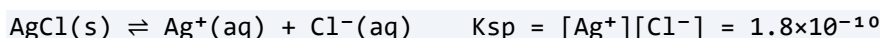
Buffer: 0.20 M CH_3COOH + 0.30 M CH_3COONa ($K_a = 1.8 \times 10^{-5}$)

SOLUTION

$$\text{pK}_a = -\log(1.8 \times 10^{-5}) = 4.74$$

$$\text{pH} = 4.74 + \log(0.30/0.20) = 4.74 + 0.18 = 4.92$$

15.3 K_{sp} and Solubility



Common Ion Effect: Adding a common ion (e.g., Cl^- from NaCl) dramatically reduces solubility.

Additional Explanation: Match the Acid–Base Model to the Question

The Arrhenius, Brønsted–Lowry, and Lewis models describe acids and bases at different levels. In aqueous calculations, $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$ is a useful concentration-based definition. Strong acids ionise nearly completely in dilute solution, while weak acids establish an equilibrium that may require an ICE table or quadratic solution. Always check whether an approximation is justified.

Worked Example: pH of a Strong Acid

Estimate the pH of $1.0 \times 10^{-3} \text{ mol L}^{-1}$ HCl at 25°C , assuming complete ionisation and ideal dilute-solution behaviour.

HCl is treated as a strong monoprotic acid, so $[\text{H}_3\text{O}^+] \approx 1.0 \times 10^{-3} \text{ M}$.

$$\text{pH} = -\log_{10}(1.0 \times 10^{-3}).$$

$$\text{pH} = 3.00.$$

Check: a tenfold decrease in $[\text{H}_3\text{O}^+]$ raises pH by one unit. At very high concentrations, activity effects make the simple concentration treatment less accurate.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 15.A: pH of a weak acid

Find the pH and percent ionisation of 0.10 M CH_3COOH ($K_a = 1.8 \times 10^{-5}$).

SOLUTION

$$K_a = x^2 / (0.10 - x). \text{ Assuming } x \ll 0.10: x = \sqrt{(1.8 \times 10^{-5})} = 1.34 \times 10^{-3} \text{ M}$$

Check: $1.34 \times 10^{-3} / 0.10 = 1.3\% < 5\% \rightarrow$ assumption valid

$$\text{pH} = -\log(1.34 \times 10^{-3}) = 2.87; \text{ percent ionisation} = 1.3\%$$

Worked Example 15.B: pH of a weak base

Find the pH of 0.20 M NH_3 ($K_b = 1.8 \times 10^{-5}$).

SOLUTION

$$[\text{OH}^-] = \sqrt{(K_b \times c)} = \sqrt{(1.8 \times 10^{-5} \times 0.20)} = 1.90 \times 10^{-3} \text{ M}$$

$$\text{pOH} = 2.72 \rightarrow \text{pH} = 14.00 - 2.72 = 11.28$$

Worked Example 15.C: Weak acid – strong base titration

25.0 mL of 0.100 M CH_3COOH is titrated with 0.100 M NaOH . Find the pH (a) at half-equivalence (b) at equivalence.

SOLUTION

(a) At 12.5 mL, $[\text{HA}] = [\text{A}^-] \rightarrow \text{pH} = \text{p}K_a = 4.74$

(b) At 25.0 mL all $\text{HA} \rightarrow \text{A}^-$: $[\text{A}^-] = 0.00250 \text{ mol} / 0.0500 \text{ L} = 0.0500 \text{ M}$

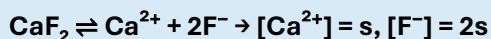
$$K_b = K_w / K_a = 5.56 \times 10^{-10}; [\text{OH}^-] = \sqrt{(5.56 \times 10^{-10} \times 0.0500)} = 5.27 \times 10^{-6} \text{ M}$$

$\text{pOH} = 5.28 \rightarrow \text{pH} = 8.72$ (basic, so phenolphthalein is a suitable indicator)

Worked Example 15.D: Solubility from K_{sp}

Find the molar solubility of CaF_2 ($K_{sp} = 3.9 \times 10^{-11}$).

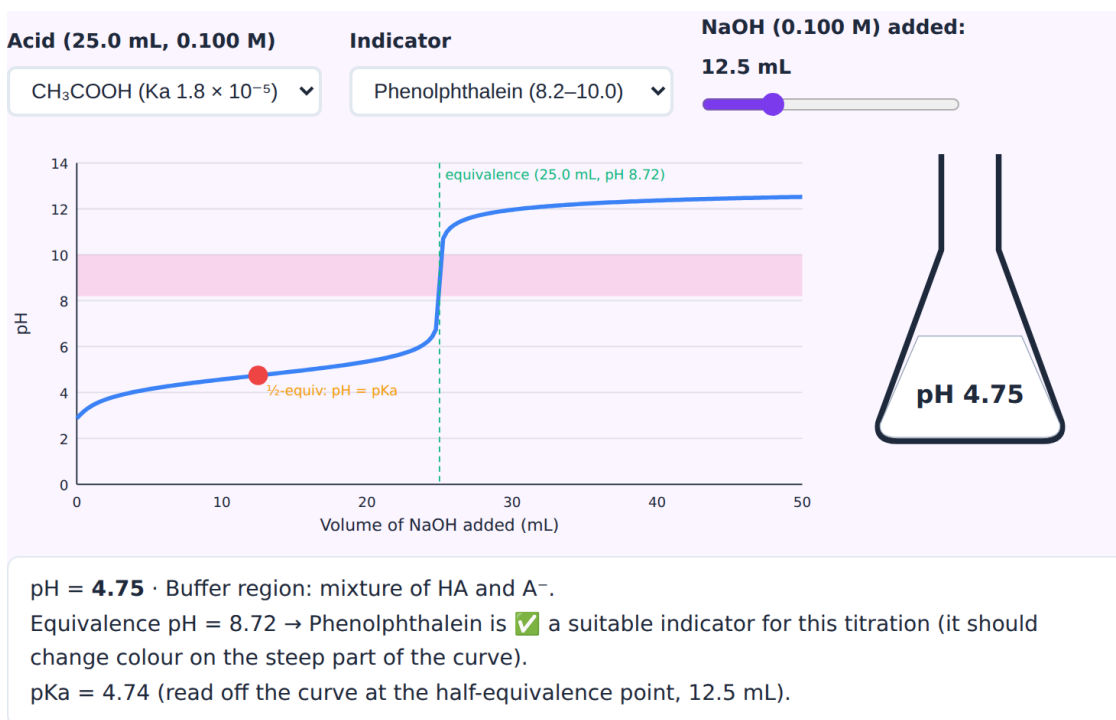
SOLUTION



$$K_{sp} = s(2s)^2 = 4s^3 \rightarrow s = \sqrt[3]{(3.9 \times 10^{-11} / 4)} = 2.1 \times 10^{-4} \text{ mol/L}$$

Interactive Simulation: Titration Curve Simulator

Titrate an acid with NaOH. Drag the burette slider to add base and watch the pH and indicator colour. Try a weak acid and see how the curve and equivalence point change.



Companion-app task — Titration Curve Simulator: locate the equivalence point and compare it with the midpoint of the buffer region. Explain why the equivalence-point pH depends on the acid and base strengths.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 15 Quiz

Q1. The pH of 0.01 M HCl is:

- A. 1
- B. 2
- C. 12
- D. 7

Q2. A buffer works best when:

- A. pH = 7
- B. $[\text{HA}] = [\text{A}^-]$
- C. pH = 14
- D. $[\text{HA}] \gg [\text{A}^-]$

Q3. The K_{sp} expression for CaF_2 is:

- A. $[\text{Ca}^{2+}][\text{F}^-]$
- B. $[\text{Ca}^{2+}][\text{F}^-]^2$
- C. $[\text{Ca}^{2+}]^2[\text{F}^-]$
- D. $[\text{Ca}][\text{F}]^2$

End-of-Chapter Exercises

Level 1 — Recall and understanding

15.1 Find the pH of 0.0050 M Ba(OH)₂.

Answer: _____

15.2 A solution has pH 4.30. Find [H⁺].

Answer: _____ M

15.3 Identify the conjugate acid–base pairs in $\text{HCO}_3^- + \text{H}_2\text{O} \rightleftharpoons \text{CO}_3^{2-} + \text{H}_3\text{O}^+$.

Level 2 — Application

15.4 Find the pH of 0.20 M HF ($K_a = 6.8 \times 10^{-4}$). Use the quadratic.

Answer: _____

15.5 Find the pH of a buffer that is 0.10 M NH₃ and 0.15 M NH₄Cl ($\text{p}K_a$ of NH₄⁺ = 9.25).

Answer: _____

Level 3 — Challenge

15.6 The molar solubility of PbI₂ is 1.52×10^{-3} mol/L. Calculate K_{sp} .

Answer: _____

15.7 Which indicator would you choose for titrating CH₃COOH with NaOH: methyl orange, bromothymol blue or phenolphthalein? Explain.

Chapter 16: Thermodynamics and Free Energy

Unit VI: Advanced Topics

16.1 Gibbs Free Energy

$$\Delta G = \Delta H - T\Delta S$$

$\Delta G < 0$: Spontaneous | $\Delta G = 0$: Equilibrium | $\Delta G > 0$: Non-spontaneous

$$\Delta G^\circ = -RT \ln K$$

ΔH	ΔS	Spontaneous?
–	+	Always (all T)
+	–	Never
–	–	Low T only
+	+	High T only

WORKED EXAMPLE

$\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$; $\Delta H^\circ = -92.2 \text{ kJ}$, $\Delta S^\circ = -198.8 \text{ J/K}$. Find ΔG° at 25°C .

SOLUTION

$$\Delta G^\circ = -92.2 - (298.15)(-0.1988) = -92.2 + 59.3 = -32.9 \text{ kJ (spontaneous)}$$

$$\text{Non-spontaneous above } T = \Delta H/\Delta S = 92.2/0.1988 = 464 \text{ K (191}^\circ\text{C)}$$

Additional Explanation: Spontaneity Depends on Both Enthalpy and Entropy

At constant temperature and pressure, Gibbs free-energy change provides a criterion for spontaneity: $\Delta G < 0$ is spontaneous in the forward direction under the stated conditions, $\Delta G = 0$ at equilibrium, and $\Delta G > 0$ means the forward direction is not spontaneous. Use consistent energy units when combining ΔH and $T\Delta S$, and remember that spontaneity does not specify how fast a process occurs.

Worked Example: Calculate ΔG

At 298 K, a process has $\Delta H = -40.0 \text{ kJ mol}^{-1}$ and $\Delta S = -100 \text{ J mol}^{-1} \text{ K}^{-1}$. Calculate ΔG .

Convert entropy to kJ units: $\Delta S = -0.100 \text{ kJ mol}^{-1} \text{ K}^{-1}$.

$$\text{Use } \Delta G = \Delta H - T\Delta S = -40.0 - [298(-0.100)] \text{ kJ mol}^{-1}.$$

$$\Delta G = -40.0 + 29.8 = -10.2 \text{ kJ mol}^{-1}.$$

Check: ΔG is negative at 298 K, so the forward process is thermodynamically favourable under these conditions. This does not guarantee a rapid rate.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 16.A: ΔS° from standard entropies

Calculate ΔS° for $\text{N}_2 + 3\text{H}_2 \rightarrow 2\text{NH}_3$. S° (J/(mol·K)): N_2 191.6, H_2 130.7, NH_3 192.5.

SOLUTION

$$\Delta S^\circ = 2(192.5) - [191.6 + 3(130.7)] = 385.0 - 583.7 = -198.7 \text{ J/K}$$

Negative, as expected: 4 mol of gas becomes 2 mol of gas.

Worked Example 16.B: ΔG° from free energies of formation

Find ΔG° for $2\text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}_2(\text{g})$. ΔG°_f (kJ/mol): NO 86.6, NO_2 51.3.

SOLUTION

$$\Delta G^\circ = 2(51.3) - [2(86.6) + 0] = -70.6 \text{ kJ} \rightarrow \text{spontaneous under standard conditions.}$$

Worked Example 16.C: Equilibrium constant from ΔG°

For the ammonia synthesis $\Delta G^\circ = -32.9 \text{ kJ}$ at 298 K. Find K.

SOLUTION

$$K = e^{-\Delta G^\circ/RT} = e^{32\,900/(8.314 \times 298.15)} = e^{13.27} = 5.8 \times 10^5$$

Worked Example 16.D: Predicting the sign of ΔS

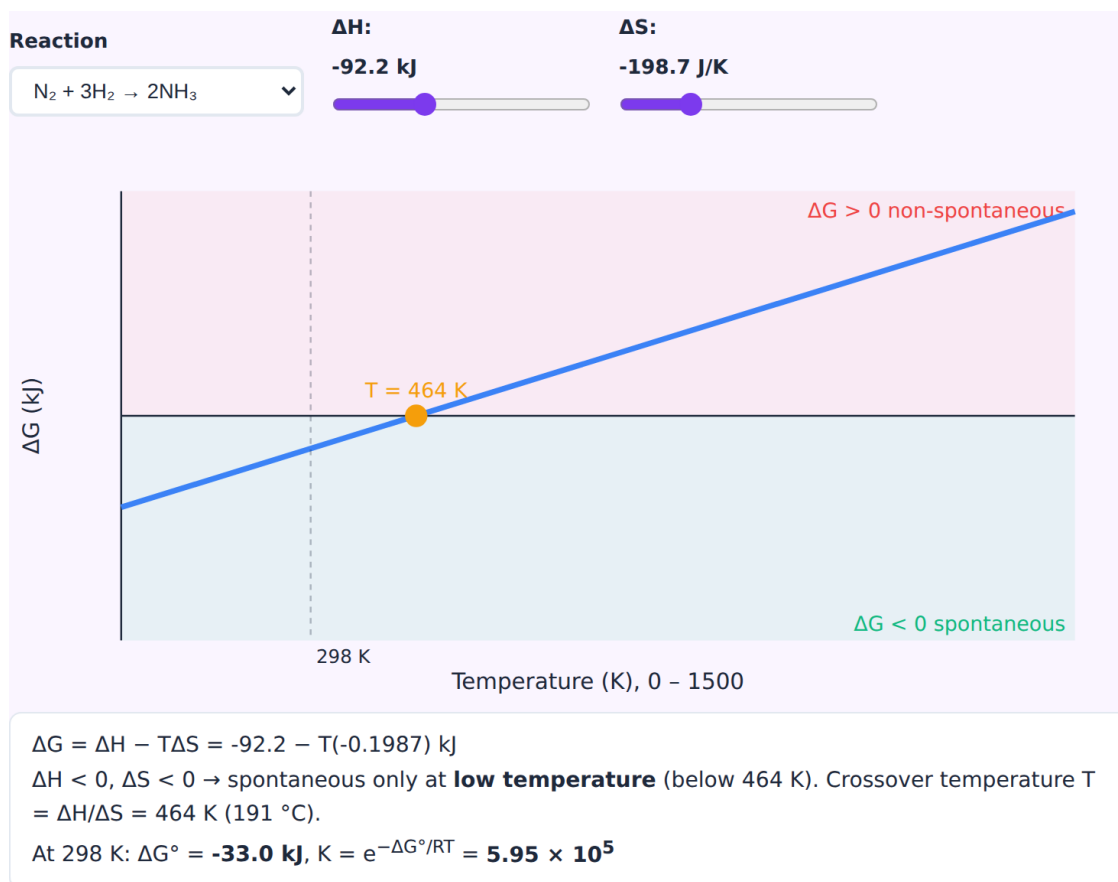
Predict the sign of ΔS for: (a) ice melting (b) $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{SO}_3(\text{g})$ (c) NaCl dissolving (d) $\text{CaCO}_3(\text{s}) \rightarrow \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$.

SOLUTION

(a) + (solid $\rightarrow$ liquid) (b) – (3 mol gas $\rightarrow$ 2 mol gas) (c) + (ions spread through the solution) (d) + (a gas is produced)

Interactive Simulation: ΔG vs Temperature

Set ΔH and ΔS (or choose a real reaction). The line shows $\Delta G = \Delta H - T\Delta S$. Below the axis the reaction is spontaneous.



Companion-app task — ΔG vs Temperature: vary temperature and locate where the sign of ΔG changes. Explain how the signs of ΔH and ΔS determine whether temperature can reverse spontaneity.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 16 Quiz

Q1. A reaction with $\Delta H < 0$ and $\Delta S > 0$ is spontaneous:

- A. Only at low T
- B. Only at high T
- C. At all temperatures
- D. Never

Q2. $\Delta G^\circ = -RT \ln K$. If $K > 1$, then ΔG° is:

- A. Positive
- B. Negative
- C. Zero
- D. Cannot determine

End-of-Chapter Exercises

Level 1 — Recall and understanding

16.1 Predict the sign of ΔS for: (a) $\text{H}_2\text{O}(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l})$ (b) $2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$.

Level 2 — Application

16.2 For $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$: $\Delta H^\circ = +178.3 \text{ kJ}$, $\Delta S^\circ = +160.5 \text{ J/K}$. Find ΔG° at 25°C .

Answer: _____ kJ

16.3 Above what temperature does the decomposition of CaCO_3 become spontaneous?

Answer: _____ K

16.4 Find K at 298 K for a reaction with $\Delta G^\circ = +10.0 \text{ kJ}$.

Answer: _____

Level 3 — Challenge

16.5 Find ΔG° at 298 K for a reaction with $K = 1.0 \times 10^{-5}$.

Answer: _____ kJ

16.6 Calculate ΔS° for $\text{CH}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l})$. $S^\circ (\text{J}/(\text{mol}\cdot\text{K}))$: CH_4 186.3, O_2 205.0, CO_2 213.7, $\text{H}_2\text{O}(\text{l})$ 69.9.

Answer: _____ J/K

Chapter 17: Electrochemistry

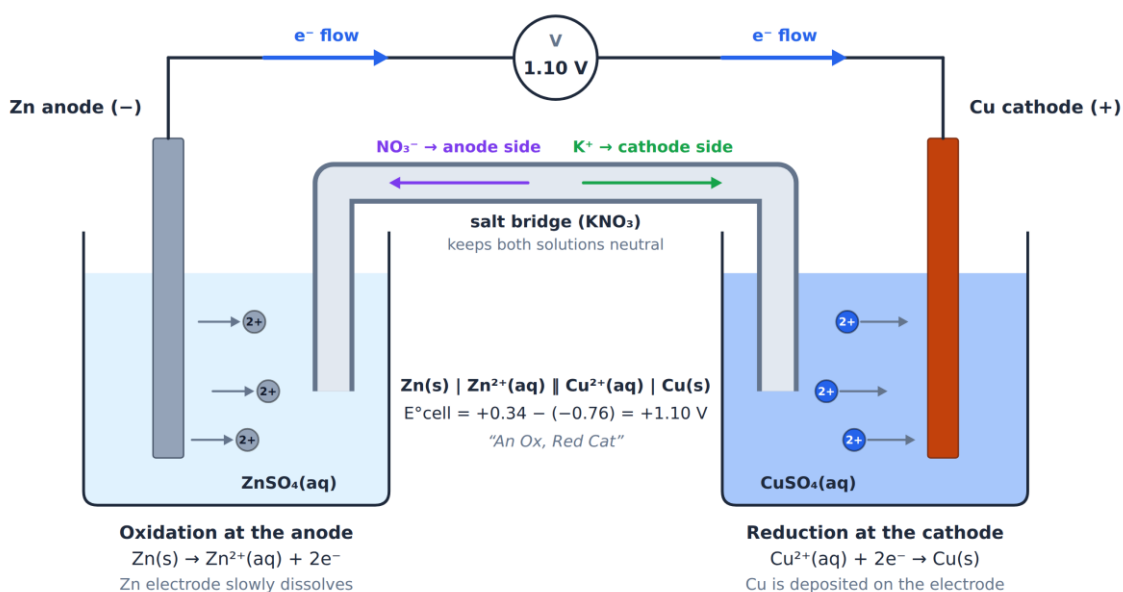
Unit VI: Advanced Topics

17.1 Galvanic Cells

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$$

$$\Delta G^{\circ} = -nFE^{\circ}_{\text{cell}} \quad (F = 96,485 \text{ C/mol})$$

$$\text{Nernst: } E = E^{\circ} - (0.0592/n) \log Q \quad (\text{at } 25^{\circ}\text{C})$$



Electrons flow through the external wire from anode to cathode; ions flow through the salt bridge to complete the circuit.

17.2 Electrolysis

$$\text{mol e}^- = It/F$$

$$\text{mass} = (It \times M)/(nF)$$

WORKED EXAMPLE

2.50 A through CuSO_4 for 45 min. Mass of Cu deposited?

SOLUTION

$$\text{mol e}^- = (2.50)(2700)/96485 = 0.06996 \text{ mol}$$

$$\text{mol Cu} = 0.06996/2 = 0.03498 \text{ mol}$$

$$\text{mass} = 0.03498 \times 63.55 = 2.22 \text{ g Cu}$$

Additional Explanation: Cell Potential Is a Difference of Reduction Potentials

Standard electrode potentials are tabulated as reduction potentials. For a galvanic cell, identify the cathode (reduction) and anode (oxidation), then calculate $E^{\circ}_{\text{cell}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}$ using the tabulated reduction potentials. A positive standard cell potential corresponds to a spontaneous cell reaction under standard conditions. Do not multiply an electrode potential by a stoichiometric coefficient when balancing electrons.

Worked Example: Zinc–Copper Galvanic Cell

Given $E^\circ(\text{Cu}^{2+}/\text{Cu}) = +0.34 \text{ V}$ and $E^\circ(\text{Zn}^{2+}/\text{Zn}) = -0.76 \text{ V}$, calculate E°_{cell} for a cell in which zinc is oxidised and copper(II) is reduced.

Cathode: $\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$, $E^\circ_{\text{cathode}} = +0.34 \text{ V}$.

Anode is zinc oxidation, but use the tabulated reduction potential in the subtraction: $E^\circ_{\text{anode}} = -0.76 \text{ V}$.

$E^\circ_{\text{cell}} = +0.34 - (-0.76) = +1.10 \text{ V}$.

Check: the positive value indicates a spontaneous cell reaction under standard conditions. The overall reaction is $\text{Zn} + \text{Cu}^{2+} \rightarrow \text{Zn}^{2+} + \text{Cu}$.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 17.A: Designing a galvanic cell

A cell is made from Zn/Zn^{2+} and Ag/Ag^+ half-cells. Find E°_{cell} , write the cell notation and overall equation, and find ΔG° .

SOLUTION

Higher E° is reduced (cathode): $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ (+0.80 V). Zn is oxidised (anode, -0.76 V).

$E^\circ_{\text{cell}} = 0.80 - (-0.76) = +1.56 \text{ V}$

Cell notation: $\text{Zn} | \text{Zn}^{2+} || \text{Ag}^+ | \text{Ag}$

Overall (multiply the Ag half by 2 — never multiply E°): $\text{Zn} + 2\text{Ag}^+ \rightarrow \text{Zn}^{2+} + 2\text{Ag}$, $n = 2$

$\Delta G^\circ = -nFE^\circ = -(2)(96\,485)(1.56) = -301 \text{ kJ}$

Worked Example 17.B: Nernst equation

Find E for the Daniell cell ($\text{Zn} | \text{Zn}^{2+} || \text{Cu}^{2+} | \text{Cu}$) when $[\text{Zn}^{2+}] = 1.0 \text{ M}$ and $[\text{Cu}^{2+}] = 0.010 \text{ M}$.

SOLUTION

$Q = [\text{Zn}^{2+}]/[\text{Cu}^{2+}] = 1.0/0.010 = 100$

$E = 1.10 - (0.0592/2) \log 100 = 1.10 - 0.0592 = 1.04 \text{ V}$

Worked Example 17.C: Time needed for electroplating

How long does it take to plate 10.0 g of silver with a current of 5.00 A?

SOLUTION

$n(\text{Ag}) = 10.0/107.87 = 0.0927 \text{ mol}$; $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$ so $n(\text{e}^-) = 0.0927 \text{ mol}$

$Q = nF = 0.0927 \times 96\,485 = 8945 \text{ C}$

$t = Q/I = 8945/5.00 = 1789 \text{ s} = 29.8 \text{ min}$

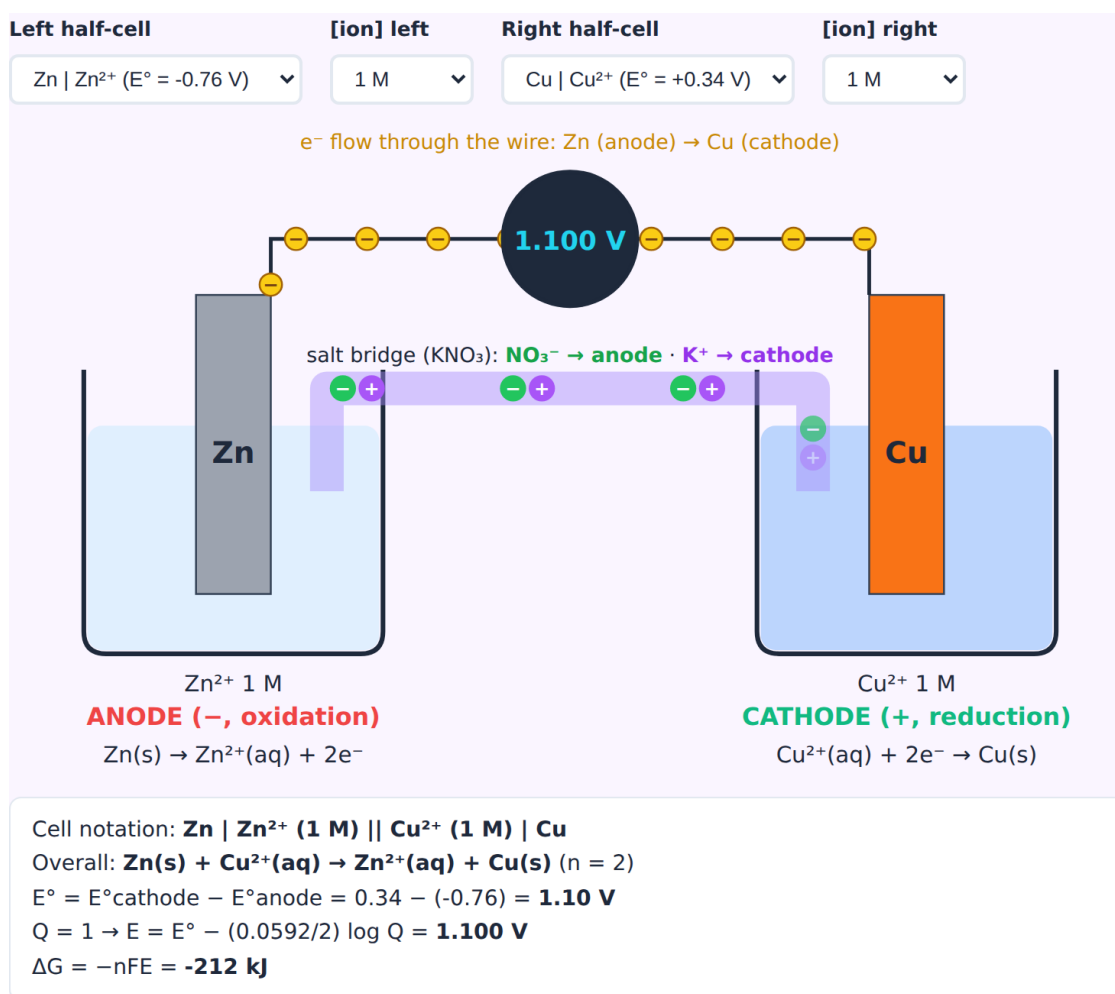
Worked Example 17.D: Equilibrium constant from E° Find K for $\text{Zn} + \text{Cu}^{2+} \rightleftharpoons \text{Zn}^{2+} + \text{Cu}$ at 25°C .**SOLUTION**

$$\log K = nE^\circ/0.0592 = (2)(1.10)/0.0592 = 37.2 \rightarrow K \approx 10^{37}$$

The reaction goes essentially to completion.

Interactive Simulation: Galvanic Cell Builder

Choose two half-cells and set the ion concentrations. The simulation identifies the anode and cathode, shows the electron flow and calculates E with the Nernst equation.



Companion-app task — Galvanic Cell Builder: identify anode, cathode, electron-flow direction, and cell potential. Explain why electrons travel through the external circuit while ions move through the electrolyte or salt bridge.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 17 Quiz

Q1. In a galvanic cell, oxidation occurs at the:

- A. Cathode
- B. Anode
- C. Salt bridge
- D. Voltmeter

Q2. E°_{cell} for $\text{Zn}|\text{Zn}^{2+}||\text{Cu}^{2+}|\text{Cu}$ is:

- A. +0.42 V
- B. +1.10 V
- C. -1.10 V
- D. +0.34 V

End-of-Chapter Exercises

Level 1 — Recall and understanding

17.1 Which is the strongest oxidising agent in the reference table of reduction potentials?

17.2 Calculate E°_{cell} for $\text{Mg}|\text{Mg}^{2+}||\text{Ag}^+|\text{Ag}$. (E° : Mg^{2+}/Mg -2.37 V, Ag^+/Ag +0.80 V)

Answer: _____ V

Level 2 — Application

17.3 Calculate ΔG° for a cell with $E^\circ = 1.10$ V and $n = 2$.

Answer: _____ kJ

17.4 Will copper metal dissolve in 1 M HCl? Explain using E° values.

17.5 Name the products at each electrode when molten NaCl is electrolysed.

Level 3 — Challenge

17.6 What mass of Al is deposited when a current of 10.0 A passes through molten Al_2O_3 for 1.00 h?

Answer: _____ g

17.7 Find E for $\text{Cu}|\text{Cu}^{2+}(1.0\text{ M})||\text{Ag}^+(0.010\text{ M})|\text{Ag}$. ($E^\circ_{\text{cell}} = 0.46$ V)

Answer: _____ V

Chapter 18: Nuclear Chemistry and Radioactivity

Unit VI: Advanced Topics

18.1 Types of Decay

Type	Particle	ΔA	ΔZ	Penetration
Alpha (α)	${}^4_2\text{He}$	-4	-2	Low (paper)
Beta (β^-)	${}^0_{-1}\text{e}$	0	+1	Medium (Al)
Gamma (γ)	${}^0_0\gamma$	0	0	High (Pb)
Positron (β^+)	${}^0_{+1}\text{e}$	0	-1	Medium

$$N = N_0 e^{(-\lambda t)} \quad t_{1/2} = 0.693/\lambda$$

$$E = \Delta mc^2 \quad (1 \text{ amu} = 931.5 \text{ MeV})$$

18.2 Half-Life Calculations

WORKED EXAMPLE

${}^{14}\text{C}$ ($t_{1/2} = 5730 \text{ yr}$): artifact has 72% of living ${}^{14}\text{C}$ activity. Age?

SOLUTION

$$t = (5730 \times \ln(1/0.720))/0.693 = (5730 \times 0.3285)/0.693 = 2720 \text{ years}$$

Additional Explanation: Half-Life Is a Fractional Rule

For radioactive decay, the half-life is the time required for half of the undecayed nuclei in a large sample to decay. After each half-life, the remaining fraction is halved: after n half-lives, the fraction remaining is $(1/2)^n$. Individual nuclear decays are random, but the behaviour of a large sample follows a predictable statistical pattern.

Worked Example: Mass Remaining After Three Half-Lives

A sample begins with 80.0 mg of a radioactive isotope. What mass remains after three half-lives?

After one half-life: $80.0 \times 1/2 = 40.0 \text{ mg}$.

After two half-lives: $40.0 \times 1/2 = 20.0 \text{ mg}$.

After three half-lives: $20.0 \times 1/2 = 10.0 \text{ mg}$, which is 12.5% of the initial sample.

Check: the mass remaining is $80.0 \times (1/2)^3 = 10.0 \text{ mg}$. This model describes the expected amount; it does not predict the exact decay time of a particular nucleus.

More Worked Examples

Try each problem yourself before reading the solution.

Worked Example 18.A: Balancing nuclear equations

Complete each: (a) $^{226}_{88}\text{Ra} \rightarrow ^4_2\text{He} + ?$ (b) $^{14}_6\text{C} \rightarrow ^0_{-1}\text{e} + ?$ (c) $^{18}_9\text{F} \rightarrow ^0_{+1}\text{e} + ?$ (d) $^{235}_{92}\text{U} + ^1_0\text{n} \rightarrow ^{141}_{56}\text{Ba} + ^{92}_{36}\text{Kr} + ?$

SOLUTION

Both the mass numbers and the atomic numbers must balance.

(a) $^{226}_{86}\text{Rn}$ (b) $^{14}_7\text{N}$ (c) $^{18}_8\text{O}$ (d) $236 = 141 + 92 + x \rightarrow 3\ ^1_0\text{n}$

Worked Example 18.B: Mass defect and binding energy

Find the binding energy per nucleon of ^4He . Masses (u): proton 1.00728, neutron 1.00866, ^4He nucleus 4.00150.

SOLUTION

Mass of separate nucleons = $2(1.00728) + 2(1.00866) = 4.03188\text{ u}$

Mass defect $\Delta m = 4.03188 - 4.00150 = 0.03038\text{ u}$

Binding energy = $0.03038 \times 931.5 = 28.3\text{ MeV} \rightarrow \text{per nucleon} = 28.3/4 = 7.07\text{ MeV}$

Worked Example 18.C: Amount remaining after a given time

I-131 (used to treat thyroid conditions) has $t_{1/2} = 8.02$ days. How much of a 100 mg dose remains after 30.0 days?

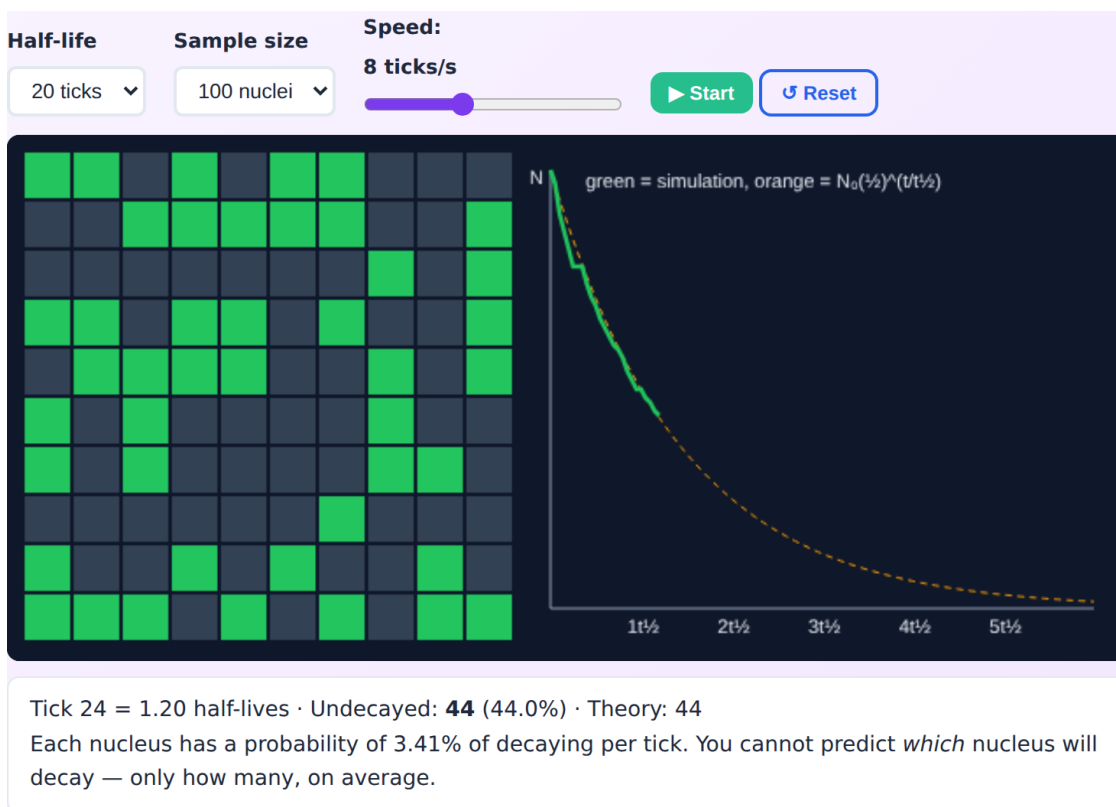
SOLUTION

Number of half-lives = $30.0/8.02 = 3.74$

$N = 100 \times (1/2)^{3.74} = 7.48\text{ mg}$

Interactive Simulation: Radioactive Decay Simulator

Each square is one nucleus. Every tick, each nucleus has a fixed chance of decaying. Watch how the random individual events produce a smooth exponential curve.



Companion-app task — Radioactive Decay Simulator: run several trials with different sample sizes. Compare the overall decay trend with the fluctuations expected in a small sample.

Screenshot of the interactive simulation. Try it yourself in the online edition at chemtextbook.neetapps.com.

Chapter 18 Quiz

Q1. Alpha decay of ^{238}U produces:

- A. ^{234}Th
- B. ^{236}U
- C. ^{234}Pa
- D. ^{238}Th

Q2. After 3 half-lives, fraction remaining is:

- A. 1/2
- B. 1/4
- C. 1/8
- D. 1/6

End-of-Chapter Exercises

Level 1 — Recall and understanding

18.1 Complete: $^{239}_{94}\text{Pu} \rightarrow ^4_2\text{He} + ?$

18.2 What is the product when ^{32}P undergoes β^- decay?

18.3 What fraction of a sample remains after 5 half-lives?

Answer: _____

Level 2 — Application

18.4 Co-60 has $t_{1/2} = 5.27$ years. How long does it take for 1.00 g to decay to 0.250 g?

Answer: _____ years

18.5 Calculate the decay constant of C-14 ($t_{1/2} = 5730$ years).

Answer: _____ per year

18.6 A bone contains 25.0% of the C-14 activity of living tissue. How old is it?

Answer: _____ years

Level 3 — Challenge

18.7 The mass defect of ^{56}Fe is 0.5285 u. Find its binding energy per nucleon.

Answer: _____ MeV

Appendix A: Periodic Table of the Elements

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57–71	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
87 Fr	88 Ra	89–103	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og
		57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
		89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Alkali Metal Alkaline Earth Transition Post-Trans. Metalloid Nonmetal Halogen Noble Gas Lanthanide
Actinide

Lanthanides (57–71) and actinides (89–103) are shown in the two rows below the main table.

Appendix B: Glossary of Key Terms

Term	Definition
Absolute zero	0 K (−273.15°C); theoretical minimum temperature
Activation energy (E_a)	Minimum energy for reaction; barrier between reactants and products
Avogadro's number	6.022×10^{23} entities per mole
Buffer	Solution resisting pH change; weak acid + conjugate base
Catalyst	Speeds reaction without being consumed; lowers E_a
Colligative properties	Depend on number of solute particles, not identity
Electronegativity	Ability to attract bonding electrons
Enthalpy (H)	Heat content at constant pressure; $\Delta H = q_p$
Entropy (S)	Measure of energy dispersal; number of microstates
Equilibrium constant (K)	Product/reactant ratio at equilibrium
Gibbs free energy (G)	$G = H - TS$; determines spontaneity
Henderson-Hasselbalch	$\text{pH} = \text{pK}_a + \log([A^-]/[HA])$
Hess's Law	ΔH is path-independent; additive over steps
Hund's Rule	Fill degenerate orbitals singly before pairing
Ideal gas law	$PV = nRT$
Le Châtelier's principle	Equilibrium shifts to counteract disturbance
Limiting reactant	Reactant consumed first; determines max product
Molarity (M)	mol solute / L solution
Mole	6.022×10^{23} entities; SI unit of amount
Nernst equation	$E = E^\circ - (0.0592/n)\log Q$
Octet rule	Atoms seek 8 valence electrons
Oxidation	Loss of electrons; increase in oxidation number
pH	$-\log[H^+]$; measure of acidity
Raoult's Law	$P_{\text{solution}} = X_{\text{solvent}} \times P^\circ_{\text{solvent}}$
Reduction	Gain of electrons; decrease in oxidation number
Resonance	Multiple valid Lewis structures; true structure is hybrid

Term	Definition
Stoichiometry	Quantitative relationships in balanced equations
Titration	Determining concentration using known solution
VSEPR	Predicts geometry by minimising e^- pair repulsion
van't Hoff factor (i)	Number of particles per formula unit in solution

Appendix C: Reference Data

Fundamental Constants

Constant	Symbol	Value
Speed of light	c	$2.998 \times 10^8 \text{ m/s}$
Planck's constant	h	$6.626 \times 10^{-34} \text{ J}\cdot\text{s}$
Avogadro's number	N_a	$6.022 \times 10^{23} \text{ mol}^{-1}$
Gas constant	R	$8.314 \text{ J}/(\text{mol}\cdot\text{K})$
Gas constant	R	$0.08206 \text{ L}\cdot\text{atm}/(\text{mol}\cdot\text{K})$
Faraday's constant	F	$96,485 \text{ C/mol}$
Boltzmann constant	k_B	$1.381 \times 10^{-23} \text{ J/K}$
Elementary charge	e	$1.602 \times 10^{-19} \text{ C}$

Standard Reduction Potentials (25°C)

Half-Reaction	$E^\circ \text{ (V)}$
$\text{F}_2 + 2\text{e}^- \rightarrow 2\text{F}^-$	+2.87
$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$	+1.51
$\text{Cl}_2 + 2\text{e}^- \rightarrow 2\text{Cl}^-$	+1.36
$\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$	+0.80
$\text{Fe}^{3+} + \text{e}^- \rightarrow \text{Fe}^{2+}$	+0.77
$\text{Cu}^{2+} + 2\text{e}^- \rightarrow \text{Cu}$	+0.34
$2\text{H}^+ + 2\text{e}^- \rightarrow \text{H}_2$	0.00
$\text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}$	-0.13
$\text{Ni}^{2+} + 2\text{e}^- \rightarrow \text{Ni}$	-0.25
$\text{Fe}^{2+} + 2\text{e}^- \rightarrow \text{Fe}$	-0.44
$\text{Zn}^{2+} + 2\text{e}^- \rightarrow \text{Zn}$	-0.76
$\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$	-1.66
$\text{Mg}^{2+} + 2\text{e}^- \rightarrow \text{Mg}$	-2.37
$\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$	-2.71

Half-Reaction	E° (V)
$\text{Li}^+ + \text{e}^- \rightarrow \text{Li}$	-3.05

Acid Dissociation Constants (K_a at 25°C)

Acid	K_a
HF	6.8×10^{-4}
CH_3COOH	1.8×10^{-5}
H_2CO_3 (K_{a1})	4.3×10^{-7}
H_3PO_4 (K_{a1})	7.5×10^{-3}
NH_4^+	5.6×10^{-10}

Key Formulas

$PV = nRT$ | $q = mc\Delta T$ | $\Delta G = \Delta H - T\Delta S$
 $\text{Rate} = k[A]^m$ | $t_{1/2} = 0.693/k$ (1st order)
 $K_c = [\text{products}]/[\text{reactants}]$ | $\text{pH} = -\log[\text{H}^+]$
 $\Delta G^\circ = -nFE^\circ$ | $E = E^\circ - (0.0592/n)\log Q$
 $N = N_0 e^{(-\lambda t)}$ | $E = mc^2$

Appendix D: Answers and Worked Solutions

Chapter 1: Matter, Measurement, and the Scientific Method

Quiz answers

- Q1. C** NaCl is a compound: two elements (Na and Cl) chemically bonded in a fixed ratio.
Q2. B Leading zeros don't count. The 4, 5, and trailing 0 (after decimal) are significant = 3 sig figs.
Q3. A $T_C = (5/9)(72-32) = (5/9)(40) = 22.2^\circ\text{C}$

Exercise solutions

- 1.1** (a) Homogeneous mixture (an alloy of Cu and Zn) (b) Compound (H_2O) (c) Element (O_2 molecules contain only O atoms) (d) Heterogeneous mixture (particles settle out).
1.2 (a) 4 — leading zeros don't count, the trailing zero after the decimal does. (b) 4 — all digits in the coefficient count. (c) Ambiguous (1 to 4). Write it in scientific notation, e.g. 5.0×10^3 for 2 sig figs.
1.3 $4.52 \times 0.0130 = 0.05876$; $\div 2.1 = 0.02798$. The least precise value (2.1) has 2 sig figs $\rightarrow$ **0.028**.
1.4 $V = 23.2 - 20.0 = 3.2 \text{ mL} = 3.2 \text{ cm}^3$ (2 sig figs). $\rho = 25.4/3.2 = \mathbf{7.9 \text{ g/cm}^3}$. This is close to iron (7.87 g/cm^3).
1.5 $T_F = (9/5)(-40) + 32 = -72 + 32 = \mathbf{-40^\circ\text{F}}$. This is the one temperature where the two scales agree.
1.6 $T_C = (5/9)(98.6 - 32) = 37.0^\circ\text{C}$; $T_K = 37.0 + 273.15 = \mathbf{310.2 \text{ K}}$.
1.7 Evaporation (liquid $\rightarrow$ gas) is endothermic. The fastest-moving water molecules escape from the skin, taking kinetic energy with them. The energy needed comes from the skin, so the skin's temperature drops.

Chapter 2: Atoms, Molecules, and Ions

Quiz answers

- Q1. A** Neutrons = $A - Z = 40 - 20 = 20$
Q2. B Fe^{3+} and O^{2-} : need 2 Fe^{3+} (+6) and 3 O^{2-} (-6) $\rightarrow \text{Fe}_2\text{O}_3$
Q3. C Isotopes have the same Z (protons) but different numbers of neutrons.

Exercise solutions

- 2.1** $Z = 19 \rightarrow$ 19 protons; neutrons = $39 - 19 = 20$; electrons = $19 - 1 = 18$.
2.2 (a) magnesium chloride (b) copper(II) oxide (c) sodium hydrogen carbonate (d) potassium permanganate.
2.3 (a) $(\text{NH}_4)_3\text{PO}_4$ (b) FeSO_4 (c) CaCO_3 .
2.4 $Z = 16 \rightarrow$ sulphur. $A = 16 + 16 = 32$. Charge = $16 - 18 = -2$. Symbol: $^{32}_{16}\text{S}^{2-}$.
2.5 $(0.6011)(68.926) + (0.3989)(70.925) = 41.431 + 28.292 = \mathbf{69.72 \text{ u}}$.
2.6 Let x = fraction of ^{11}B : $11.009x + 10.013(1 - x) = 10.81 \rightarrow 0.996x = 0.797 \rightarrow x = 0.800 \rightarrow \mathbf{80.0\% \text{ } ^{11}\text{B}}$ (and 20.0% ^{10}B).

Chapter 3: Stoichiometry and the Mole Concept

Quiz answers

- Q1. A** $3.50 \text{ g} \times (1 \text{ mol}/18.02 \text{ g}) \times (6.022 \times 10^{23} \text{ /mol}) = 1.17 \times 10^{23} \text{ molecules}$

Q2. B 4 mol H₂ needs 2 mol O₂, but only 1 mol available. O₂ is limiting.

Q3. C Divide all subscripts by 6: CH₂O

Exercise solutions

3.1 $n = 50.0/100.09 = \mathbf{0.500 \text{ mol}}$.

3.2 $2\text{Al} + 6\text{HCl} \rightarrow 2\text{AlCl}_3 + 3\text{H}_2$. Balance Cl with 6HCl (needs 2AlCl₃), then H gives 3H₂, then Al gives 2Al.

3.3 Each H₂SO₄ has 4 O atoms: $0.250 \times 4 \times 6.022 \times 10^{23} = \mathbf{6.02 \times 10^{23} \text{ atoms}}$.

3.4 C: $85.6/12.01 = 7.13 \text{ mol}$; H: $14.4/1.008 = 14.3 \text{ mol}$. Ratio 1 : 2 → **CH₂**.

3.5 $n = 5.85/58.44 = 0.1001 \text{ mol}$; $c = 0.1001/0.500 = \mathbf{0.200 \text{ mol/L}}$.

3.6 $n(\text{Al}) = 5.00/26.98 = 0.185 \text{ mol}$; $n(\text{Cl}_2) = 10.0/70.90 = 0.141 \text{ mol}$. Cl₂ needed for all Al = $0.185 \times 3/2 = 0.278 \text{ mol} > 0.141$, so **Cl₂ is limiting**. $n(\text{AlCl}_3) = 0.141 \times 2/3 = 0.0940 \text{ mol}$; $m = 0.0940 \times 133.33 = \mathbf{12.5 \text{ g}}$.

3.7 % yield = $10.9/12.54 \times 100 = \mathbf{86.9\%}$.

Chapter 4: Chemical Reactions in Aqueous Solution

Quiz answers

Q1. C NaCl is an ionic compound that dissociates 100% in water.

Q2. C $\text{K}(+1) + \text{Mn}(x) + 4\text{O}(-2) = 0 \rightarrow x = +7$

Q3. B BaSO₄ is insoluble (Ba²⁺ exception for SO₄²⁻)

Exercise solutions

4.1 HNO₃ strong (strong acid); NH₃ weak (weak base); C₂H₅OH non-electrolyte (molecular, doesn't ionise); KBr strong (soluble ionic).

4.2 H₂S: -2; SO₂: +4; H₂SO₄: +6; S₈: 0 (free element).

4.3 $\text{Ba}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightarrow \text{BaSO}_4(\text{s})$. K⁺ and Cl⁻ are spectators.

4.4 Mg: 0 → +2 (oxidised) so Mg is the **reducing agent**. O: 0 → -2 (reduced) so O₂ is the **oxidising agent**.

4.5 $n(\text{Ag}^+) = 0.200 \times 0.0500 = 0.0100 \text{ mol} = n(\text{AgCl})$; $m = 0.0100 \times 143.32 = \mathbf{1.43 \text{ g}}$.

4.6 $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$. $n(\text{NaOH}) = 0.150 \times 0.0324 = 4.86 \times 10^{-3} \text{ mol}$; $n(\text{H}_2\text{SO}_4) = \frac{1}{2} \times 4.86 \times 10^{-3} = 2.43 \times 10^{-3} \text{ mol}$; $c = 2.43 \times 10^{-3}/0.0200 = \mathbf{0.122 \text{ mol/L}}$.

4.7 Balance O with H₂O, H with H⁺, charge with e⁻: **$\text{MnO}_4^- + 8\text{H}^+ + 5\text{e}^- \rightarrow \text{Mn}^{2+} + 4\text{H}_2\text{O}$** . Check charge: -1 + 8 - 5 = +2.

Chapter 5: Thermochemistry

Quiz answers

Q1. B Exothermic means heat is released to surroundings, so ΔH is negative.

Q2. B $q = 50 \times 4.18 \times 20 = 4180 \text{ J} = 4.18 \text{ kJ}$

Q3. B Enthalpy depends only on initial and final states, not the pathway.

Exercise solutions

5.1 $q = mc\Delta T = 250 \times 4.18 \times 75.0 = 78\,375 \text{ J} = \mathbf{78.4 \text{ kJ}}$.

5.2 $n(\text{H}_2) = 10.0/2.016 = 4.96 \text{ mol}$. Heat = $4.96 \times (571.6/2) = \mathbf{1.42 \times 10^3 \text{ kJ released}}$.

- 5.3** Products: $2(-393.5) + 2(-285.8) = -1358.6$; reactants: $+52.3$. $\Delta H^\circ = -1358.6 - 52.3 = -1411 \text{ kJ}$.
- 5.4** Keep eq. 1; halve eq. 2: $\text{SO}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{SO}_3$, $\Delta H = -98.9 \text{ kJ}$. Add: $-296.8 + (-98.9) = -395.7 \text{ kJ}$.
- 5.5** Broken: $4(413) + 2(498) = 2648$. Formed: $2(799) + 4(463) = 3450$. $\Delta H \approx 2648 - 3450 = -802 \text{ kJ}$.
- 5.6** $q = (200)(4.18)(6.8) = 5685 \text{ J}$. $n(\text{H}_2\text{O}) = 0.100 \text{ mol}$. $\Delta H = -5.685 \text{ kJ}/0.100 \text{ mol} = -56.8 \text{ kJ/mol}$.

Chapter 6: Electronic Structure of Atoms

Quiz answers

Q1. B Max electrons = $2n^2 = 2(9) = 18$

Q2. B Cu is an exception to the Aufbau order: $[\text{Ar}]4s^1 3d^{10}$ (a completely filled 3d subshell) is lower in energy than $[\text{Ar}]4s^2 3d^9$.

Q3. B ℓ (angular momentum) determines shape: $0=s$, $1=p$, $2=d$, $3=f$

Exercise solutions

- 6.1** S: $1s^2 2s^2 2p^6 3s^2 3p^4$. S^{2-} gains 2 electrons: $1s^2 2s^2 2p^6 3s^2 3p^6$ (isoelectronic with Ar).
- 6.2** (a) $2n^2 = 32$ (b) $\ell = 3 \rightarrow 7 \text{ orbitals} \times 2 = 14$.
- 6.3** Mn: $[\text{Ar}]4s^2 3d^5$. Remove the two 4s electrons $\rightarrow \text{Mn}^{2+} = [\text{Ar}]3d^5 \rightarrow$ **5 unpaired**.
- 6.4** (b) is not allowed because ℓ must be less than n . (d) is not allowed because $|m_\ell|$ cannot be larger than ℓ . (a) and (c) are allowed.
- 6.5** $\nu = c/\lambda = (2.998 \times 10^8)/(500 \times 10^{-9}) = 6.00 \times 10^{14} \text{ Hz}$.
- 6.6** $E = hc/\lambda = (6.626 \times 10^{-34})(2.998 \times 10^8)/(400 \times 10^{-9}) = 4.97 \times 10^{-19} \text{ J}$.
- 6.7** Expected $[\text{Ar}]4s^2 3d^4$; actual **$[\text{Ar}]4s^1 3d^5$** . The 4s and 3d energies are very close, and a half-filled 3d⁵ subshell (six unpaired electrons, minimum electron–electron repulsion, maximum exchange energy) is lower in energy overall.

Chapter 7: Periodic Properties of the Elements

Quiz answers

Q1. D Radius increases down a group. I is lowest in Group 17.

Q2. D IE generally increases across Period 3, so Si is highest ($\approx 786 \text{ kJ/mol}$). Mg (≈ 738) is higher than Al (≈ 578) because Al's 3p electron is easier to remove than Mg's 3s electron, but this dip does not lift Mg above Si.

Q3. C Cations are smaller (lost e^- , more p^+/e^- ratio). Anions are larger (gained e^- , more repulsion).

Exercise solutions

- 7.1** **N** — it is furthest up and to the right of the four.
- 7.2** (a) S^{2-} — extra electrons increase repulsion. (b) Fe^{2+} — it has one more electron for the same nuclear charge.
- 7.3** **K < Ca < Br < Kr** (IE increases across Period 4).
- 7.4** Si: $Z = 14$, core $[\text{Ne}] = 10 \text{ electrons} \rightarrow Z_{\text{eff}} \approx +4$.
- 7.5** The big jump is after $\text{IE}_3 \rightarrow 3 \text{ valence electrons} \rightarrow$ **Group 13** (it is Al).
- 7.6** **$\text{S}^{2-} > \text{Cl}^- > \text{K}^+ > \text{Ca}^{2+}$** . All have 18 electrons; nuclear charge increases $16 \rightarrow 20$, pulling the same electron cloud in tighter.

- 7.7** N is $2p^3$ with one electron in each p orbital. O is $2p^4$, so one p orbital contains a pair. Repulsion between the paired electrons makes one of them easier to remove, despite O's higher nuclear charge.

Chapter 8: Chemical Bonding I: Lewis Structures

Quiz answers

- Q1. B** $S(6) + 4 \times O(6) + 2(\text{charge}) = 6 + 24 + 2 = 32$
Q2. B $FC = 5 - 0 - \frac{1}{2}(8) = 5 - 4 = +1$
Q3. C O_3 has two equivalent Lewis structures (double bond on either side).

Exercise solutions

- 8.1** PCl_3 : $5 + 21 = 26$. NO_3^- : $5 + 18 + 1 = 24$. NH_4^+ : $5 + 4 - 1 = 8$. XeF_4 : $8 + 28 = 36$.
8.2 BF_3 (B has $6 e^-$), PCl_5 (P has $10 e^-$) and NO (11 valence electrons — an odd number, so one atom must be short). CH_4 obeys the rule.
8.3 C: $4 - 2 - 3 = -1$. O: $6 - 2 - 3 = +1$. (The negative charge sits on the less electronegative atom, which is unusual.)
8.4 $24 e^-$; N has one $N=O$ and two $N-O$ bonds; the double bond can be at any of three positions $\rightarrow$ 3 resonance structures. Bond order = $4/3 \approx 1.33$.
8.5 S: $FC = 6 - 0 - \frac{1}{2}(8) = +2$. With two $S=O$ double bonds (expanded octet) the FC on S becomes 0, which is why that structure is often drawn.

Chapter 9: Molecular Geometry & Hybridisation

Quiz answers

- Q1. B** NH_3 has 3 bonds + 1 lone pair = 4 domains (tetrahedral e^- geometry), but molecular shape is trigonal pyramidal.
Q2. C 4 electron domains around C $\rightarrow sp^3$ hybridisation.
Q3. C CCl_4 is tetrahedral and symmetric — all four C-Cl dipoles cancel.

Exercise solutions

- 9.1** $BeCl_2$ sp ; SO_3 sp^2 ; NH_4^+ sp^3 ; PF_5 sp^3d ; SF_6 sp^3d^2 .
9.2 (a) 5 σ (4 C-H + 1 C-C) and 1 π . (b) 2 σ (H-C, C-N) and 2 π (in the triple bond).
9.3 $26 e^- \rightarrow$ 3 bonds + 1 lone pair on P (AX_3E) $\rightarrow$ **trigonal pyramidal**, angle less than 109.5° (about 100°).
9.4 BF_3 nonpolar (trigonal planar, symmetric); NF_3 polar (pyramidal); CH_2Cl_2 polar; XeF_4 nonpolar (square planar).
9.5 $7 + 21 = 28 e^- \rightarrow$ 3 bonds + 2 lone pairs (AX_3E_2). Lone pairs take equatorial positions of a trigonal bipyramid $\rightarrow$ **T-shaped**.
9.6 $3(7) + 1 = 22 e^-$. Central I has 2 bonds + 3 lone pairs (AX_2E_3). Lone pairs occupy all three equatorial positions $\rightarrow$ **linear**, 180° .

Chapter 10: Gases and Kinetic Molecular Theory

Quiz answers

- Q1. B** At STP ($0^\circ C$, 1 atm): $V = nRT/P = (1)(0.08206)(273.15)/1 = 22.4 \text{ L}$
Q2. B Gay-Lussac's Law: $P \propto T$ at constant V and n.

Q3. B $r_{\text{H}_2}/r_{\text{O}_2} = \sqrt{(32/2)} = \sqrt{16} = 4$

Exercise solutions

10.1 Boyle: $V_2 = P_1V_1/P_2 = (100)(4.00)/250 = \mathbf{1.60\text{ L}}$.

10.2 Charles: $V_2 = V_1T_2/T_1 = 500 \times 373.15/293.15 = \mathbf{636\text{ mL}}$.

10.3 $P = nRT/V = (0.500)(0.08206)(300.15)/10.0 = \mathbf{1.23\text{ atm}}$ (= 125 kPa).

10.4 $P(\text{N}_2) = 0.78 \times 101.3 = \mathbf{79.0\text{ kPa}}$.

10.5 $n(\text{Zn}) = 6.54/65.38 = 0.100\text{ mol} = n(\text{H}_2)$; $V = 0.100 \times 22.4 = \mathbf{2.24\text{ L}}$.

10.6 $n = PV/RT = (1.00)(0.500)/(0.08206 \times 298.15) = 0.02044\text{ mol}$. $M = 0.856/0.02044 = \mathbf{41.9\text{ g/mol}}$ (possibly propene, C_3H_6).

10.7 $r(\text{He})/r(\text{CH}_4) = \sqrt{M(\text{CH}_4)/M(\text{He})} = \sqrt{(16.04/4.003)} = \mathbf{2.00}$.

Chapter 11: Liquids, Solids, and Intermolecular Forces

Quiz answers

Q1. C London dispersion forces exist in all molecules due to temporary electron fluctuations.

Q2. B O is small and highly electronegative, so O–H bonds are very polar and water molecules form hydrogen bonds. S is larger and much less electronegative, so H_2S has only weaker dipole-dipole and London forces.

Q3. C At the triple point, solid, liquid, and gas phases all coexist in equilibrium.

Exercise solutions

11.1 (a) London dispersion (b) hydrogen bonding (c) dipole–dipole (d) ion–dipole.

11.2 $\text{CH}_4 < \text{C}_2\text{H}_6 < \text{C}_3\text{H}_8 < \text{C}_4\text{H}_{10}$. All are nonpolar; larger molecules have more electrons and stronger London forces.

11.3 $n = 25.0/18.02 = 1.387\text{ mol}$; $q = 1.387 \times 6.01 = \mathbf{8.34\text{ kJ}}$.

11.4 $n = 10.0/18.02 = 0.555\text{ mol}$; $q = 0.555 \times 40.7 = \mathbf{22.6\text{ kJ}}$.

11.5 Diamond: covalent network (very hard). Cu: metallic (conducts). NaCl: ionic (brittle, high mp). I_2 : molecular (soft, sublimates).

11.6 The solid–liquid line of water slopes to the left, so increasing pressure at $-5\text{ }^\circ\text{C}$ crosses into the liquid region: **the ice melts**.

11.7 Hydrogen bonding locks molecules in ice into an open hexagonal lattice, so ice is less dense (0.92 g/cm^3) than liquid water (1.00 g/cm^3).

Chapter 12: Solutions and Their Properties

Quiz answers

Q1. D AlCl_3 has $i=4$, producing most particles $\rightarrow$ greatest ΔT_f .

Q2. B Solvent flows from low solute concentration to high solute concentration through a semipermeable membrane.

Exercise solutions

12.1 $n = 10.0/40.00 = 0.250\text{ mol}$; $c = 0.250/0.250 = \mathbf{1.00\text{ mol/L}}$.

12.2 $V_1 = c_2V_2/c_1 = (0.30)(500)/6.0 = \mathbf{25\text{ mL}}$.

12.3 $n = 20.0/60.06 = 0.333\text{ mol}$; $m = 0.333/0.500 = \mathbf{0.666\text{ mol/kg}}$.

12.4 $n(\text{ethanol}) = 1.000 \text{ mol}$; $n(\text{water}) = 72.06/18.02 = 4.00 \text{ mol}$. $X = 1.000/5.00 = \mathbf{0.200}$.

12.5 $\Delta T_f = (3)(1.86)(0.250) = 1.40^\circ\text{C} \rightarrow T_f = \mathbf{-1.40^\circ\text{C}}$.

12.6 $m = \Delta T_f/K_f = 1.28/5.12 = 0.250 \text{ mol/kg}$. $n = 0.250 \times 0.0500 = 0.0125 \text{ mol}$. $M = 2.50/0.0125 = \mathbf{200 \text{ g/mol}}$.

12.7 $\Pi = iMRT = (1)(0.0500)(0.08206)(310.15) = \mathbf{1.27 \text{ atm}}$.

Chapter 13: Chemical Kinetics

Quiz answers

Q1. C $30 \text{ min} = 3 \text{ half-lives}$. $(1/2)^3 = 1/8$

Q2. B Catalysts lower activation energy, increasing rate without changing ΔH or K .

Exercise solutions

13.1 $t_{1/2} = 0.693/k = 0.693/0.0231 = \mathbf{30.0 \text{ min}}$.

13.2 $t_{1/2} = [A]_0/(2k) = 0.50/(2 \times 0.010) = \mathbf{25 \text{ s}}$.

13.3 $[A] = 0.80 e^{-(0.10)(15)} = 0.80 \times 0.223 = \mathbf{0.18 \text{ M}}$.

13.4 $1/[A] - 1/[A]_0 = kt \rightarrow 10 - 5 = 0.50t \rightarrow t = \mathbf{10 \text{ s}}$.

13.5 Doubling $[A]$ doubles the rate $\rightarrow$ first order in A . Doubling $[B]$ quadruples the rate $\rightarrow$ second order in B .
Rate = $k[A][B]^2$; $k = 4.0 \times 10^{-3}/((0.10)(0.10)^2) = \mathbf{4.0 \text{ M}^{-2}\text{s}^{-1}}$.

13.6 $\ln 4 = (E_a/8.314)(1/300 - 1/320) \rightarrow 1.386 = (E_a/8.314)(2.083 \times 10^{-4}) \rightarrow E_a = \mathbf{55.3 \text{ kJ/mol}}$.

13.7 The catalyst provides an alternative pathway with a lower activation energy (lower peak). The energies of reactants and products — and therefore ΔH — are unchanged. A larger fraction of collisions now have enough energy, so the rate increases.

Chapter 14: Chemical Equilibrium

Quiz answers

Q1. B $Q > K$ means too much product relative to equilibrium $\rightarrow$ shifts left (reverse).

Q2. D Catalysts speed both forward and reverse equally; K and position unchanged.

Exercise solutions

14.1 $K_c = [\text{NO}]^4[\text{H}_2\text{O}]^6 / ([\text{NH}_3]^4[\text{O}_2]^5)$

14.2 (a) Reverse $\rightarrow 1/4.0 = 0.25$. (b) Doubled coefficients $\rightarrow (4.0)^2 = 16$.

14.3 $K_c = (0.60)^2/((0.20)^2(0.10)) = 0.36/0.0040 = \mathbf{90}$.

14.4 (a) right (b) left — the forward reaction is exothermic (c) right — 2 mol gas vs 4 mol gas (d) no shift — partial pressures are unchanged.

14.5 $(2x)^2/(0.500 - x)^2 = 49.0 \rightarrow 2x/(0.500 - x) = 7.00 \rightarrow 2x = 3.50 - 7x \rightarrow x = 0.389$. $[\text{HI}] = 2x = \mathbf{0.778 \text{ M}}$.

14.6 $\Delta n = 2 - 4 = -2$. $K_p = 0.060 \times (0.08206 \times 500)^{-2} = 0.060/1683 = \mathbf{3.6 \times 10^{-5}}$.

Chapter 15: Acids and Bases

Quiz answers

Q1. B Strong acid: $[\text{H}^+] = 0.01 \text{ M}$. $\text{pH} = -\log(0.01) = 2$

Q2. B When $[\text{HA}] = [\text{A}^-]$, $\text{pH} = \text{pK}_a$, and the buffer resists both acid and base additions equally.

Q3. B $\text{CaF}_2 \rightleftharpoons \text{Ca}^{2+} + 2\text{F}^-$, so $K_{\text{sp}} = [\text{Ca}^{2+}][\text{F}^-]^2$

Exercise solutions

15.1 $[\text{OH}^-] = 2 \times 0.0050 = 0.010 \text{ M}$; $\text{pOH} = 2.00$; $\text{pH} = \mathbf{12.00}$.

15.2 $[\text{H}^+] = 10^{-4.30} = \mathbf{5.0 \times 10^{-5} \text{ M}}$.

15.3 HCO_3^- (acid) / CO_3^{2-} (conjugate base); H_2O (base) / H_3O^+ (conjugate acid).

15.4 $x^2 + 6.8 \times 10^{-4}x - 1.36 \times 10^{-4} = 0 \rightarrow x = 0.0113 \text{ M}$. $\text{pH} = \mathbf{1.95}$. (The shortcut gives 1.93; 5.8% ionisation, so the quadratic is needed.)

15.5 $\text{pH} = 9.25 + \log(0.10/0.15) = 9.25 - 0.18 = \mathbf{9.07}$.

15.6 $K_{\text{sp}} = [\text{Pb}^{2+}][\text{I}^-]^2 = s(2s)^2 = 4s^3 = 4(1.52 \times 10^{-3})^3 = \mathbf{1.40 \times 10^{-8}}$.

15.7 Phenolphthalein (range 8.2–10). The equivalence point is basic (about pH 8.7) because the acetate ion is a weak base.

Chapter 16: Thermodynamics and Free Energy

Quiz answers

Q1. C $\Delta G = (-) - T(+)$ is always negative regardless of T.

Q2. B If $K > 1$, $\ln K > 0$, so $\Delta G^\circ = -(\text{positive})(\text{positive}) = \text{negative}$.

Exercise solutions

16.1 (a) negative — gas to liquid. (b) positive — a gas is formed.

16.2 $\Delta G^\circ = 178.3 - (298.15)(0.1605) = 178.3 - 47.9 = \mathbf{+130 \text{ kJ}}$ $\rightarrow$ not spontaneous at 25 °C.

16.3 $T = \Delta H/\Delta S = 178.3/0.1605 = \mathbf{1111 \text{ K}}$ (838 °C) — why lime kilns run so hot.

16.4 $K = e^{-10\,000/(8.314 \times 298.15)} = e^{-4.03} = \mathbf{0.0177}$.

16.5 $\Delta G^\circ = -RT \ln K = -(8.314)(298.15)(-11.51) = 28\,500 \text{ J} = \mathbf{+28.5 \text{ kJ}}$.

16.6 $\Delta S^\circ = [213.7 + 2(69.9)] - [186.3 + 2(205.0)] = 353.5 - 596.3 = \mathbf{-242.8 \text{ J/K}}$.

Chapter 17: Electrochemistry

Quiz answers

Q1. B An Ox, Red Cat: Oxidation at Anode, Reduction at Cathode.

Q2. B $E^\circ = E^\circ\text{Cu}(0.34) - E^\circ\text{Zn}(-0.76) = \mathbf{+1.10 \text{ V}}$

Exercise solutions

17.1 F_2 — it has the most positive E° (+2.87 V), so it is the most easily reduced.

17.2 $E^\circ = 0.80 - (-2.37) = \mathbf{+3.17 \text{ V}}$.

17.3 $\Delta G^\circ = -nFE^\circ = -(2)(96\,485)(1.10) = \mathbf{-212 \text{ kJ}}$.

17.4 $\text{Cu} + 2\text{H}^+ \rightarrow \text{Cu}^{2+} + \text{H}_2$; $E^\circ = 0.00 - 0.34 = -0.34 \text{ V}$. E° is negative, so the reaction is **not spontaneous**: Cu does not dissolve.

17.5 Cathode (reduction): $\text{Na}^+ + \text{e}^- \rightarrow \text{Na(l)}$. Anode (oxidation): $2\text{Cl}^- \rightarrow \text{Cl}_2(\text{g}) + 2\text{e}^-$.

17.6 $Q = 10.0 \times 3600 = 36\,000 \text{ C}$; $n(\text{e}^-) = 36\,000/96\,485 = 0.373 \text{ mol}$; $\text{Al}^{3+} + 3\text{e}^- \rightarrow \text{Al}$ so $n(\text{Al}) = 0.124 \text{ mol}$; $m = 0.124 \times 26.98 = \mathbf{3.36 \text{ g}}$.

17.7 $\text{Cu} + 2\text{Ag}^+ \rightarrow \text{Cu}^{2+} + 2\text{Ag}$, $n = 2$. $Q = 1.0/(0.010)^2 = 1.0 \times 10^4$. $E = 0.46 - (0.0592/2)(4) = \mathbf{0.34 \text{ V}}$.

Chapter 18: Nuclear Chemistry and Radioactivity

Quiz answers

Q1. A A: $238 - 4 = 234$, Z: $92 - 2 = 90 \rightarrow {}^{234}_{90}\text{Th}$

Q2. C $(1/2)^3 = 1/8$

Exercise solutions

18.1 ${}^{235}_{92}\text{U}$

18.2 ${}^{32}_{15}\text{P} \rightarrow {}^0_{-1}\text{e} + {}^{32}_{16}\text{S}$ (Z increases by 1, A unchanged).

18.3 $(1/2)^5 = 1/32 = 0.0313$ (3.13%).

18.4 $0.250/1.00 = 1/4 = (1/2)^2 \rightarrow 2$ half-lives = **10.5 years**.

18.5 $\lambda = 0.693/t_{1/2} = 0.693/5730 = 1.21 \times 10^{-4} \text{ y}^{-1}$.

18.6 $25\% = (1/2)^2 \rightarrow 2 \times 5730 = 11\,460$ years.

18.7 $\text{BE} = 0.5285 \times 931.5 = 492.3 \text{ MeV}$; per nucleon = $492.3/56 = 8.79 \text{ MeV}$. Fe-56 is near the top of the binding-energy curve, which is why both fission and fusion release energy.