

ESAT

Chemistry

Moles and reacting masses

Free lesson notes

Covers

C4.1 to C4.7

Format

Lesson, worked examples, self-check

Price

Free

About this lesson

Chemical equations tell us how *particles* combine: two molecules of hydrogen with one of oxygen. But in a laboratory we cannot count particles, we can only weigh things. The **mole** is the bridge between the two. It lets us turn a mass in grams into a number of particles, and then use the equation to say how much of everything else is needed or made.

This topic is the quantitative backbone of GCSE-to-A-level chemistry, and the ESAT tests it heavily because it can be examined without a calculator, provided the numbers are chosen well. We start from relative atomic and formula masses, define the mole and the Avogadro constant, and then use the single equation $n = m / M_r$ in a series of applications: percentage composition, empirical and molecular formulae, reacting masses, limiting reactants, and finding an equation from experimental masses or gas volumes. (Concentration, gas volumes at rtp and percentage yield come in the next lesson.)

How to use it. Cover each worked example and try it before reading the solution. In the real ESAT, relative atomic masses are given in the question when needed; here we use the values below.

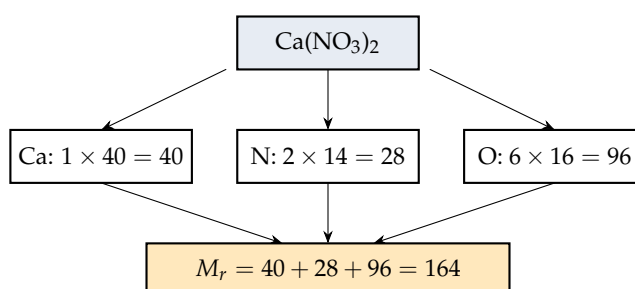
H	C	N	O	Na	Mg	Al	S	Cl	K	Ca	Fe	Cu	Zn
1	12	14	16	23	24	27	32	35.5	39	40	56	64	65

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1 Relative atomic mass and relative formula mass (C4.1)

Atoms are far too light to weigh in grams one at a time, so we compare them instead. The **relative atomic mass** A_r of an element is the average mass of its atoms compared with $\frac{1}{12}$ of the mass of a carbon-12 atom. By this scale carbon-12 has $A_r = 12$, hydrogen about 1, oxygen 16. It is a ratio, so it has no units. Where an element has several isotopes (such as chlorine, which is about three-quarters ^{35}Cl and one-quarter ^{37}Cl) the A_r is the weighted average, which is why chlorine's A_r is 35.5.

The **relative formula mass** M_r of a compound is found by adding the relative atomic masses of all the atoms in its formula. (For a simple molecule it is called the relative molecular mass; the method is identical.) Atoms in brackets are multiplied by the number after the bracket.



Worked example: relative formula masses

Find M_r for (a) H_2O , (b) CaCO_3 , (c) $(\text{NH}_4)_2\text{SO}_4$.

(a) $2 \times 1 + 16 = 18$. (b) $40 + 12 + 3 \times 16 = 100$. (c) Two ammonium ions, each $14 + 4 \times 1 = 18$, so $2 \times 18 = 36$; sulfate $32 + 4 \times 16 = 96$; total $36 + 96 = 132$.

Common mistakes

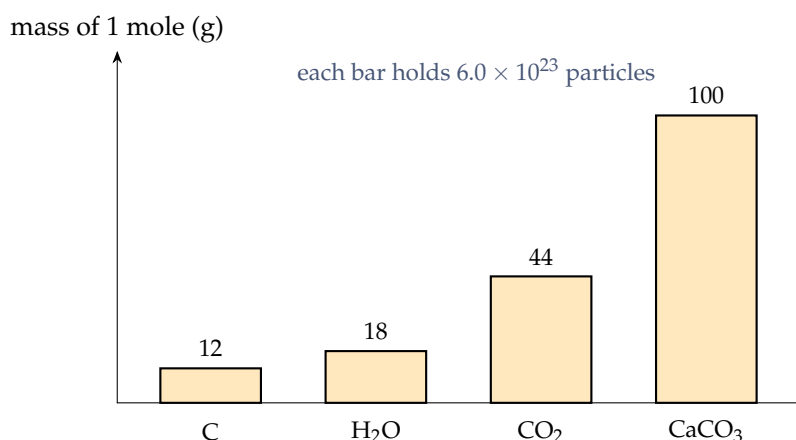
- Forgetting to multiply everything inside a bracket: in $\text{Ca}(\text{OH})_2$ there are 2 O and 2 H, so $M_r = 40 + 2 \times 17 = 74$.

- Giving units for A_r or M_r . They are ratios and have none (the *molar mass* in g/mol has the same number).
- Adding the number of atoms instead of their masses: CO_2 is $12 + 2 \times 16 = 44$, not $1 + 2$.

2 The mole and the Avogadro constant (C4.2)

Chemists count particles in a unit big enough to be useful: the **mole**, which is a very large number, like a “dozen” of particles. One mole contains 6.0×10^{23} particles (more exactly 6.02×10^{23}). This is the **Avogadro constant**, with the symbol N_A and units of “per mole”. The particles can be atoms, molecules, ions or electrons: always say which.

The number 6.0×10^{23} is chosen for a reason. It is the number of atoms in exactly 12 g of carbon-12. Because A_r compares atoms with carbon-12, it follows that **the mass of one mole of a substance, in grams, is its A_r or M_r** . This is called the *molar mass*: one mole of carbon is 12 g, of water 18 g, of calcium carbonate 100 g. Every one of these samples holds the same number of particles.



To convert between amount and number of particles:

$$\text{number of particles} = n \times 6.0 \times 10^{23}.$$

So 0.5 mol of water holds $0.5 \times 6.0 \times 10^{23} = 3.0 \times 10^{23}$ molecules. Each molecule has three atoms, so there are 9.0×10^{23} atoms in all. Counting the right kind of particle is the whole point: one mole of CaCl_2 contains one mole of Ca^{2+} ions but *two* moles of Cl^- ions.

Worked example: counting particles

How many hydrogen atoms are there in 0.25 mol of methane, CH_4 ?

Molecules: $0.25 \times 6.0 \times 10^{23} = 1.5 \times 10^{23}$. Each has 4 H atoms, so $4 \times 1.5 \times 10^{23} = 6.0 \times 10^{23}$ hydrogen atoms.

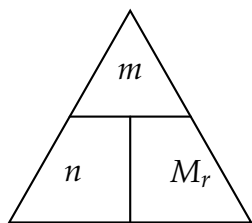
Common mistakes

- Mixing up the number of moles and the number of particles. A mole is a *number* of particles, not a mass.
- Forgetting that formulae contain several atoms: one mole of H_2O contains 2 moles of H atoms.
- Believing that one mole of every substance has the same mass. They have the same *number* of particles, but the mass differs.

3 Moles from mass (C4.3)

Combining the last section's facts gives the main equation of the topic. If one mole has a mass of M_r grams, then a sample of mass m grams contains $m \div M_r$ moles:

$$n = \frac{m}{M_r} \quad \Longleftrightarrow \quad m = n \times M_r \quad \Longleftrightarrow \quad M_r = \frac{m}{n}.$$



$$n = \frac{m}{M_r}$$

n in mol

$$m = n \times M_r$$

m in g

$$M_r = \frac{m}{n}$$

M_r in g/mol

Cover the quantity you want in the triangle and read off the formula. Always write down what you are calculating, with units: n in mol, m in g.

Worked example: mass to moles and back

(a) How many moles are there in 36 g of water? (b) What is the mass of 0.25 mol of carbon dioxide?
 (a) $M_r(\text{H}_2\text{O}) = 18$, so $n = 36/18 = 2$ mol. (b) $M_r(\text{CO}_2) = 44$, so $m = 0.25 \times 44 = 11$ g.

Worked example: finding M_r from a measurement

3.2 g of a gas is 0.05 mol. What is its relative molecular mass, and which of CH_4 , O_2 , CO_2 is it?
 $M_r = m/n = 3.2/0.05 = 64$. Neither: CH_4 is 16, O_2 is 32, CO_2 is 44. (It could be SO_2 : $32 + 32 = 64$.)

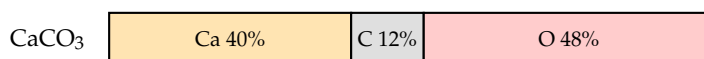
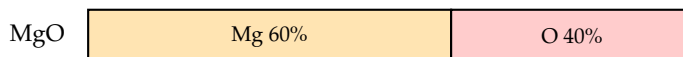
Common mistakes

- Using the mass of one atom when the substance is a molecule: for oxygen gas, O_2 , $M_r = 32$, not 16.
- Dividing the wrong way round. Sense-check: a small mass of a heavy substance should be a small fraction of a mole.
- Forgetting that masses must be in grams. $1 \text{ kg} = 1000 \text{ g}$.

4 Percentage composition by mass (C4.4)

The **percentage by mass** of an element in a compound tells you what fraction of the mass is that element:

$$\% \text{ of element} = \frac{(\text{number of atoms of the element in the formula}) \times A_r}{M_r} \times 100.$$



$$\text{percentage by mass} = \frac{\text{mass of the element in one formula unit}}{M_r} \times 100$$

In magnesium oxide, MgO , $M_r = 24 + 16 = 40$, so $\% \text{ Mg} = \frac{24}{40} \times 100 = 60\%$ and $\% \text{ O} = \frac{16}{40} \times 100 = 40\%$. The percentages of all the elements add to 100%, which is a useful check. In calcium carbonate: Ca 40%, C 12%, O 48%, adding to 100%.

The same idea gives the percentage of water in a hydrated salt, or of an element in a mixture or an impure sample. Note that percentage composition depends only on the formula, not on how much of the compound you have.

Worked example: nitrogen in a fertiliser

What is the percentage by mass of nitrogen in ammonium nitrate, NH_4NO_3 ?
 $M_r = 14 + 4 \times 1 + 14 + 3 \times 16 = 80$. There are 2 N atoms: $2 \times 14 = 28$. $\% \text{ N} = \frac{28}{80} \times 100 = 35\%$.

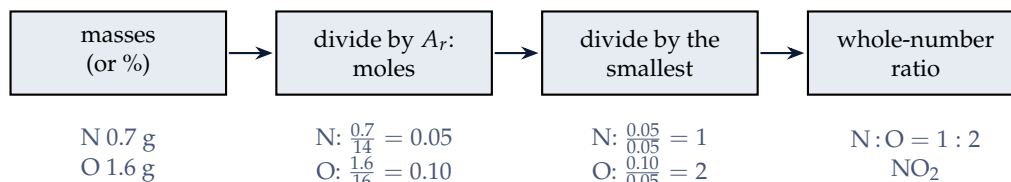
Common mistakes

- Counting only one nitrogen atom in NH_4NO_3 : it has two, in different parts of the formula.
- Dividing by the mass of one element instead of M_r .
- Forgetting that the percentages should add up to 100%; if they do not, check the arithmetic.

5 Empirical and molecular formulae (C4.5)

The **empirical formula** is the simplest whole-number ratio of atoms of each element in a compound. The **molecular formula** is the actual number of atoms in a molecule. Ethene has molecular formula C_2H_4 and empirical formula CH_2 . Ionic compounds are always written as empirical formulae.

Finding the empirical formula. In an experiment you measure masses (or percentages). Masses cannot be compared directly because atoms differ in mass; dividing by A_r turns each mass into moles, and moles are proportional to numbers of atoms.



1. Write the mass (or percentage) of each element.
2. Divide each by its A_r to get moles.
3. Divide all the answers by the smallest one.
4. If the ratios are not whole numbers, multiply them all by a small integer (e.g. $1.5 \rightarrow \times 2$; $1.33 \rightarrow \times 3$).

Molecular formula. Calculate the M_r of the empirical formula, divide the given M_r by it to get a whole number k , and multiply every subscript by k .

Water of crystallisation. Some salts are **hydrated**: their crystals contain water molecules, written $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$. The same method finds x : heat the salt, measure the mass lost (water), and find the mole ratio of water to the anhydrous salt.

Combustion data. If a hydrocarbon burns completely, all its carbon appears in CO_2 and all its hydrogen in H_2O . Moles of C = moles of CO_2 , and moles of H = $2 \times$ moles of H_2O .

Worked example: empirical formula

A compound of iron and oxygen contains 5.6 g of iron and 2.4 g of oxygen. Find its empirical formula.
Fe: $5.6/56 = 0.10$ mol. O: $2.4/16 = 0.15$ mol. Divide by 0.10: 1 : 1.5. Multiply by 2: 2 : 3. The formula is Fe_2O_3 .

Worked example: molecular formula

A hydrocarbon has empirical formula CH_2 and $M_r = 56$. What is its molecular formula?
 $M_r(\text{CH}_2) = 14$. $56/14 = 4$, so the molecular formula is C_4H_8 .

Worked example: water of crystallisation

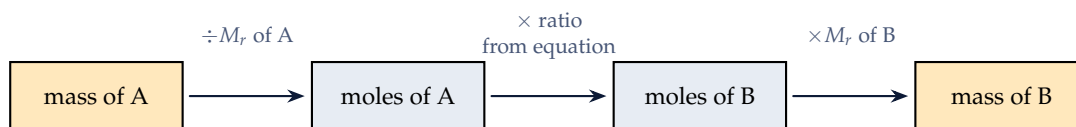
12.5 g of hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot x\text{H}_2\text{O}$, is heated until all the water has been driven off, leaving 8.0 g of anhydrous CuSO_4 . Find x .
Water lost = $12.5 - 8.0 = 4.5$ g, which is $4.5/18 = 0.25$ mol. CuSO_4 : $M_r = 64 + 32 + 64 = 160$, so $8.0/160 = 0.05$ mol. Ratio water : salt = $0.25 : 0.05 = 5 : 1$, so $x = 5$.

Common mistakes

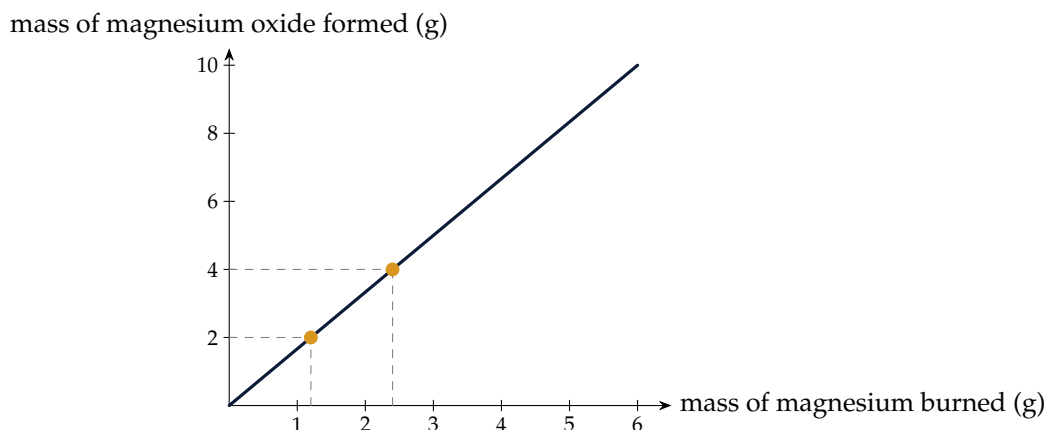
- Using the masses as if they were ratios of atoms, without dividing by A_r .
- Rounding 1.5 down to 1 or up to 2. A ratio such as 1 : 1.5 means multiply by two.
- Giving the empirical formula when the molecular formula is asked for (or vice versa).

6 Reacting masses and limiting reactants (C4.6)

A balanced equation gives the ratio of *moles* of reactants and products. That ratio, never a ratio of masses, is what we use. To find the mass of one substance from the mass of another:

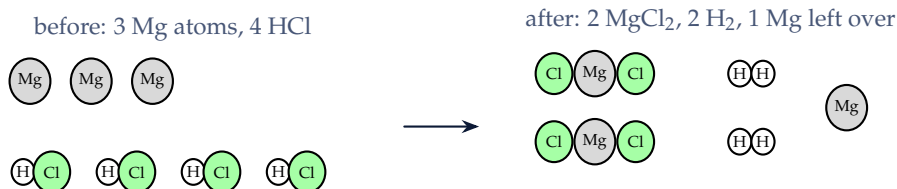
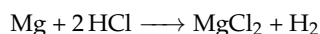


Worked through. What mass of magnesium oxide is made by burning 2.4 g of magnesium? $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$. Moles of Mg = $2.4/24 = 0.10$. The equation says 2 mol Mg gives 2 mol MgO, so 0.10 mol MgO. Mass = $0.10 \times 40 = 4.0$ g. Since the mass of product is proportional to the mass of reactant, the graph is a straight line through the origin:



Magnesium oxide forms at $\frac{40}{24} = \frac{5}{3}$ times the mass of the magnesium, because the extra mass is the oxygen that combined. So the oxygen mass is $\frac{2}{3}$ of the magnesium mass (e.g. 1.6 g for 2.4 g).

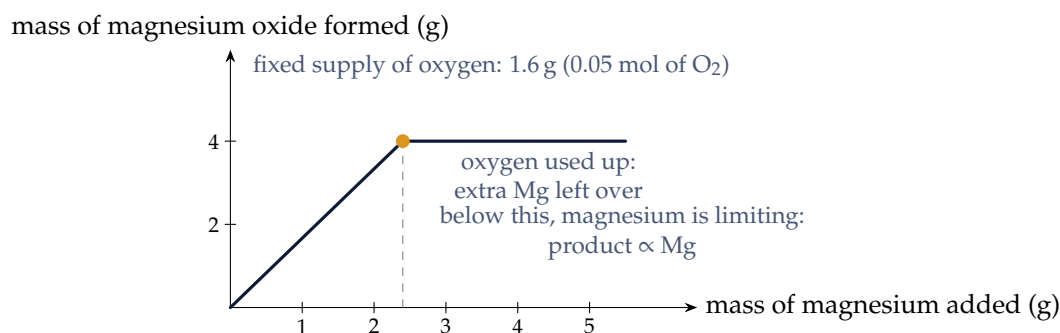
Limiting reactants. Reactants are not always mixed in exactly the right proportions. The reactant that is used up first is the **limiting reactant**: it controls how much product forms. The other is **in excess**, and some is left over.



HCl runs out first (needs 2 per Mg), so HCl is limiting

In the picture, $\text{Mg} + 2\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$: each magnesium atom needs two HCl. Three Mg would need six HCl, but only four are available, so HCl limits the reaction: $4/2 = 2$ units of Mg react, giving two H₂, and one Mg is left.

Method: (1) convert both masses to moles; (2) divide each by its coefficient; the smaller answer is the limiting reactant; (3) use the limiting reactant's moles to find the product. If the supply of one reactant is fixed, the graph of product against the other rises and then levels off:



Worked example: reacting masses

What mass of calcium oxide is made by heating 50 g of calcium carbonate? $\text{CaCO}_3 \longrightarrow \text{CaO} + \text{CO}_2$
 $M_r(\text{CaCO}_3) = 100$: $50/100 = 0.5$ mol. The ratio is 1 : 1, so 0.5 mol of CaO. $M_r(\text{CaO}) = 56$, mass = $0.5 \times 56 = 28$ g.

Worked example: a limiting reactant

5.6 g of iron and 2.0 g of sulfur are heated together: $\text{Fe} + \text{S} \longrightarrow \text{FeS}$. Which is limiting, and what mass of iron(II) sulfide forms?

Fe: $5.6/56 = 0.100$ mol. S: $2.0/32 = 0.0625$ mol. The ratio is 1 : 1, so whichever has fewer moles runs out first: *sulfur* is limiting and iron is in excess. Moles of FeS = 0.0625; $M_r = 88$, so mass = $0.0625 \times 88 = 5.5$ g.

Common mistakes

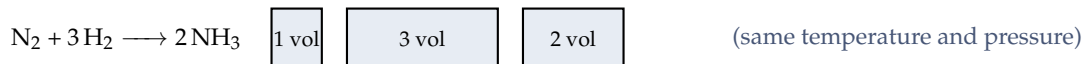
- Using mass ratios instead of mole ratios. The equation $2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$ does *not* say that 2 g of Mg reacts with 1 g of O_2 .
- Assuming the reactant with the smaller *mass* is limiting. Compare moles divided by coefficients.
- Calculating the product from the reactant in excess.

7 Finding an equation from masses or gas volumes (C4.7)

The mole ratios in an equation can be *measured*, and so used to find the equation itself.

From reacting masses. Convert each mass to moles, then divide by the smallest to get the ratio of the coefficients. For example, 4.0 g of a metal M ($A_r = 40$) combine with 3.2 g of oxygen atoms (O, $A_r = 16$): M: $4.0/40 = 0.10$ mol; O: $3.2/16 = 0.20$ mol. The ratio is 1 : 2 so the oxide is MO_2 , and the equation is $\text{M} + \text{O}_2 \longrightarrow \text{MO}_2$.

From gas volumes. Equal volumes of different gases, measured at the same temperature and pressure, contain the same number of molecules (Avogadro's law). So the volumes of gases in a reaction are in the ratio of the *coefficients* in the equation, with no need to convert to moles.



For example, 20 cm^3 of nitrogen reacts with 60 cm^3 of hydrogen to give 40 cm^3 of ammonia: $20 : 60 : 40 = 1 : 3 : 2$, so $\text{N}_2 + 3\text{H}_2 \longrightarrow 2\text{NH}_3$. This works only for gases, all measured under the same conditions; solids and liquids do not follow the volume rule.

Worked example: equation from volumes

10 cm^3 of a gas X_2 reacts with 30 cm^3 of chlorine gas, Cl_2 , to form 20 cm^3 of a gaseous product, all at the same temperature and pressure. Write the equation.

Ratio $10 : 30 : 20 = 1 : 3 : 2$. Atoms: $\text{X}_2 + 3\text{Cl}_2 \longrightarrow 2\text{XCl}_3$ (six Cl atoms and two X atoms on each side; the product is XCl_3).

Worked example: ratio of moles from masses

When 2.7 g of aluminium ($A_r = 27$) reacts completely with chlorine, 13.35 g of aluminium chloride forms. Find the formula of aluminium chloride.

Mass of chlorine = $13.35 - 2.7 = 10.65$ g. Al: $2.7/27 = 0.10$ mol. Cl atoms: $10.65/35.5 = 0.30$ mol. Ratio 1 : 3, so AlCl_3 .

Common mistakes

- Using gas-volume ratios for solids or liquids. Only gases (at the same T and P) follow the volume rule.
- Forgetting that gases such as H_2 , N_2 , O_2 , Cl_2 are diatomic: 1 volume of Cl_2 supplies two chlorine atoms per molecule.
- Stopping at the ratio of moles of *atoms* without checking the equation balances.

Check yourself

Try these without looking back, then compare with the answers on the next page.

1. Find M_r of $\text{Ca}(\text{OH})_2$.
2. How many moles are in 88 g of carbon dioxide?
3. How many molecules are in 0.5 mol of oxygen gas?
4. What is the percentage of carbon in CO_2 ?
5. A compound has 1.2 g C and 0.2 g H. Find its empirical formula.
6. What mass of MgO is made from 12 g of Mg? ($2\text{Mg} + \text{O}_2 \longrightarrow 2\text{MgO}$)
7. In $\text{Mg} + 2\text{HCl} \longrightarrow \text{MgCl}_2 + \text{H}_2$, 3 mol Mg is added to 4 mol HCl. Which is limiting?
8. 50 cm^3 of H_2 reacts with 25 cm^3 of O_2 to form steam. What is the ratio of the volumes?

Answers to the self-check

1. $40 + 2 \times (16 + 1) = 74$.
2. $88/44 = 2 \text{ mol}$.
3. $0.5 \times 6.0 \times 10^{23} = 3.0 \times 10^{23}$.
4. $12/44 = 27.3\%$ (about 27%).
5. C: 0.1; H: 0.2; ratio 1 : 2: CH_2 .
6. $12/24 = 0.5 \text{ mol Mg} \rightarrow 0.5 \text{ mol MgO} = 0.5 \times 40 = 20 \text{ g}$.
7. HCl: $4/2 = 2$ is less than 3. HCl is limiting.
8. $50 : 25 = 2 : 1$, so $2 \text{ H}_2 + \text{O}_2 \longrightarrow 2 \text{ H}_2\text{O}$.

Next step

Heurisco's *ESAT Chemistry: Moles and Reacting Masses Practice Set* has twenty original exam-style questions on this lesson, with graphs and particle diagrams, in order of difficulty, with a full worked solution and a note on the trap in each.