Chapter 1 Working scientifically

1.1 Questioning and predicting

1.1 KEY QUESTIONS

1 A hypothesis is a statement that can be tested. This involves making a prediction based on previous knowledge and evidence or observations.

2 A. ‘How are chemicals in solutions measured?’ is an inquiry question.

3
- **a** electrical conductivity
- **b** concentration of lead
- **c** electrical conductivity
- **d** pH

4 qualitative observation

5 C. A calibrated pH meter at a particular temperature provides the most accurate, quantitative data. Litmus paper and universal indicator provide a qualitative indication of pH. A conductivity meter measures conductivity, not pH.

6 C. A hypothesis should test only one independent variable, and it should predict the relationship between the independent and dependent variables. Hypothesis A tests two independent variables. Hypothesis B does not predict the type of relationship between the independent and dependent variables.

1.2 Planning investigations

1.2 KEY QUESTIONS

1 B

2
- **a** In a controlled experiment, two groups of subjects are tested; the groups, or the tests performed on them, are identical except for a single factor (the independent variable).
- **b** The dependent variable is the variable that is measured to determine the effect of changes in the independent variable. The independent (experimental) variable is the variable that is changed in an experiment. For example, in an experiment testing the effect of soil pH on flower colour, the independent variable would be soil pH and the dependent variable would be flower colour.

3
- **a** type of soft drink
- **b** pH
- **c** temperature of solutions, type of equipment used, storage and preparation of solutions.

4
- **a** Litmus paper and universal indicator give qualitative information about pH (e.g. acidic, basic, or neutral) through colour, and therefore is not an accurate method for determining pH.
- **b** A calibrated pH meter will give quantitative information and is more accurate than using litmus paper or universal indicator.

5
- **a** valid
- **b** reliable
- **c** accurate

1.3 Conducting investigations

1.3 KEY QUESTIONS

1 Data set A: mistake, because there is one unexpected value (1.5) in the data.

Data set B: systematic error, because the error is not obvious and may be due to a consistent equipment or operator error.

2 **a** systematic errors

b random errors

3 17.34 mL, 17.38 mL and 17.44 mL

4 **a** systematic

b mistake

c random

5 There could be many reasons why the same experimental results cannot be obtained. The experimental design may be poor because of a lack of objectivity, clear and simple instructions and appropriate equipment, or a failure to control variables.

Other problems not specifically related to the experiment could be a poor hypothesis that could not be tested objectively, conclusions that do not agree with the results and interpretations that are subjective.

1.4 Processing data and information

Worked example: Try yourself 1.4.1

CALCULATING UNCERTAINTY

The temperature of the water in the Murrumbidgee River at Wagga Wagga was measured in February 2017. The temperatures (in °C) recorded were:

23, 22, 24, 21, 26, 24 and 23

Find the uncertainty for these values.

Thinking	**Working**
Calculate the average temperature.	average = $\frac{23+22+24+21+26+24+23}{7} = 23°\text{C}$
Calculate the maximum variance from the mean.	26 is 3 above the average, so the uncertainty is 3.
Write the average temperature and include the uncertainty.	Average temperature is 23 ± 3°C

1.4 KEY QUESTIONS

1 **a** mean = $\frac{21+28+19+19+25+24}{6} = 23$

b mode = 19

c median = 23

2 Add a trend line or line of best fit.

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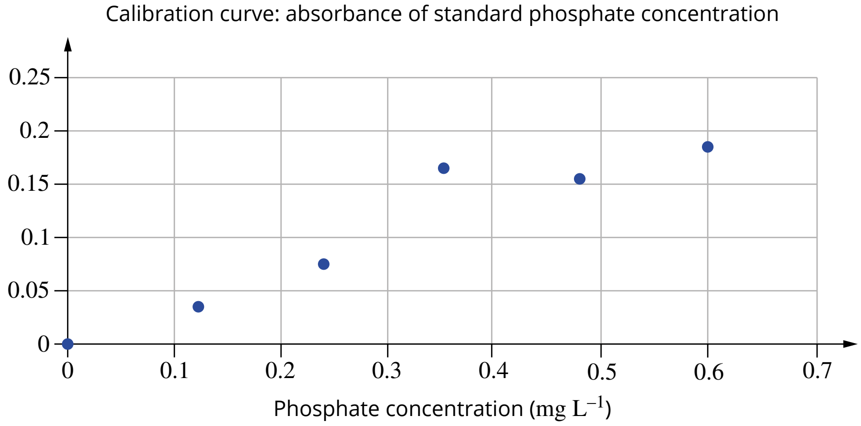
3 **a**



b Data point 4 (0.36, 0.159) is an outlier.

c An outlier is a point in the data that does not fit the trend and may be the result of error.

4 The mean is 23 with an uncertainty of 5. The mean is $\frac{21+28+19+19+25+24}{6} = 23$. The uncertainty is the maximum variance from the mean, 28 – 23 = 5.

1.5 Analysing data and information

1.5 KEY QUESTIONS

1 A sloping linear graph shows a directly proportional relationship between two variables.
2 inversely proportional relationship
3 directly proportional
4 time restraints and limited resources
5 An increase in the amount of limescale on the heating element of the kettle from 10% to 20% produced a 14% decrease in the efficiency of the kettle.

1.6 Problem solving

1.6 KEY QUESTIONS

1 B. A scientific article should always be written using objective language. A concluding paragraph must summarise the information presented in the article and connect it with the title. It should also include limitations, possible applications of the research and potential future research.

2 Implications. Generalisations apply facts to a broader context that they have not been specifically tested for. Implications use facts and accepted knowledge to logically make a connection between two things, such as the results of a chemical experiment and a possible commercial use of that chemical.

3 The statement ‘Many repeats of the procedure were conducted’ is unquantified. ‘Thirty repeats of the procedure were conducted’ is better because the number of trials is quantified.

4 A. A hypothesis can never be proved to be correct; it can be supported or refuted by the results of the investigation. Options B and C are vague. A strong conclusion will concisely summarise the results of the investigation and specifically link them to the hypothesis, including relevant chemistry where appropriate.

1.7 Communicating

1.7 KEY QUESTIONS

1 B. Scientific writing should not use biased or absolute language.
2 D. First-person narrative uses the pronoun ‘I’.
3 **a** $\mathrm{g\,mol^{-1}}$
b $\mathrm{J\,g^{-1}K^{-1}}$
c $\mathrm{mol\,L^{-1}}$

4 Divide the value in g by 1000.

5 This might need to be done to reflect how we use quantities of varying units of measurement in the laboratory or to make the numbers easier to comprehend.

CHAPTER 1 REVIEW

1 **a** concentration of lead—dependent variable
  b analytical technique, temperature of water sample, type of sampling container—controlled variables
  c source and location of water—independent variable

2 **a** reflect
  b create
  c analyse
  d investigate
  e apply
  f identify
  g describe

3 independent variable: source of the water; dependent variable: phosphate concentration; controlled variables: temperature, time of testing, method of testing

4 **a** bar graph
  b line graph
  c scatter graph (with line of best fit)
  d pie chart

5 **a** It can cause severe burns, and dissolve or eat away at substances, including tissues such as your skin or lungs.
  b It is poisonous if inhaled or ingested.
  c It is a highly combustible liquid that could catch fire.

6 Accuracy refers to the ability of the method to allow the measurement to be close to a true or accepted value. Validity refers to whether an experiment or investigation enables testing of the set hypothesis and achievement of the purpose.

7 The uncertainty of a set of data is the maximum variance from the mean of the data. In this example, the mean is 6.63. The maximum variance is 7.20 – 6.63 = 0.57 = the uncertainty.

8 Mean. The mode is the most common value in a set of data. The median is the middle value. The mean is the average of the values in a data set. The inclusion of an outlier will have the most effect on the mean value; therefore, outliers are excluded from calculating the mean value.

9 **a** mistake
  b random error
  c systematic error

10 a non-linear relationship, e.g. exponential, inverse

11 **a** reliability
  b validity
  c accuracy
  d precision

12 Evaluate the method; identify issues that could affect validity, accuracy, precision and reliability of data; state systematic sources of error and uncertainty; and recommend improvements to the investigation if it is to be repeated.

13 A trend is a pattern or relationship that can be seen between the dependent and independent variables. It may be linear, in which case the variables change in direct proportion to each other to produce a straight trend line. The relationship may be in indirect proportion to each other and non-linear, giving a curved trend line. The relationship may also be inverse, when one variable decreases in response to the other variable increasing. This relationship can also be linear or non-linear.

14 Limitations are issues that could have affected the validity, accuracy, precision or reliability of the data, plus any sources of error or uncertainty.

15 Bias is a form of systematic error resulting from a researcher’s personal preferences or motivations.

16 The purpose of referencing and acknowledgements is to ensure creators and sources are properly credited for their work, and to demonstrate that prior research has been undertaken to support the validity of the hypothesis.

17 C. An interactive periodic table, an article published in a science magazine, a science documentary and this Year 11 textbook contain summaries of information published elsewhere, for example, in scientific journals. A practical report written by a Year 11 student and an article published in a peer-reviewed scientific journal contain primary sources of data.

18 A. In APA style, the in-text citation should list the first author's last name and the year of publication.

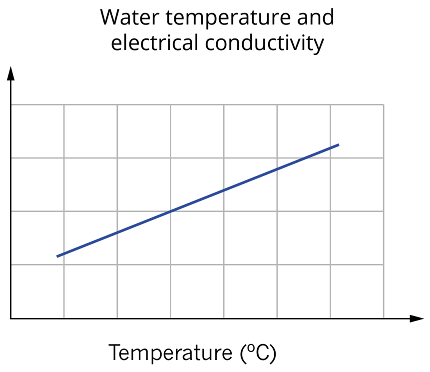
19 0.03000 L or 3.000×10^{-2} L

20 **a** Purpose: to determine the effect of increasing water temperature on the electrical conductivity of water.

b independent variable: water temperature; dependent variable: electrical conductivity of water; controlled variables: pH, water source, type of sampling container

c The data collected would be electrical conductivity, measured using a probe, and therefore it would be quantitative.

d Raw data is data collected in the field and recorded as measurements are taken. Processed data is tabulated in a form in which the reader can clearly see the temperature of the water and the conductivity at each separate temperature value. This can then be processed and graphed with the independent variable on the *x*-axis (temperature) and dependent variable on the *y*-axis (conductivity). If the hypothesis is correct, the graph for this experiment could look like this:



A non-linear increase could also fit the hypothesis, because a directly proportional relationship between conductivity and temperature has not been specified.

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Chapter 2 Properties of matter

2.1 Types of matter

2.1 KEY QUESTIONS

1 An element is made up of just one type of atom. A compound is a substance that contains at least two different types of atoms, bonded together in a definite proportion.

2 **a** element
  b element
  c compound
  d compound
  e element
  f compound
  g element
  h compound

3 **a** solvent
  b solution
  c solute
  d solute

4 They all have water as the solvent.

5 **a** The batter has different compositions depending on where it is sampled. The batter is a heterogeneous mixture.
  b The vigorous mixing would make the batter a homogeneous mixture. A sample of batter taken from the middle, left or right of the bowl should now taste the same.

6 The gases are combining physically to form a mixture. The component gases retain their individual character. To form water, hydrogen needs to react with oxygen and combine chemically, not just mix physically.

7 A compound has two or more different types of atoms in a definite proportion, and any sample of a homogeneous mixture (wherever it is sampled from) will have the same combination of materials in the same proportion. In this case, you cannot tell if it is a chemically combined compound or a physically combined mixture. You can only tell that it is neither a heterogeneous mixture nor an element.

2.2 Physical properties and changes of state

2.2 KEY QUESTIONS

1 The particles are (1) hard, (2) indivisible, (3) attracted to each other and (4) constantly in motion.

2 liquid and gas

3 solid

4 1 kg of steam. A change of state does not create or destroy mass.

5 The particle model assumes particles are attracted to one another. The particles in solids and liquids are in contact and there is limited space between the particles to allow them to come closer together. The particles in a gas are not in contact, and there is space that allows the particles to come closer together.

6 The hot day means the temperature is higher, and the particles of water in the liquid move more rapidly and have enough energy to overcome the forces of attraction between particles to form water vapour. As the wind blows, the particles in the water vapour are blown away from the puddle, increasing the rate of evaporation. Eventually the puddle will dry out.

2.3 Separating mixtures

2.3 KEY QUESTIONS

1 A. All the others are methods of filtration.

2 by (fractional) distillation

3 A magnet. Neither iron nor sand is a volatile material.

4 Heating the pot containing the solution will evaporate the water. Eventually the magnesium sulfate will precipitate from solution. When cool, the magnesium sulfate crystals can be filtered from the solution.

5 The sodium phosphate is soluble and remains dissolved in the unevaporated water. Recovering all the material from solution is difficult.

2.4 Calculating percentage composition

Worked example: Try yourself 2.4.1

CALCULATING THE MASS OF ELEMENTS IN A COMPOUND, BASED ON PERCENTAGE COMPOSITION BY MASS

An 85 g sample of sodium nitrate was analysed and found to contain 16% nitrogen and 56% oxygen, and the remainder was sodium. Determine the mass of each element in the compound.

Thinking	**Working**
Calculate the mass of each element in the compound using the formula: mass of element = $\frac{\text{element}\% \times \text{mass of compound}}{100}$	Mass of nitrogen: mass of nitrogen = $\frac{\text{nitrogen}\% \times \text{mass of sodium nitrate}}{100}$ $= \frac{16 \times 85}{100}$ = 14 g Mass of oxygen: mass of oxygen = $\frac{\text{oxygen}\% \times \text{mass of sodium nitrate}}{100}$ $= \frac{56 \times 85}{100}$ = 48 g
The mass of sodium can be determined by subtracting the mass of nitrogen and oxygen from the mass of sodium nitrate.	Mass of sodium: mass of sodium = mass of sodium nitrate − mass of nitrogen − mass of oxygen = 85 − 14 − 48 = 23 g
State the masses of each element in the sample.	The sample is made up of 14 g of nitrogen, 48 g of oxygen and 23 g of sodium.

Worked example: Try yourself 2.4.2

CALCULATING THE PERCENTAGE COMPOSITION BY MASS OF EACH ELEMENT IN A MIXTURE

A 75 g glucose solution contains 5% glucose (40% C, 7% H, 53% O) and 95% water (11% H, 89% O). Determine the percentage composition by mass of each element in this mixture.

Thinking	Working
Calculate the mass of each component in the mixture using the formula: mass of component $= \frac{\text{component\%} \times \text{mass of mixture}}{100}$	Mass of glucose: mass of glucose $= \frac{\text{glucose\%} \times \text{mass of mixture}}{100}$ $= \frac{5 \times 75}{100}$ $= 4\text{ g}$ Mass of water: mass of water $= \frac{\text{water\%} \times \text{mass of mixture}}{100}$ $= \frac{95 \times 75}{100}$ $= 71\text{ g}$
Calculate the mass of carbon and hydrogen in glucose using the formula: mass of element $= \frac{\text{element\%} \times \text{mass of compound}}{100}$	Mass of carbon: massof carbon $= \frac{\text{carbon\%} \times \text{mass of glucose}}{100}$ $= \frac{40 \times 4}{100}$ $= 1.6\text{ g}$ Mass of hydrogen: mass of hydrogen $= \frac{\text{hydrogen\%} \times \text{mass of glucose}}{100}$ $= \frac{7 \times 4}{100}$ $= 0.3\text{ g}$
Calculate the mass of hydrogen in water using the formula: mass of element $= \frac{\text{element\%} \times \text{mass of compound}}{100}$	Mass of hydrogen: mass of hydrogen $= \frac{\text{hydrogen\%} \times \text{mass of water}}{100}$ $= \frac{11 \times 71}{100}$ $= 7.8\text{ g}$
Calculate the percentage by mass of carbon in the mixture.	percentage by mass of carbon $= \frac{\text{mass of carbon}}{\text{mass of mixture}} \times 100$ $= \frac{1.6}{75} \times 100$ $= 2.1\%$
Calculate the percentage by mass of hydrogen in the mixture.	percentage by mass of hydrogen $= \frac{\text{mass of hydrogen in glucose} + \text{mass of hydrogen in water}}{\text{mass of mixture}} \times 100$ $= \frac{0.3 + 7.8}{75} \times 100$ $= 10.8\%$
The remaining balance must be oxygen.	oxygen% = 100 − carbon% − hydrogen% = 100 − 2.1 − 10.8 = 87.1%
State the percentage composition by mass of each element in the sample.	The sample is made up of 2.1% carbon, 10.8% hydrogen and 87.1% oxygen.

2.4 KEY QUESTIONS

1 Carbon is $\frac{24}{88} \times 100 = 27\%$, oxygen is $\frac{64}{88} \times 100 = 73\%$.

2 Nitrogen is $\frac{28}{106} \times 100 = 26\%$, chlorine is $\frac{71}{106} \times 100 = 67\%$. The remainder must be from hydrogen, and the balance of mass is $100 - (26 + 67) = 7\%$.

3 No, because the total returned mass is 63 g, more than the original mass of the sample. There is an error: either the sample was contaminated during analysis or the original mass is incorrect.

4 Carbon dioxide is $\frac{10}{13} \times 100 = 77\%$. The remainder is oxygen: $100 - 77 = 23\%$.

5 Gold: $\frac{75}{100} \times 20 = 15\,\text{g}$

Copper: $\frac{20}{100} \times 20 = 4\,\text{g}$

Silver: $\frac{5}{100} \times 20 = 1\,\text{g}$

6 **a** From carbon monoxide, the oxygen percentage is $100 - 43 = 57\%$. In a 5 g sample, oxygen is $5 \times \frac{57}{100} = 3\,\text{g}$. From nitrous oxide, the oxygen percentage is $100 - 30 = 70\%$. In a 5 g sample, oxygen is $5 \times \frac{70}{100} = 4\,\text{g}$. The total mass of oxygen is $3 + 4 = 7\,\text{g}$.

b From 5 g of carbon monoxide, the mass of oxygen is $(100\% - 43\%) \times 5 = 2.9\,\text{g}$. From 4 g of nitrous oxide, the mass of oxygen is $(100\% - 30\%) \times 4 = 2.8\,\text{g}$. From 3 g of sulfur oxide, the mass of oxygen is $(100\% - 50\%) \times 3 = 1.5\,\text{g}$. The total mass of oxygen is $2.9 + 2.8 + 1.5 = 7.2\,\text{g}$. The total mass of the sample is $5 + 4 + 3 = 12\,\text{g}$, so the percentage by mass of oxygen is $\frac{7.2}{12} \times 100 = 60\%$. The mass of sulfur in sulfur dioxide is $\frac{50}{100} \times 3 = 1.5\,\text{g}$, so the percentage by mass of sulfur in the mixture is $\frac{1.5}{12} \times 100 = 13\%$.

7 In a 100 g sample, nitrogen is $100 \times \frac{10}{100} = 10\,\text{g}$. Nitrogen is 30% of nitrogen dioxide by mass. The mass of nitrogen dioxide will be $10 \times \frac{100}{30} = 33\,\text{g}$. The balance of the mixture, carbon monoxide, is $100 - 33.3 = 67\,\text{g}$. Carbon monoxide is 43% carbon by mass. The mass of carbon is $67 \times \frac{43}{100} = 29\,\text{g}$. The percentage of carbon by mass is $\frac{29}{100} \times 100 = 29\%$.

2.5 Elements and the periodic table

2.5 KEY QUESTIONS

1 **a** iron
  b potassium
  c tungsten
  d lead

2 Malleable. The property of being ductile is similar, but refers to being able to be drawn into a wire.

3 Oxygen. Oxygen is a non-metal. Rhodium and calcium are metals, so are good conductors of heat and electricity. Silicon is a metalloid, so will also conduct heat and electricity.

4 The unreactivity of copper with water is a chemical property. The colour of the element and the melting point are physical properties. An alloy is a physical mixture and not a chemical property.

5 The properties of being a good thermal conductor and having a high melting point are useful. Cookware needs to transfer heat from the stove or oven to the food, but not melt itself. Materials that are a good conductor of electricity and are ductile are useful for making wires.

6 Sulfur then silicon then scandium. Sulfur is a non-metal; silicon, a metalloid; and scandium, a metal. A metal is a good conductor of electricity, so it is a poor insulator; a non-metal is a poor conductor of electricity, so it is a good insulator.

CHAPTER 2 REVIEW

1 Compounds are made of two or more different types of atom in fixed ratios.

2 Atoms are hard, indivisible structures.

3 Atoms are the basic building blocks of matter. Molecules are composed of a number of atoms chemically bonded together.

4 A discrete molecule has a definite number of atoms chemically bonded together. A network has an indefinite number of atoms chemically bonded together.

5 monatomic: helium, neon, krypton

molecules: sulfur, oxygen, nitrogen

large network: copper, gold, carbon, tin

6 A compound consists of two or more elements chemically bonded together in a fixed ratio. A mixture has two or more pure substances physically combined in no fixed ratio.

7 The milk with cream is a mixture that differs depending on whether it is sampled from the top or with the bottom. It is a heterogeneous mixture.

8 In general, no. Dalton defined matter as being composed of indivisible particles. Particle theory views matter as particles in motion, but does not assume that the particles are atoms. In Dalton's atomic theory, an atom of sulfur is fundamental. In particle theory, the ring of eight sulfur atoms is the fundamental unit.

9 More dense. The balloon when blown up is taut and squeezes the air inside slightly. The air inside is slightly more pressured, and has a greater mass of air in the volume of the balloon compared with the volume of unpressurised air.

10 This is a physical change. The particles in the metal are rearranged and ordered differently, but the particles themselves have not changed. If the particles had changed, this would be a chemical change.

11 **a** The temperature stays the same during the transition.

b The attractive forces change—they are being formed as the temperature decreases (condensing, freezing), and being broken as the temperature increases (melting, vaporising).

12 Toothpaste keeps its shape like a solid, but when squeezed it flows like a liquid.

13 No. The ink solution is a homogeneous mixture, so filtration is not the best choice.

14 Most filters are too coarse and the blood cells would remain in the filtrate. A very fine filter would retain the blood cells, but it would be a slow process. Techniques such as centrifugation are simpler.

15 The puddle will evaporate more quickly. Evaporation takes place on the surface of the water. The cup of water has a much smaller surface exposed to air than the puddle.

16 The liquid with the lower boiling point is more volatile than the other liquid. Therefore, the vapour will contain a greater proportion of the liquid with the lower boiling point.

17 Na 31% + O 21% means the balance, 48%, must be Cl.

18 If 48% of the original mass is 3.0 g, the original mass must be $3.0 \times \frac{100}{48} = 6.3\,\text{g}$.

19 Fe: $\frac{111}{150} \times 100 = 74\%$

Cr: $\frac{27}{150} \times 100 = 18\%$

Ni: $100 - 74 - 18 = 8\%$

20 mass of calcium phosphate $= \frac{12}{100} \times 20 = 2.4\,\text{g}$

21 Assume there is 100 g of methylated spirits = 10 g methanol and 90 g ethanol.

mass of carbon in ethanol $= \frac{52}{100} \times 10 = 5.2\,\text{g}$

mass of carbon in methanol $= \frac{38}{100} \times 90 = 34.2\,\text{g}$

total mass of carbon in methylated spirits $= 5.2 + 34.2 = 39.4\,\text{g}$

percentage composition by mass of carbon in methylated spirits $= \frac{39.4}{100} \times 100 = 39.4 = 39\%$

22 **a** O

b C

c He

d Ni

e K

f Au

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23 **a** nitrogen
   b calcium
   c chlorine
   d silver
   e mercury

24 **a** A metalloid. It is shiny but a poor conductor of heat, so it is not a metal.
   b the element that conducts electricity

25 Iron is a metal and is malleable. A fine, sharp edge can be formed by flattening the metal. Sulfur is a non-metal and not malleable. When hammered, the sulfur becomes a powder.

26 They are all non-metals, so they are all likely to be poor conductors. The other properties are properties of metals.

27 Properties **a**, **b** and **d** are physical properties, because the diamond, sugar and aluminium have not changed. Properties **c** and **e** are chemical properties, because the carbon in the wood burns and changes into carbon dioxide, and the lithium reacts with water and changes into lithium hydroxide.

28 **a** Lithium and sodium are positioned in the same group, so their chemical properties are likely to be similar. Magnesium is in a different group so its chemical properties will differ. All of these elements are metals, so their physical properties are likely to be similar, in that all are malleable, ductile, good conductors etc.

   b These elements are, respectively, a metal, a metalloid and a non-metal. You would not expect their physical properties to be similar.

   c From group 1, potassium, to group 13, gallium, the elements are metals and likely to be good conductors of heat. In group 14, germanium, and group 15, arsenic, the elements are metalloids and moderate conductors of heat. From group 16 onwards the non-metals are expected to be poor conductors of heat.

29 The crust is a heterogeneous mixture of minerals and rocks, and the rocks are in turn mixtures of minerals. The first problem was in separating the mixtures to generate a sufficient quantity of pure aluminium-bearing minerals from the soil and other minerals. The main problem in isolating metallic aluminium from its ores was that it is a very reactive metal and therefore hard to separate chemically from the other elements in its compounds.

For modern industry, bauxite (a mix of $Al(OH)_3$ and $AlO(OH)$) is the most common aluminium-bearing mineral. Through the Bayer process, the bauxite is separated from other minerals and purified to alumina (Al_2O_3). Molten alumina is converted to elemental aluminium by electrolysis. This means the large-scale production of aluminium was only possible from the late 18th century onwards.

30 The mixture is a gas. Filtration is inappropriate as there are no solids. Evaporation is inappropriate as there are no liquids. However, on cooling, the gases will become liquids. Oxygen boils at −183°C (90 K), argon at −186°C (87 K), and nitrogen at −196°C (77 K). Cooling the mixture to less than the boiling point of the least volatile compound (nitrogen) means that all of the components will become liquids. Distillation is now possible. Allowing temperatures to increase slightly, first nitrogen will evaporate, then argon, leaving behind liquid oxygen.

Argon freezes at −189°C (84 K), nitrogen at −210°C (63 K) and oxygen at −219°C (54 K). In principle, the frozen solids could be collected and filtered, but with much more difficulty than for simple distillation.

In industrial practice, this fractional distillation process is actually used, but the precise boiling points of the gas mixtures shift a little when the liquid mixture is rich in oxygen.

31 Malleability. This property supports aluminium being a metal, because metals can be beaten into different shapes without breaking or shattering.

Chapter 3 Atomic structure and atomic mass

3.1 Inside atoms

3.1 KEY QUESTIONS

1 10 000–100 000 times larger

2 protons and neutrons, found in the nucleus

3 The electrons are held within the cloud surrounding the nucleus by the electrostatic attraction of the protons; the negative electrons are attracted to the positive protons and pulled towards them.

4 **a** −1
  b +1
  c 0

3.2 Classifying atoms

Worked example: Try yourself 3.2.1

CALCULATING THE NUMBER OF SUBATOMIC PARTICLES

Calculate the number of protons, neutrons and electrons for the atom with this atomic symbol: ${}^{235}_{92}\mathrm{U}$

Thinking	**Working**
The atomic number is equal to the number of protons.	number of protons = Z = 92
Find the number of neutrons. number of neutrons = mass number – atomic number	number of neutrons = $A - Z$ = 235 – 92 = 143
Find the number of electrons. The number of electrons is equal to the atomic number because the total negative change is equal to the total positive charge.	number of electrons = Z = 92

3.2 KEY QUESTIONS

1 mass number

2 Z = 15, so protons = 15, electrons = 15. Neutrons = 31 – 15 = 16.

3 Its atomic number is 7 because it has 7 protons. The element that contains 7 protons is nitrogen.

4 Isotopes of the same element have the same atomic number and therefore the same number of protons and electrons. The number of neutrons differs between isotopes of the same element; therefore, they have different mass numbers.

5 Carbon-14 has a mass number A = 14. The number of neutrons is:
  mass number – atomic number = 14 – 6 = 8
  Carbon-12 has a mass number of 12 and therefore has 6 neutrons.
  Therefore, the difference in the number of neutrons is 8 – 6, which is 2 neutrons.

6 Alpha particles are made up of two protons and two neutrons (a helium nucleus). Beta particles are high-energy electrons. Gamma radiation is a type of high-energy electromagnetic radiation.

7 **a** ${}^{234}_{90}\mathrm{Th} \rightarrow {}^{234}_{91}\mathrm{Pa} + {}^{0}_{-1}\mathrm{e}$
  b ${}^{234}_{91}\mathrm{Pa} \rightarrow {}^{230}_{89}\mathrm{Ac} + {}^{4}_{2}\mathrm{He}$
  c ${}^{220}_{86}\mathrm{Rn} \rightarrow {}^{216}_{84}\mathrm{Po} + {}^{4}_{2}\mathrm{He}$
  d ${}^{216}_{84}\mathrm{Po} \rightarrow {}^{216}_{85}\mathrm{At} + {}^{0}_{-1}\mathrm{e}$

3.3 Masses of particles

Worked example: Try yourself 3.3.1

CALCULATING RELATIVE ATOMIC MASS FROM ISOTOPIC MASSES AND PERCENTAGE ABUNDANCES

Boron has two isotopes. Their relative isotopic masses and percentage abundances are provided. Calculate the relative atomic mass of boron correct to 2 decimal places.

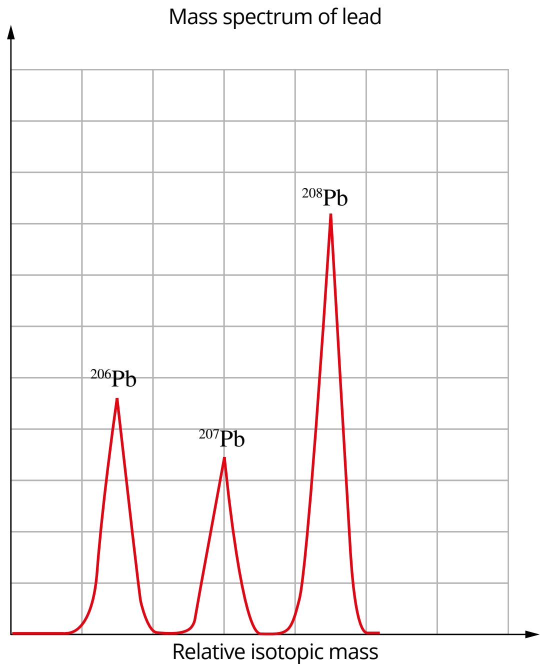
Isotope	I_r	**Relative abundance (%)**
^{10}B	10.013	19.91
^{11}B	11.009	80.09

Thinking	**Working**
Determine the relative isotopic masses (I_r) and abundances for each isotope.	First isotope: I_r 10.013; abundance 19.91% Second isotope: I_r 11.009; abundance 80.09%
Substitute the relative isotopic masses and abundance into the formula for calculating relative atomic mass: $A_r = \frac{(I_r \times \% \text{ abundance}) + (I_r \times \% \text{ abundance})}{100}$	$A_r = \frac{(10.013 \times 19.91) + (11.009 \times 80.09)}{100}$
Calculate the relative atomic mass.	$A_r = \frac{199.36 + 881.71}{100} = 10.811$
Express the answer to 2 decimal places.	$A_r(\text{B}) = 10.81$

Worked example: Try yourself 3.3.2

CALCULATING PERCENTAGE ABUNDANCE OF EACH ISOTOPE FROM THE MASS SPECTRUM

The graph shows a simplified mass spectrum of lead. From this mass spectrum, calculate the percentage abundances of each of its three isotopes.



Thinking	**Working**
Measure the peak height for each isotope using a ruler.	From the spectrum, the height of each peak is: $^{206}Pb = 3.1\ cm$ $^{207}Pb = 2.4\ cm$ $^{208}Pb = 5.6\ cm$
Calculate the total peak height for the three isotopes by adding the individual peak heights.	total peak height = $3.1 + 2.4 + 5.6 = 11.1\ cm$
Substitute the peak height for each isotope into the formula: $\%\ abundance = \frac{peak\ height}{total\ peak\ height} \times 100$	$\%\ abundance\ ^{206}Pb = \frac{3.1}{11.1} \times 100 = 28\%$ $\%\ abundance\ ^{207}Pb = \frac{2.4}{11.1} \times 100 = 22\%$ $\%\ abundance\ ^{208}Pb = \frac{5.6}{11.1} \times 100 = 50\%$

Worked example: Try yourself 3.3.3

CALCULATING PERCENTAGE ABUNDANCES FROM RELATIVE ATOMIC MASS AND RELATIVE ISOTOPIC MASSES

The relative atomic mass of copper is 63.54. The relative isotopic masses of its two isotopes are 62.95 and 64.95. Calculate the relative abundances of the isotopes in naturally occurring copper.

Thinking	**Working**
State the relative abundances of the two isotopes in terms of *x*, where *x* is the abundance of the lighter isotope. Abundance of lighter isotope = *x*. The abundance of the heavier isotope must equal 100 – *x*.	abundance of 62.95 isotope = *x* abundance of 64.95 isotope = 100 – *x*
Substitute the relative isotopic masses, relative abundances and relative atomic mass into the formula: $A_r = \frac{(I_r \times \%\ abundance) + (I_r \times \%\ abundance)}{100}$	$63.54 = \frac{62.95x + (64.95(100 - x))}{100}$
Expand the top line of the equation.	$63.54 = \frac{62.95x + (6495 - 64.95x)}{100}$
Solve the equation to find *x*, the relative abundance of the lightest isotope.	$6354 = 62.95x + 6495 - 64.95x$ $6354 - 6495 = 62.95x - 64.95x$ $-141 = -2x$ $x = 70.50\%$
Determine the abundance of the heavier isotope.	Abundance of 64.95 isotope $= 100 - x$ $= 100 - 70.50$ $= 29.50\%$

Worked example: Try yourself 3.3.4

CALCULATING THE RELATIVE MOLECULAR MASS OF MOLECULES

Calculate the relative molecular mass of nitric acid (HNO_3).

Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(H) = 1.008$ $A_r(N) = 14.01$ $A_r(O) = 16.00$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × H atom 1 × N atom 3 × O atoms
Determine the relative molecular mass by adding the appropriate relative atomic masses.	$M_r = 1 \times A_r(H) + 1 \times A_r(N) + 3 \times A_r(O)$ $= 1 \times 1.008 + 1 \times 14.01 + 3 \times 16.00$ $= 63.02$

Worked example: Try yourself 3.3.5

CALCULATING THE RELATIVE FORMULA MASS OF IONIC COMPOUNDS

Calculate the relative formula mass of copper(II) nitrate ($Cu(NO_3)_2$).

Thinking	**Working**
Use the periodic table to find the relative atomic mass for the elements represented in the formula.	$A_r(Cu) = 63.55$ $A_r(N) = 14.01$ $A_r(O) = 16.00$
Determine the number of atoms of each element present, taking into consideration any brackets in the formula.	1 × Cu atom 1 × 2 = 2 N atoms 3 × 2 = 6 O atoms
Determine the relative formula mass by adding the appropriate relative atomic masses.	Relative formula mass $= 1 \times A_r(Cu) + 2 \times A_r(N) + 6 \times A_r(O)$ $= 63.55 + 2 \times 14.01 + 6 \times 16.00$ $= 187.57$

3.3 KEY QUESTIONS

1 B.

A is incorrect because the relative atomic mass is a weighted average of the masses of the two isotopes, taking into account their abundances.

C is incorrect because the lighter isotope, ^{35}Cl, is more abundant than the heavier isotope, ^{37}Cl.

D is incorrect because isotopes have the same number of protons and different numbers of neutrons.

2 **a** $A_r(O) = \frac{(15.995 \times 99.76) + (16.999 \times 0.04) + (17.999 \times 0.20)}{100} = 15.999$

b $A_r(Ag) = \frac{(106.9 \times 51.8) + (108.9 \times 48.2)}{100} = 107.9$

c $A_r(H) = \frac{(1.008 \times 99.986) + (2.014 \times 0.014) + (3.016 \times 0.0001)}{100} = 1.008$

3 Let the percentage abundance of the lighter isotope be x.

∴ percentage abundance of the heavier isotope will be $100 - x$

$6.94 = \frac{6.02x + (7.02 \times (100 - x))}{100}$

$694 = 6.02x + 702 - 7.02x$

$x = 8\%$

Percentage abundance of the lighter isotope is 8%.

4 **a** $\%\ \text{abundance}\ {}^{90}\text{Zr} = \frac{\text{peak height}}{\text{total peak height}} \times 100 = 51\%$

$\%\ \text{abundance}\ {}^{91}\text{Zr} = \frac{\text{peak height}}{\text{total peak height}} \times 100 = 11\%$

$\%\ \text{abundance}\ {}^{92}\text{Zr} = \frac{\text{peak height}}{\text{total peak height}} \times 100 = 17\%$

$\%\ \text{abundance}\ {}^{94}\text{Zr} = \frac{\text{peak height}}{\text{total peak height}} \times 100 = 17\%$

$\%\ \text{abundance}\ {}^{96}\text{Zr} = \frac{\text{peak height}}{\text{total peak height}} \times 100 = 4\%$

b $A_r(\text{Zr}) = \frac{(90 \times 51) + (91 \times 11) + (92 \times 17) + (94 \times 17) + (96 \times 4)}{100} = 91$

5 **a** $M_r = 2 \times A_r(\text{H}) + A_r(\text{S}) + 4 \times A_r(\text{O})$
$= 2 \times 1.008 + 32.07 + 4 \times 16.00$
$= 98.09$

b $M_r = A_r(\text{N}) + 3 \times A_r(\text{H})$
$= 14.01 + 3 \times 1.008$
$= 17.03$

c $M_r = 2 \times A_r(\text{C}) + 6 \times A_r(\text{H})$
$= 2 \times 12.01 + 6 \times 1.008$
$= 30.07$

6 **a** $M_r = A_r(\text{K}) + A_r(\text{Cl})$
$= 39.10 + 35.45$
$= 74.55$

b $M_r = 2 \times A_r(\text{Na}) + A_r(\text{C}) + 3 \times A_r(\text{O})$
$= 2 \times 22.99 + 12.01 + 3 \times 16.00$
$= 105.99$

c $M_r = 2 \times A_r(\text{Al}) + 3 \times A_r(\text{S}) + 12 \times A_r(\text{O})$
$= 2 \times 26.98 + 3 \times 32.07 + 12 \times 16.00$
$= 342.17$

3.4 Electronic structure of atoms

3.4 KEY QUESTIONS

1 When a sample containing copper is heated in the flame of a Bunsen burner, the flame turns a green colour. This is because the *electrons* in the copper atoms absorb energy and move to *higher* energy levels and then *emit* light that corresponds to a green colour as they return to *lower* energy levels.

2 An emission spectrum is the spectrum of the light emitted by an element when it is heated. Atoms are heated so that electrons move to higher energy levels, before returning to lower levels and emitting light.

3 Each line in an emission spectrum corresponds to a specific amount of energy. This energy is emitted when electrons from higher-energy electron shells transition to a lower-energy shell. The different spectral lines indicate that there are energy differences between shells. This is evidence that electrons are found in shells with discrete energy levels.

4 Electrons revolve around the nucleus in fixed, circular orbits.

Electrons’ orbits correspond to specific energy levels in the atom.

Electrons can only occupy fixed energy levels and cannot exist between two energy levels.

Orbits of larger radii correspond to higher energy levels.

5 Energy is emitted as light.

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3.5 Electronic configuration and the shell model

Worked example: Try yourself 3.5.1

CALCULATING ELECTRONIC CONFIGURATIONS FOR UP TO 36 ELECTRONS

Apply the rules of the shell model to determine the electronic configuration of an atom with 34 electrons.

Thinking	Working
Recall the maximum number of electrons that each shell can hold.	Shell (*n*) Maximum number of electrons 1 2 2 8 3 18 4 32
Place the first 18 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.	Shell (*n*) Electrons in atom 1 2 2 8 3 8 4
Place the next two electrons in the fourth shell.	Shell (*n*) Electrons in atom 1 2 2 8 3 8 4 2
Continue filling the third shell until it holds up to 18 electrons. Put any remaining electrons in the fourth shell.	Shell (*n*) Electrons in atom 1 2 2 8 3 18 4 6 The 10 remaining electrons from the previous step have gone into the third shell.
Write the electronic configuration by listing the number of electrons in each shell separated by commas (with no space between).	The electronic configuration is 2,8,18,6

Worked example: Try yourself 3.5.2

CALCULATING THE NUMBER OF VALENCE ELECTRONS IN AN ATOM

How many valence electrons are present in an atom of sulfur, which has an atomic number of 16?

Thinking	Working
Recall the maximum number of electrons that each shell can hold.	Shell (*n*) Maximum number of electrons 1 2 2 8 3 18 4 32
Place 16 electrons in the shells from the lowest energy to the highest energy. Do not exceed the maximum number of electrons allowed.	Shell (*n*) Electrons in atom 1 2 2 8 3 6 4
Write the electronic configuration by listing the number of electrons in each shell, separated by commas.	The electronic configuration is 2,8,6.
Determine the number of electrons in the outer shell or valence shell.	The number of electrons in the valence shell is 6.

3.5 KEY QUESTIONS

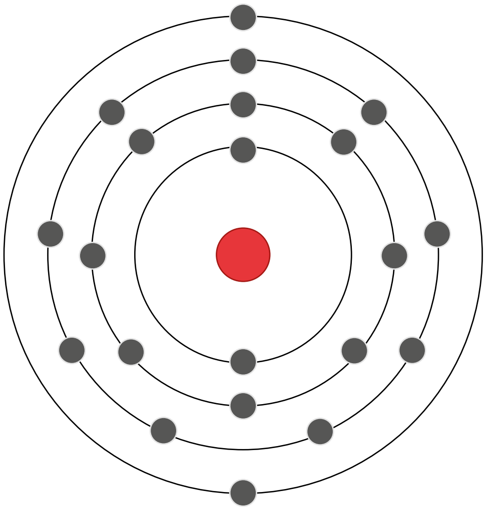
1 Each shell (*n*) can hold up to $2n^2$ electrons. The fourth shell always accepts 2 electrons after the third electron has received 8 electrons. Once the fourth shell has 2 electrons, the third shell will continue to fill.

- **a** 2,3
- **b** 2,8,2
- **c** 2,8,8,2
- **d** 2,8,18,7

2

- **a** 2,2
- **b** 2,8,6
- **c** 2,8,8
- **d** 2,8,2
- **e** 2,8

3



Sc 2,8,9,2

4

- **a** helium, He
- **b** fluorine, F
- **c** aluminium, Al
- **d** nitrogen, N
- **e** chlorine, Cl

5 The electronic configuration is 2,8,18,7. Therefore, the number of valence electrons is 7.

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3.6 The Schrödinger model of the atom

Worked example: Try yourself 3.6.1

WRITING ELECTRONIC CONFIGURATIONS USING THE SCHRÖDINGER MODEL

Write the Schrödinger model of electronic configuration for a titanium atom with 22 electrons.

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Recall the order in which the subshells fill by listing them from lowest energy to highest energy and the number of orbitals in each.</td><td>$1s$, 1 orbital
$2s$, 1 orbital
$2p$, 3 orbitals
$3s$, 1 orbital
$3p$, 3 orbitals
$4s$, 1 orbital
$3d$, 5 orbitals
$4p$, 3 orbitals</td></tr>
<tr><td>Fill the subshells by assigning two electrons per orbital, starting from the lowest energy subshells until you have reached the total number of electrons in your atom.</td><td><table>
<tr><th>Subshell</th><th>Electrons in subshell</th><th>Progressive total of electrons</th></tr>
<tr><td>$1s$</td><td>2</td><td>2</td></tr>
<tr><td>$2s$</td><td>2</td><td>4</td></tr>
<tr><td>$2p$</td><td>6</td><td>10</td></tr>
<tr><td>$3s$</td><td>2</td><td>12</td></tr>
<tr><td>$3p$</td><td>6</td><td>18</td></tr>
<tr><td>$4s$</td><td>2</td><td>20</td></tr>
<tr><td>$3d$</td><td>2</td><td>22</td></tr>
</table></td></tr>
<tr><td>Write the electronic configuration by writing each subshell with the number of electrons as a superscript. Remember to group subshells from the same shell.</td><td>$1s^22s^22p^63s^23p^63d^24s^2$</td></tr>
</table>

3.6 KEY QUESTIONS

1 B. According to the Aufbau principle, the $4s$-subshell is filled before the $3d$-subshell. The $4p$ and $5s$-subshells are also higher in energy.

2

Element (atomic number)	Electronic configuration using the shell model	Electronic configuration using the subshell model
boron (5)	2,3	$1s^22s^22p^1$
lithium (3)	2,1	$1s^22s^1$
chlorine (17)	2,8,7	$1s^22s^22p^63s^23p^5$
sodium (11)	2,8,1	$1s^22s^22p^63s^1$
neon (10)	2,8	$1s^22s^22p^6$
potassium (19)	2,8,8,1	$1s^22s^22p^63s^23p^64s^1$
scandium (21)	2,8,9,2	$1s^22p^22p^63s^23p^63d^14s^2$
iron (26)	2,8,14,2	$1s^22s^22p^63s^23p^63d^64s^2$
bromine (35)	2,8,18,7	$1s^22s^22p^63s^23p^63d^{10}4s^24p^5$

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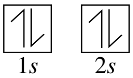
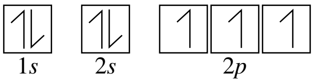
3 The subshell model is a refinement of the shell model. The shell model proposed that all electrons in the one shell were of equal energy. Evidence from emission spectra indicated that there were different electronic energy levels (called subshells) within a shell.

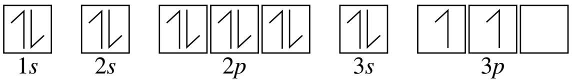
4 **a**



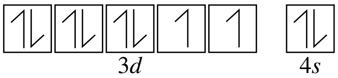
b



c



d



CHAPTER 3 **REVIEW**

1 in the nucleus

2 The protons and neutrons form the nucleus. The electrons are grouped in shells and occupy the space around the nucleus.

3 The mass of a proton is approximately equal to the mass of a neutron and is about 1800 times the mass of an electron. The proton and electron have equal but opposite charges, and the neutron has no charge.

4 **a** 1.20×10^{-10} m

b 2.16×10^{-10} m

c 3.48×10^{-10} m

5 **a** atomic number is 24; mass number is 52

b 24 electrons, 24 protons, 52 – 24 = 28 neutrons

6 No. Isotopes have the same number of protons (atomic number) but different numbers of neutrons (and therefore different mass numbers). These atoms have different atomic numbers and different mass numbers.

7 Atoms are electrically neutral. The positive charge on one proton balances the negative charge on one electron. Therefore, for electrical neutrality, there must be an equal number of protons and electrons.

8 Most elements have more than one isotope, so they will have more than one mass number. All bromine atoms have 35 protons in their nuclei. No other type of atom has 35 protons in its nucleus (i.e. no other atom has an atomic number of 35). Isotopes of bromine, however, differ in their mass numbers, so mass number is not fixed for an element (except for those elements such as sodium, which have only one naturally occurring isotope). In addition, an isotope of one element may have the same mass number as an isotope of another element.

9 **a** The relative molecular mass, M_r, is the sum of the relative atomic masses, A_r, of the elements in the compound.

$$\therefore M_r(\text{H}_2\text{O}) = 2 \times 1.008 + 16.00$$
$$= 18.02$$

b $M_r(\text{P}_4) = 4 \times 30.97$
$= 123.88$

c $M_r(\text{CO}) = 12.01 + 16.00$
$= 28.01$

10 **a** Relative formula mass of zinc bromide = 65.38 + 2 × 79.90 = 225.18

b Relative formula mass of barium hydroxide = 137.3 + 2 × 16.00 + 2 × 1.008 = 171.3

c Relative formula mass of iron(III) carbonate = 2 × 55.85 + 3 × 12.01 + 9 × 16.00 = 291.73

11 **a** ${}^{231}_{87}\text{Fr} \rightarrow {}^{231}_{88}\text{Ra} + {}^{0}_{-1}\text{e}$

b ${}^{239}_{94}\text{Pu} \rightarrow {}^{235}_{92}\text{U} + {}^{4}_{2}\text{He}$

c ${}^{231}_{91}\text{Pa} \rightarrow {}^{227}_{89}\text{Ac} + {}^{4}_{2}\text{He}$

d ${}^{214}_{84}\text{Po} \rightarrow {}^{210}_{82}\text{Pb} + {}^{4}_{2}\text{He}$

e ${}^{149}_{62}\text{Sm} \rightarrow {}^{145}_{61}\text{Pm} + {}^{4}_{2}\text{He} + {}^{0}_{-1}\text{e}$

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12 The relative atomic mass of carbon is the weighted average of the isotopic masses of all carbon isotopes (i.e. ^{12}C, ^{13}C and ^{14}C). Small amounts of ^{13}C and ^{14}C make this average slightly greater than 12, the relative isotopic mass of the ^{12}C isotope.

13 $A_r(\text{Pd}) = \frac{(101.9049 \times 0.9600) + (103.9036 \times 10.97) + (104.9046 \times 22.23) + (105.9032 \times 27.33) + (107.9039 \times 26.71) + (109.9044 \times 11.800)}{100} = 106.4$

14 **a** $A_r(\text{Ar}) = \frac{(35.978 \times 0.307) + (37.974 \times 0.060) + (39.974 \times 99.633)}{100} = 39.96$

$A_r(\text{K}) = \frac{(38.975 \times 93.3) + (39.976 \times 0.011) + (40.974 \times 6.69)}{100} = 39.11$

b Although potassium atoms have one more proton than argon atoms, the most abundant isotope of argon has 22 neutrons, giving it a relative atomic mass close to 40. The most abundant isotope of potassium has only 20 neutrons, giving it a relative atomic mass close to 39.

15 **a** Peak heights: ^{50}Cr = 0.3 units, ^{52}Cr = 12 units, ^{53}Cr = 1 unit, ^{54}Cr = 0.2 units; total height = 13.5 units

Percentages: ^{50}Cr = 2.2%, ^{52}Cr = 88.9%, ^{53}Cr = 7.4%, ^{54}Cr = 1.5%

b $A_r(\text{Cr}) = \frac{(50 \times 2.2) + (52 \times 88.9) + (53 \times 7.4) + (54 \times 1.5)}{100} = 52$

16 Let the percentage abundance of the lighter isotope be x.

$\therefore$ percentage abundance of the heavier isotope will be $(100 - x)$

$\therefore 151.96 = \frac{(x \times 150.92) + ((100 - x) \times 152.92)}{100}$

$\therefore x = 48.0$

$\therefore$ proportions of the isotopes are 48.0% and 52.0%.

17 **a** Let the percentage abundance of the lighter isotope be x.

$\therefore$ percentage abundance of the heavier isotope will be $100 - x$

$\therefore 69.72 = \frac{68.95x + ((100 - x) \times 70.95)}{100}$

$\therefore x = 61.5$

$\therefore$ proportion of the lighter isotope = 61.5%

b 20.2% (calculated using the same process as for part **a**)

18 $n = 1$

19 $n = 3$. The first 2 electrons fill the first shell, the next 8 the second shell, the next 8 the third shell. The next 2 electrons go into the fourth shell and the remaining 10 electrons go into the third shell.

20 Magnesium. A total of 12 electrons means the element has 12 protons. The element with 12 protons and atomic number 12 is magnesium.

21 **a** $1s^2$

b $1s^22s^22p^2$

c $1s^22s^22p^5$

d $1s^22s^22p^63s^23p^1$

e $1s^22s^22p^63s^23p^6$

f $1s^22s^22p^63s^23p^63d^84s^2$

g $1s^22s^22p^63s^23p^63d^{10}4s^24p^5$

22 A fluorine atom contains nine electrons. The electrons are arranged in energy levels called shells; two electrons are in the first shell and seven electrons are in the second shell, which has higher energy. The electron arrangement in the shells can be written as 2,7.

Shells are regarded as being made up of energy levels called subshells. The first shell contains an *s*-type subshell, which is labelled '1*s*'. The second shell contains both *s*- and *p*-type subshells, labelled '2*s*' and '2*p*', respectively.

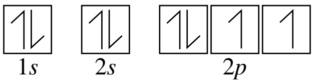
Within subshells, electrons occupy regions of space known as orbitals. An orbital can hold up to two electrons. Subshells of an *s*-type contain one orbital, whereas *p*-type subshells contain three orbitals. The electron arrangement in the subshells of a fluorine atom can be represented as $1s^22s^22p^5$.

23 $1s^22s^22p^63s^23p^4$

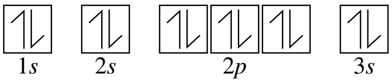
24 In the Schrödinger model of the atom, electron shells are divided into subshells, and each subshell can have a different energy level. According to the Schrödinger model, the 4*s*-subshell is lower in energy than the 3*d*-subshell. Therefore, the 4*s*-subshell begins filling after the 3*s*- and 3*p*-subshells, but before the 3*d*-subshell.

25 There is very little difference between the energy levels of the 3*d*- and 4*s*-subshells. As orbitals fill in sequence, one electron at a time, it is more stable for chromium to have all *d*-orbitals exactly half-filled than to have one empty *d*-orbital. Likewise, for copper it is more stable to have all *d*-orbitals completely filled than to have one half-filled and a full 4*s*-orbital.

26 **a**



b



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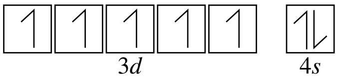

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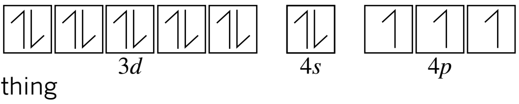

27 nothing

28 **a** Atomic models before Bohr's accounted for a number of atomic properties, but were not able to account for the characteristic emission spectrum of each element. Previous models were also in conflict with the principles of classical physics, which suggested that electrons moving in circular orbits should continuously lose energy and spiral into the nucleus.

b The Bohr model of the atom did not adequately explain why electrons adopted some energy levels but not others. In addition, calculated frequencies for lines in the emission spectra of atoms with more than one electron gave poor agreement with measured values.

29 Student answers will vary. Possible answers include the lack of a magnetic field and that the balls representing the particles are not charged.

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Chapter 4 **Periodicity**

4.1 The periodic table

4.1 KEY QUESTIONS

1 row
2 groups 1, 2 and 13–18
3 **a** 1
    b 15 – 10 = 5
    c 17 – 10 = 7
    d 2
4 The element is in period 3 and therefore has three occupied shells. As the element is in group 2, it will have two valence electrons. This gives an electron configuration of 2,8,2 or $1s^22s^22p^63s^2$.
5 **a** **i** 13
        ii 17
        iii 1
        iv 18
        v 14
        vi 14
    b **i** 4
        ii 2
        iii 1
        iv 1
        v 7
        vi 3
    c **i** silicon, Si, 2,8,4 or $1s^22s^22p^63s^23p^2$
        ii beryllium, Be, 2,2 or $1s^22s^2$
        iii argon, Ar, 2,8,8 or $1s^22s^22p^63s^23p^6$
6 Atoms are listed in the periodic table in order of atomic number because the atomic number determines the number of protons and thus electrons, and this in turn sets the properties that make each element unique. Relative atomic mass does not necessarily increase as atomic number increases.

4.2 Trends in the periodic table: Part 1

Worked example: Try yourself 4.2.1

DETERMINING CORE CHARGE

Determine the core charge of an atom of fluorine.

Thinking	Working
Determine the number of electrons in an atom of the element, using the periodic table as a reference.	The atomic number of fluorine is 9. Therefore, an atom of fluorine has 9 protons and 9 electrons.
Use the number of electrons to determine the electronic configuration.	With 9 electrons, the electronic configuration is $1s^22s^22p^5$.
Determine the core charge Core charge = number of protons – number of inner-shell electrons	There are 2 inner-shell electrons. Core charge = 9 – 2 = +7

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4.2 KEY QUESTIONS

1 Carbon has 6 protons and 6 electrons. Its electronic configuration is $1s^22s^22p^2$. Carbon has two inner-shell electrons. Core charge = 6 – 2 = +4.

2 As core charge increases, electrons are more strongly attracted to the nucleus of an atom, so electronegativity increases.

3 **a** **i** fluorine, F
  ii francium, Fr
 b **i** group 17
  ii group 1
 c Elements in group 18, the noble gases, have a very stable electronic configuration and so are unreactive.

4 F, O, P, Mg, Ca

5 hydrogen, helium, nitrogen, oxygen, fluorine, neon, chlorine, argon, krypton, xenon and radon

4.3 Trends in the periodic table: Part 2

Worked example: Try yourself 4.3.1

DETERMINING FIRST IONISATION ENERGY

Determine which of these atoms has the greatest first ionisation energy: Al, B, C, Si

Thinking	Working
Determine the group and period for each element.	Al is in group 13, period 3 B is in group 13, period 2 C is in group 14, period 2 Si is in group 14, period 3
Recall that first ionisation energies generally increase across a period and decrease down a group.	Of the four elements, C and Si are in group 14 and are expected to have greater first ionisation energies than the elements in group 13. C is in period 2, so is expected to have a greater first ionisation energy than Si, which is in period 3. C has the greatest first ionisation energy.

4.3 KEY QUESTIONS

1 **a** First ionisation energy is the amount of energy needed to remove an electron from an atom in the gas phase.
 b The factors that affect first ionisation energy across a period are the size of the atom (i.e. the distance of the outermost (highest energy) electron from the nucleus) and the core charge.

2 A metalloid is an element that exhibits both metallic and non-metallic properties.

3 Across a period, the number of occupied shells in the atoms remains constant, but core charge increases. The valence electrons become more strongly attracted to the nucleus, so more energy is required to remove an electron from an atom. Therefore, the first ionisation energy increases across a period.

4 The valence electrons in strontium are less tightly held than those of beryllium. They are easily lost in reactions, such as with water.

5 **a** fluorine
 b Aluminium and magnesium are metals. Fluorine has the highest core charge (it is located farthest right on the periodic table), while also containing the least number of electron shells. This means that fluorine attracts an electron more strongly than other elements and is therefore the most reactive non-metal.

CHAPTER 4 REVIEW

1 The atomic number is determined by the number of protons in the nucleus of an atom.

2 **a** period 1, *s*-block
- **b** period 2, *p*-block
- **c** period 3, *p*-block
- **d** period 4, *d*-block
- **e** period 7, *f*-block

3 **a** period 2, group 2
- **b** period 3, group 14
- **c** period 4, group 13
- **d** period 1, group 18

4 **a** Elements in the *s*-block are filling an *s*-subshell. As an s-subshell accommodates a maximum of two electrons, there are only two groups of elements in the *s*-block. They have outer-shell configurations of s^1 and s^2.
- **b** Elements in the *p*-block are filling a *p*-subshell. As a *p*-subshell accommodates a maximum of six electrons, there are six groups of elements in the *p*-block. They have outer-shell configurations of s^2p^1 to s^2p^6.
- **c** Elements in the *d*-block are filling a *d*-subshell. As a *d*-subshell accommodates a maximum of 10 electrons, there are 10 elements in each transition series.
- **d** The lanthanoids and actinoids are filling an *f*-subshell. As an *f*-subshell accommodates a maximum of 14 electrons, there are 14 elements in each of these series.

5 **a** silicon
- **b** potassium, caesium
- **c** bromine
- **d** nitrogen, arsenic

6 Melting points increase across periods from groups 1 to 14 and then are drastically lower in groups 15–18.

7 bromine and mercury

8 As you move from left to right across groups 1, 2 and 13–17, the charge on the nucleus increases. Each time the atomic number increases by one, the electrons are attracted to an increasingly positive nucleus. Within a period, the outer electrons are in the same shell—that is, they have the same number of inner-shell electrons shielding them from the nucleus. Therefore, the additional nuclear charge attracts the electrons more strongly, drawing them closer to the nucleus and so decreasing the size of the atom.

9 **a** Magnesium and phosphorus, with outer electrons in the third shell, are in the same period. Magnesium has a nuclear charge of +12, but with completed inner shells of $1s^22s^22p^6$, the outer electrons experience the attraction of a core charge of +2. The outer-shell electrons of phosphorus, which has a nuclear charge of +15 and the same number of inner shells as magnesium, are attracted by a core charge of +5. The stronger attraction of the phosphorus electrons to the core means that more energy is required to remove an electron from a phosphorus atom than from a magnesium atom.
- **b** Both fluorine and iodine are in group 17, so the outer electrons of each atom experience the attraction of the same core charge. Because the outer-shell electrons of a fluorine atom are closer to the nucleus than those of an iodine atom, they are attracted more strongly and so more energy is needed to remove one.

10 **a** $1s^22s^22p^3$
- **b** period 2 and group 15
- **c** 5
- **d** +5

11 As the atomic number increases across a period of the periodic table, the number of protons in the nucleus increases, pulling the electrons added to the outermost shell towards the nucleus more strongly. This results in a decreased atomic radius. Once each shell is filled, a new period is started and additional electrons are added to a new shell, making the radius increase.

12 $\frac{120}{2} = 60$ pm

13 **a** The radius of the atoms decreases as the core charge increases.
- **b** There is a trend from metals (lithium, beryllium) to metalloids (boron) to non-metals (carbon, nitrogen, oxygen, fluorine and neon).
- **c** Electronegativity increases as the core charge increases and size of the atoms decreases.

14 **a** lithium, sodium, potassium, rubidium

b The reactivity of metals increases down a group. This is because the number of electron shells increases down a group, so the valence electrons are farther from the nucleus and more easily lost. Since metals lose electrons in their reactions, those that lose electrons most easily will be most reactive.

15 **a** nitrogen

b chlorine

c chlorine

16 The reactivity decreases then increases. Metal atoms lose electrons in chemical reactions, and it is easier for atoms with a low core charge to do this. These atoms are found on the left-hand side of the periodic table. Moving across the period from left to right, core charge increases and reactivity decreases. However, in the non-metals, which are located on the right-hand side of the period, electrons are gained or shared by their atoms in chemical reactions, so as the core charge increases from left to right across the periodic table, the reactivity of non-metals increases.

17 Metals conduct electricity, are usually solid at room temperature and are located on the left side of the periodic table. They often contain one, two or three valence electrons.

18 The first ionisation energy decreases from Mg to Al because of their electronic configurations. Mg ($1s^22s^22p^63s^2$) has a completely filled 3*s*-subshell. In Al ($1s^22s^22p^63s^23p^1$), the valence electron is in the 3p-subshell and slightly farther from the nucleus, and it experiences a little bit more shielding from the nuclear charge due to the electrons in the 3*s*-subshell. Therefore, the valence electron in Al are not as strongly attracted to the nucleus, so Al has a lower first ionisation energy.

19 Student answers will vary. Possible answer: Periodicity refers to the periodic, or regularly recurring, pattern of properties of the elements when they are listed in order of atomic number.

20 Student answers will vary. Possible answer: The first periodic table, Benfey's spiral periodic table, illustrates the sizes of different periods (two periods of eight elements, two periods of 18 elements, two periods of 32 elements, etc.). The second, the Janet form of the periodic table, arranges the elements in order of the filling of electron subshells. Student answers will vary for additional forms of the periodic table.

21 Student answers will vary. Ensure the assigned labels are internally consistent and accurately reflect atomic radii.

Chapter 5 Bonding

5.1 Metallic bonding

Worked example: Try yourself 5.1.1

EXPLAINING THE ELECTRICAL CONDUCTIVITY OF MAGNESIUM

With reference to the electronic configuration of magnesium, explain why solid magnesium can conduct electricity.

Thinking	**Working**
Using the atomic number of the element, determine the electronic configuration of its atoms. (You may need to refer to a periodic table.)	Mg has an atomic number of 12. This means that a neutral atom of Mg has 12 electrons. The electronic configuration is $1s^{2}2s^{2}2p^{6}3s^{2}$.
From the electronic configuration, find how many outer-shell electrons are lost to form cations that have a stable octet of valence electrons. These electrons become delocalised.	Mg has two electrons in its outer shell (the $3s^{2}$ electrons). Mg atoms will tend to lose these two valence electrons to form a cation with a charge of +2. The outer-shell electrons become delocalised and form the sea of delocalised electrons within the metal lattice.
An electric current occurs when there are free-moving charged particles.	If the Mg is part of an electric circuit, the delocalised electrons can move through the lattice towards a positively charged electrode.

5.1 KEY QUESTIONS

1. **a** Li atoms have three electrons. The electronic structure is $1s^{2}2s^{1}$. There is one electron in the outer shell. The charge of the cation will therefore be +1.
   b Mg atoms have 12 electrons. The electronic structure is $1s^{2}2s^{2}2p^{6}3s^{2}$. There are two electrons in the outer shell. The charge of the cation will therefore be +2.
   c Ga atoms have 31 electrons. The electronic structure is $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{1}$. There are three electrons in the outer shell. The charge of the cation will therefore be +3.
   d Ba atoms have 56 electrons. The electronic structure is $1s^{2}2s^{2}2p^{6}3s^{2}3p^{6}3d^{10}4s^{2}4p^{6}4d^{10}5s^{2}5p^{6}6s^{2}$. There are two electrons in the outer shell. The charge of the cation will therefore be +2. Another way to approach this is to notice that barium is in the same group as magnesium. All alkaline earth metals form +2 cations by losing their two outer-shell electrons.
2. **a** Both potassium and gold have good thermal and electrical conductivity. However, gold has a higher density and higher melting and boiling points than potassium.
   b sodium
   c silver
   d Sodium and potassium are in group 1. Gold and silver are in the transition metals.
3. **a** silver, copper, gold, aluminium
   b Availability and cost need to be considered. Other properties might include malleability and ductility.
4. Sodium belongs to the alkali metals which have relatively low melting and boiling points, relatively low density and are relatively soft. Iron is a transition metal. Transition metals have relatively high melting and boiling points, relatively high density and are relatively hard.
5. tensile strength, cost, availability

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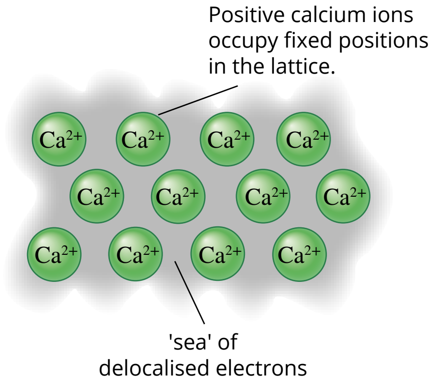
6 **a**



b strong electrostatic forces of attraction between Ca^{2+} ions and the delocalised valence electrons

7 Barium has a high melting point because there are strong attractive forces between the positive ions and the delocalised electrons. Barium conducts electricity because the delocalised electrons in the outer shell are free to move.

8 **a** Both graphite and metals are lustrous and conduct heat and electricity.

b These properties are explained by the presence of free-moving electrons. Both graphite and metals must contain delocalised electrons.

5.2 Ionic bonding

Worked example: Try yourself 5.2.1

WRITING A CHEMICAL FORMULA

Determine the chemical formula of the ionic compound formed between barium and fluoride ions. You may need to refer to Tables 5.2.3 and 5.2.4 on page 156 of the Student Book.

Thinking	Working
Write the symbol and charge of the two ions forming the ionic compound.	Ba^{2+} and F^-
Calculate the lowest common multiple of the two numbers in the charges of the ions.	$2 \times 1 = 2$
Calculate how many positive ions are needed to equal the lowest common multiple.	one Ba^{2+} ion
Calculate how many negative ions are needed to equal the lowest common multiple.	two F^- ions
Use the answers from the previous two steps to write the formula for the ionic compound. Write the symbol of the positive ion first. (Note that 1 is not written as a subscript.)	BaF_2

5.2 KEY QUESTIONS

1 B. Substance B has a melting point of 308°C. B will not conduct electricity at 250°C, when it is solid, but will when it is at 350°C, when it is molten. This is characteristic of ionic compounds.

2 **a** The diagram shows that in solid sodium chloride, the sodium and chloride ions are held in fixed positions in the crystal lattice and are not free to move and conduct electricity.

b Molten sodium chloride contains sodium and chloride ions that are free to move and, therefore, it can conduct electricity.

3 The electrostatic forces of attraction between the positive and negative ions holding the lattice together are very strong, and a lot of energy is required to break them apart.

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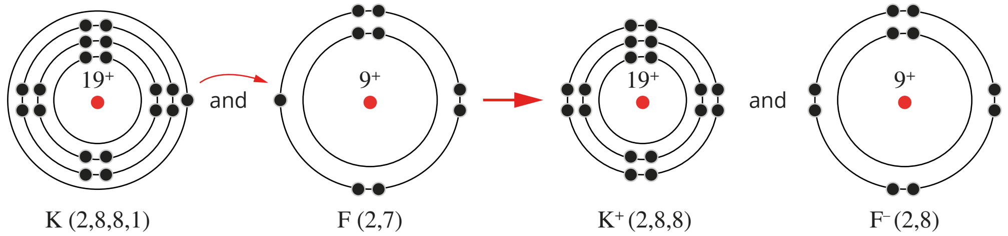
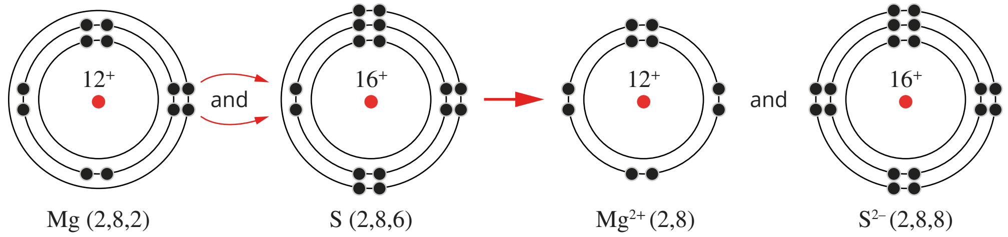
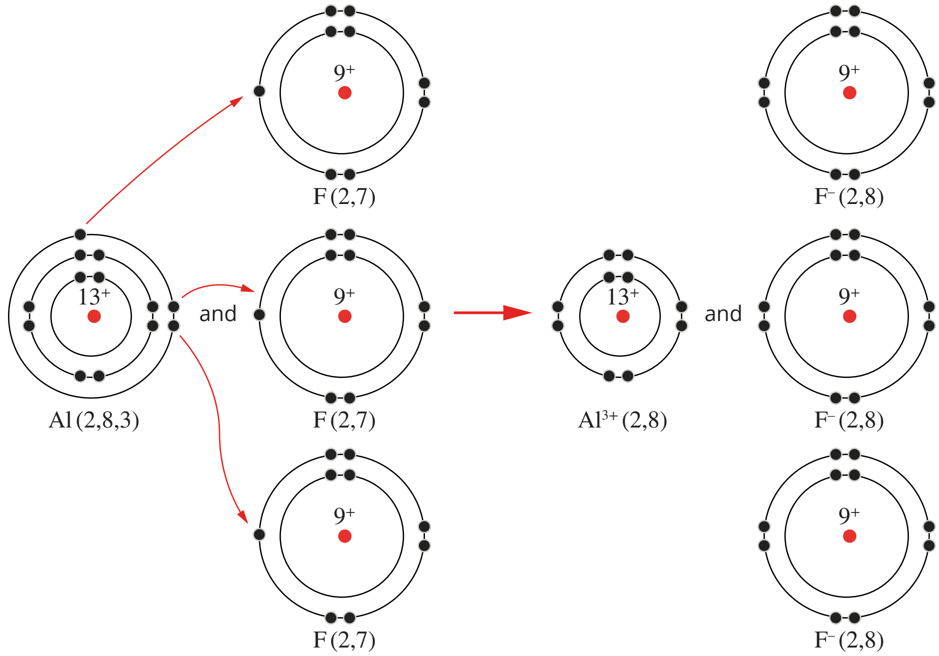
4 a



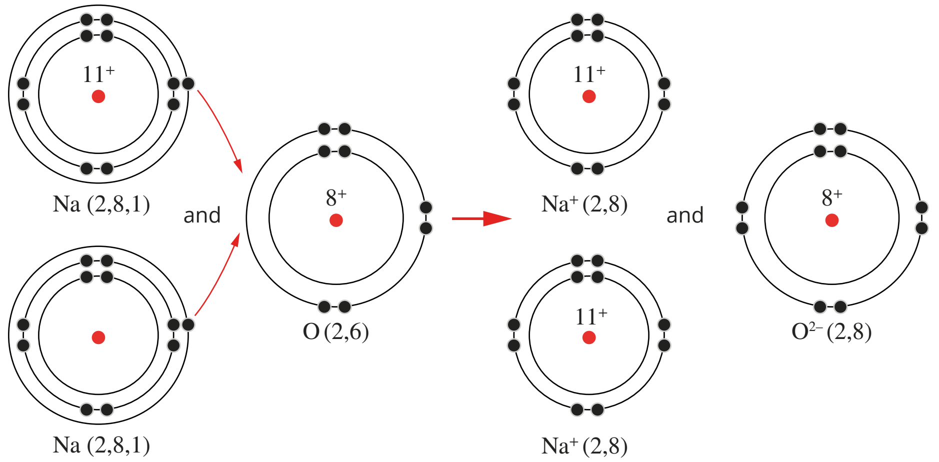
b



c



d



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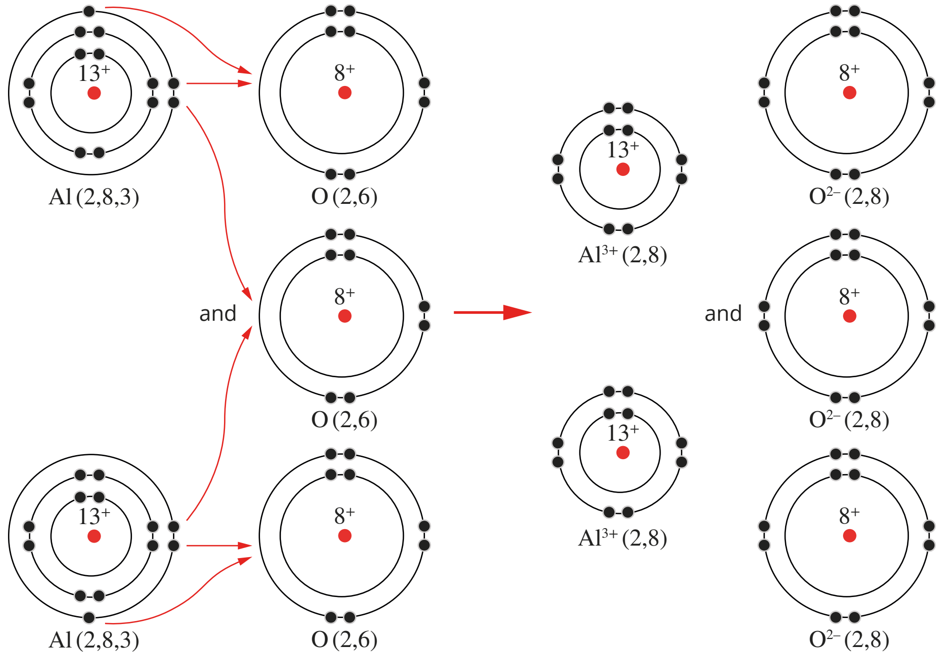
e



5 The electronic configuration of an atom of potassium is $1s^22s^22p^63s^23p^64s^1$ and that of an atom of chlorine is $1s^22s^22p^63s^23p^5$. Because an atom of K has one more electron than an atom of a noble gas and an atom of Cl has one less, in a reaction one K atom can donate one electron to one Cl atom to give K^+ ($1s^22s^22p^63s^23p^6$) and Cl^- ($1s^22s^22p^63s^23p^6$), both of which have the electronic configuration of the noble gas argon. The formula of this compound is, therefore, KCl.

Calcium, however, has the electronic configuration of $1s^22s^22p^63s^23p^64s^2$. A Ca atom will lose two electrons to gain a noble gas configuration. Because each Cl atom will gain only one electron, there needs to be two chlorine atoms for each Ca atom. This reaction will therefore produce Ca^{2+} ($1s^22s^22p^63s^23p^6$) ions and Cl^- ($1s^22s^22p^63s^23p^6$) ions. The formula of the compound is, therefore, $CaCl_2$.

6 **a** NaCl
b KBr
c $ZnCl_2$
d K_2O
e $BaBr_2$
f Al_2I_3
g AgBr
h ZnO
i BaO
j Al_2S_3

7 **a** potassium chloride
b calcium oxide
c magnesium sulfide
d potassium oxide
e sodium fluoride

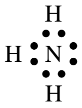
8 **a** Na_2CO_3
b $Ba(NO_3)_2$
c $Al(NO_3)_3$
d $Ca(OH)_2$
e $Zn(SO_4)_2$
f KOH
g KNO_3
h $ZnCO_3$
i K_2SO_4
j $Ba(OH)_2$

5.3 Covalent bonding

Worked example: Try yourself 5.3.1

DRAWING LEWIS DOT DIAGRAMS

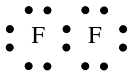
Draw a Lewis dot diagram of ammonia (NH_3).

Thinking	**Working**
Write the electronic configuration of the atoms in the molecule.	N—electronic configuration: 2,5 H—electronic configuration: 1
Determine how many electrons each atom requires for a stable outer shell.	N requires 3 electrons H requires 1 electron
Draw a Lewis dot diagram of the likely molecule, ensuring that each atom has a stable outer shell. Any electrons not involved in bonding will be in non-bonding (lone) pairs.	Draw a Lewis dot diagram of the molecule. 

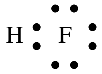
5.3 KEY QUESTIONS

1 **a** 1. Hydrogen requires 1 electron to complete its outer shell, so it forms a single covalent bond.
  b 3. Nitrogen requires 3 electrons to complete its outer shell, so it forms a triple covalent bond.
  c 2. Oxygen requires 2 electrons to complete its outer shell, so it forms a double covalent bond.
  d 1. Fluorine requires 1 electron to complete its outer shell, so it forms a single covalent bond.

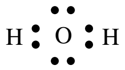
2 fluorine (F_2)



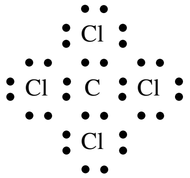
hydrogen fluoride (HF)



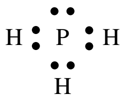
water (H_2O)



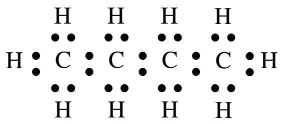
tetrachloromethane (CCl_4)



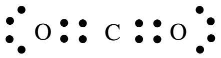
phosphine (PH_3)



butane (C_4H_{10})



carbon dioxide (CO_2)



3 **a** 1
  b 2
  c 3
  d 4
  e 1
  f 0

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4 To complete its outer shell, the oxygen atom uses two of its outer-shell electrons to form two single bonds or a double bond with suitable non-metal atoms. The remaining four electrons in the outer shell are not required for bonding, as the outer shell is now complete, and they arrange themselves as two lone pairs around the oxygen atom.

5 **a** CCl_4, carbon tetrachloride
b NBr_3, nitrogen tribromide
c SiO_2, silicon dioxide
d HF, hydrogen fluoride
e PF_3, phosphorus trifluoride

6 A conductor must contain free-moving charged particles, such as ion or electrons. Covalent compounds do not contain ions or delocalised electrons and therefore do not conduct an electric current.

7 **a** Non-metallic atoms bond together by sharing one or more valence electrons so that there are eight electrons in the outer shell of each atom.
b The hydrogen atom has only a single electron. This electron is in the first shell, which can hold a maximum of two electrons. The first shell is filled when hydrogen shares one electron from a non-metal atom.

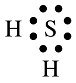
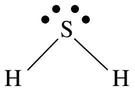
8 **a** Allotropes are different forms of the same element in the same physical state. Allotropes differ in the way atoms are bonded within their structures.
b **i** oxygen and ozone
ii red, white and black phosphorus

5.4 Intermolecular forces

Worked example: Try yourself 5.4.1

PREDICTING THE SHAPE OF MOLECULES

Predict the shape of a molecule of hydrogen sulfide (H_2S).

Thinking	Working
Draw the electron dot diagram for the molecule.	
Count the number of bonds and lone pairs on the central atom.	There are two bonds and two lone pairs.
Determine how the groups of electrons will be arranged to get maximum separation.	Because there are four electron pairs, the groups will be arranged in a tetrahedral arrangement.
Deduce the shape of the molecule by considering the arrangement of just the atoms.	The sulfur and hydrogen atoms are arranged in a bent or V-shape. 

Worked example: Try yourself 5.4.2

COMPARING THE POLARITY OF COVALENT BONDS

Compare the polarity of the covalent bonds in nitrogen monoxide (NO) and hydrogen chloride (HCl).

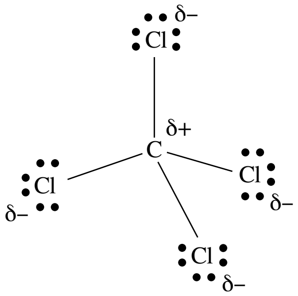
Thinking	Working
Use the table of electronegativity values in Figure 5.4.12 on page 176 of the Student Book to find the electronegativities of the atoms in each molecule.	NO: nitrogen 3.04; oxygen 3.44 HCl: hydrogen 2.20; chlorine 3.16
For each molecule, subtract the lowest electronegativity value from the highest value.	NO: 3.44 – 3.04 = 0.40 HCl: 3.16 – 2.20 = 0.96
Determine which molecule has the biggest difference in electronegativity in order to determine the more polar molecule.	HCl is more polar than NO.

5.4 KEY QUESTIONS

1 **a** V-shaped
  b linear
  c tetrahedral
  b pyramidal
  e linear

2 **a** O
  b C
  c N
  d N
  e F
  f F

3



4 **a** polar
  b polar
  c polar
  d polar
  e non-polar

5 **a** ionic
  b polar covalent
  c non-polar covalent

6 B and E. Hydrogen chloride and chloromethane would form dipole-dipole forces between molecules. They both contain polar covalent bonds and have asymmetrical molecular shapes.

7 C. Hydrogen chloride

8 dipole–dipole forces: b, c, d, e, f; hydrogen bonds: a, g, h; neither: i (dispersion forces between H_2 molecules).

5.5 Covalent network structures

5.5 KEY QUESTIONS

1 **a** Diamond is hard because it has strong covalent bonds throughout the network, with all atoms being held in fixed positions. Graphite is soft because there are weak dispersion forces between the layers in graphite, so layers can be made to slide over each other easily.

  b Diamond is a non-conductor of electricity because all of its electrons are localised in covalent bonds and are not free to move. Graphite can conduct electricity because it has delocalised electrons between its layers of carbon atoms.

2 **a** Graphite is used as a dry lubricant because the dispersion forces between the layers in graphite enable the layers to slide over each other easily and to reduce the friction between moving parts.

  b The strong covalent bonding throughout the network means that the carbon atoms are fixed in place. This makes the diamond very hard and suitable as a material for cutting other less hard materials.

3 Fullerenes are similar to graphite in that they are allotropes of carbon in which each carbon atom has bonds to 3 other carbon atoms. Fullerenes conduct electricity and heat, similar to graphite. Fullerenes differ from graphite in that they are nanomaterials, which give them different physical properties.

4 Each carbon atom in a buckyball is covalently bonded to 3 other carbon atoms. Each carbon atom has one free electron, which is shared between the other carbon atoms in the sphere. The structure consists of hexagonal and pentagonal rings of atoms.

5 Each carbon atom in a graphene sheet is covalently bonded to 3 other carbon atoms. Each carbon atom has one free electron that is shared with the other carbon atoms in the sheet. The structure consists of hexagonal rings of atoms. The sheet is one atom thick but may be of any length and width.

6 The structure of silicon dioxide is a covalent network structure. Each silicon atom sits at the centre of a tetrahedron and is covalently bonded to four oxygen atoms. Each oxygen atom is bonded to two silicon atoms.

CHAPTER 5 **REVIEW**

1 Electrical conductivity in the solid state.

2 **a** When a current is applied to copper wire, the free-moving, delocalised electrons move from one end to the other, so the copper wire conducts electricity.

b The delocalised electrons in the metal spoon obtain energy from the boiling mixture and move more quickly. These electrons move freely throughout the spoon, colliding with other electrons and metal ions, transferring energy, so the spoon becomes warmer and eventually too hot to hold.

c A lot of energy is required to overcome the strong forces of attraction between the iron ions and the delocalised electrons in the metal lattice.

d Because of the strong forces of attraction between them, the lead ions and the delocalised electrons form a closely packed three-dimensional structure. Also, the lead atom itself has a higher mass-to-volume ratio than the sulfur atom. This means that the density—the mass per volume—is high.

e As the copper is drawn out, the copper ions are forced apart. The delocalised electrons rearrange themselves around these ions and re-establish strong forces of attraction.

3 **a** **i** valence electrons that are not restricted to a region between two atoms

ii a regular three-dimensional arrangement of a very large number of positive ions (cations)

iii the electrostatic attraction between a lattice of cations and delocalised electrons

b valence electrons

4 **a** Assemble equipment to test conductivity. Add a globe to the circuit. When the electrodes are touching the solid magnesium chloride, the globe will not light up.

b Using the same equipment with molten sodium chloride, the globe will glow. Care is needed, as sodium chloride melts at 801°C.

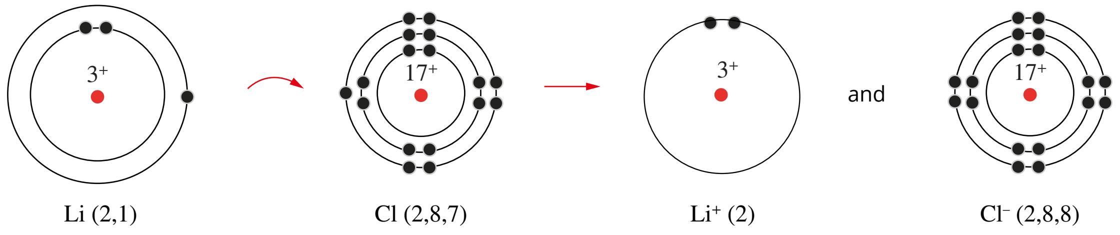
c If a crystal of sodium chloride was hit firmly with a hammer, it would shatter. Again, care is needed—safety glasses must be worn.

5 **a** The electrostatic forces of attraction between the positive and negative ions are strong and will be overcome only at high temperatures.

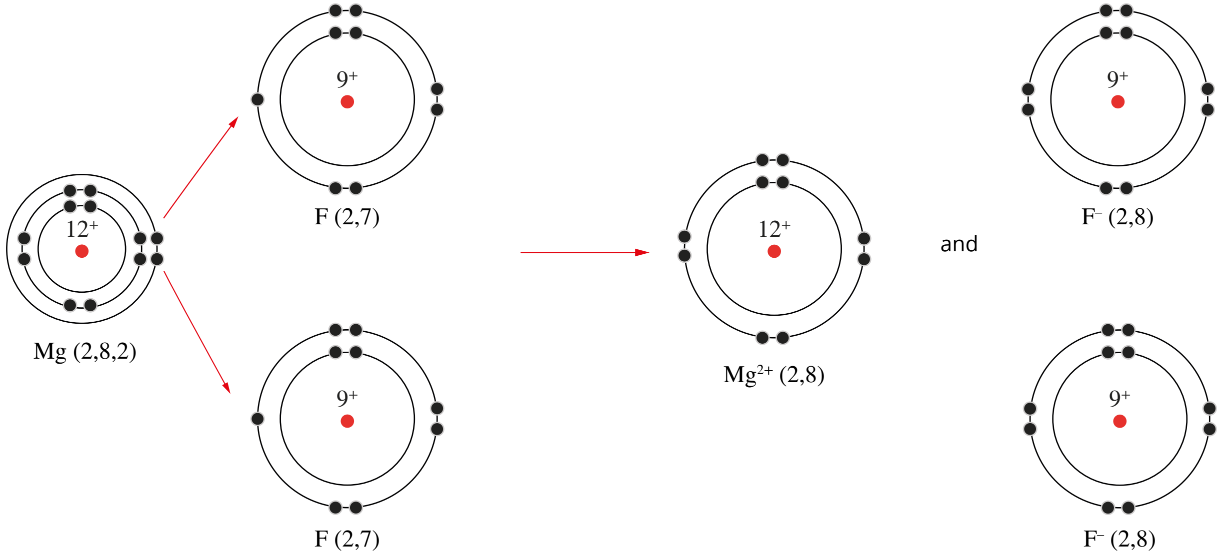
b The strong electrostatic forces of attraction between the ions mean that a strong force is needed to break up the lattice, giving the ionic crystals the property of hardness. However, the crystal lattice will shatter when a strong force is applied, suddenly causing ions of like charge to become adjacent to one another and be repelled.

c In the solid state, the ions are not free to move. However, when the solid melts or dissolves in water, the ions are free to move and conduct electricity.

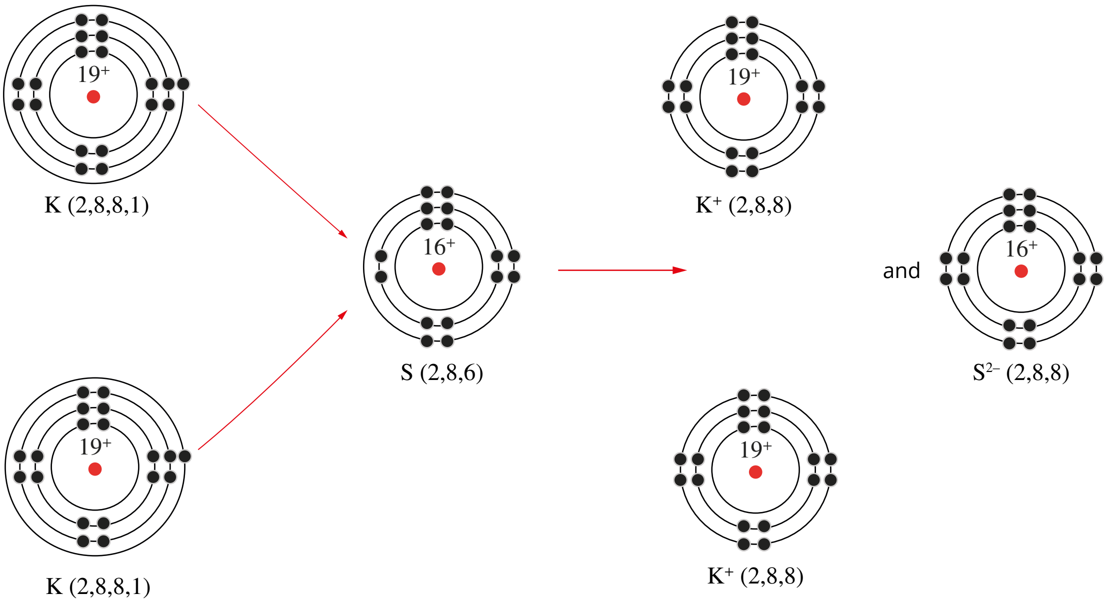
6 **a**



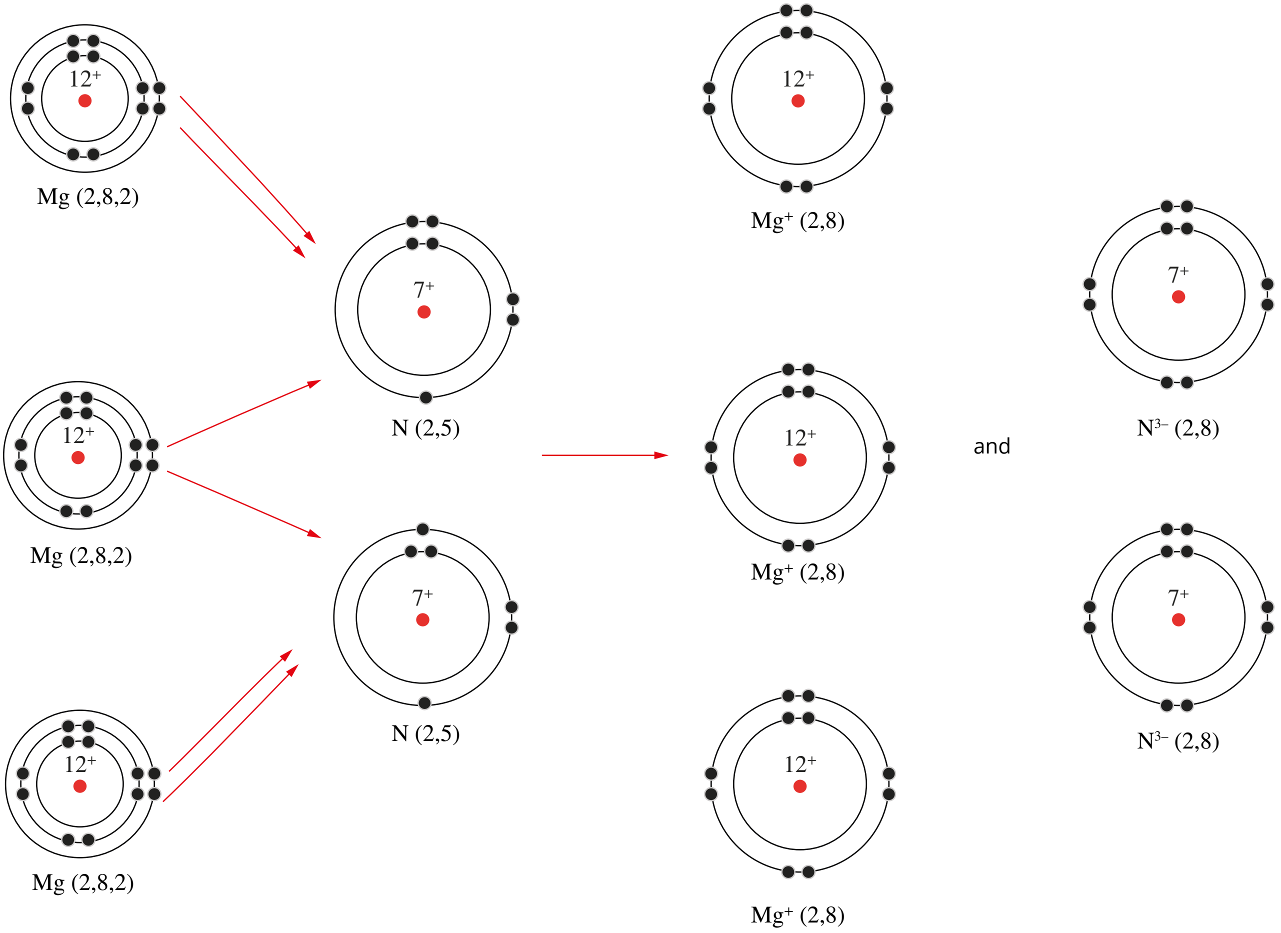
b


12+
Mg (2,8,2)
9+
F (2,7)
9+
F (2,7)
12+
Mg2+ (2,8)
and
9+
F− (2,8)
9+
F− (2,8)

c


19+
K (2,8,8,1)
19+
K (2,8,8,1)
16+
S (2,8,6)
19+
K+ (2,8,8)
19+
K+ (2,8,8)
and
16+
S2− (2,8,8)

d



7 **a** KBr. The potassium ion has a charge of +1, and the bromide ion has a charge of –1.
b MgI_2. The magnesium ion has a charge of +2, and the iodide ion has a charge of –1.
c CaO. The calcium ion has a charge of +2, and the oxide ion has a charge of –2.
d AlF_3. The aluminium ion has a charge of +3, and the fluoride ion has a charge of –1.
e Ca_3N_2. The calcium ion has a charge of +2, and the nitride ion has a charge of –3.

8 **b** CD_3
c EF
d G_3H
e KL

9 **a** CuCl
b Ag_2O
c Li_3N
d KI
e $Cu(NO_3)_2$
f CrF_2
g K_2CO_3
h $Mg(HCO_3)_2$
i $Ni_3(PO_4)_2$

10 **a** ammonium carbonate
b copper(II) nitrate
c chromium(III) bromide

11 The strength of the intermolecular bonds in pure hydrogen chloride must be relatively weak. Since pure hydrochloride exists as a gas at room temperature, it must have a low boiling temperature, indicating that not much energy is required to break the intermolecular bonds between molecules.

12 D. Each hydrogen atom has one electron only. If two atoms share an electron each they both obtain a complete outer shell. A single covalent bond is formed.

13 **a** Six. Each nitrogen atom shares three electrons with the other nitrogen atom.

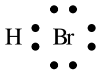
b Three. Each nitrogen atom shares three electrons with the other nitrogen atom, so three bonding pairs are formed, which is equal to a triple covalent bond.

c Each nitrogen atom has two non-bonding electrons, making a total of four when both atoms are counted.

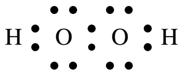
14 Atom X requires 2 electrons to form a stable outer shell. Each atom of element Y can share 1 electron. Therefore, the molecular formula will be XY_2.

15 The diagrams of each molecule below show the number of electrons.

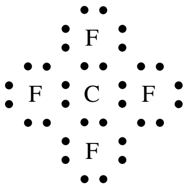
a 2 bonding electrons, 6 non-bonding electrons



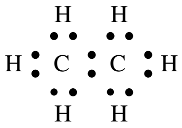
b 6 bonding electrons, 8 non-bonding electrons



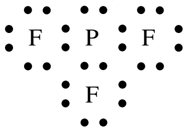
c 8 bonding electrons, 24 non-bonding electrons



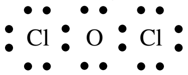
d 14 bonding electrons, 0 non-bonding electrons



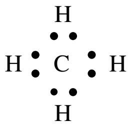
e 6 bonding electrons, 20 non-bonding electrons



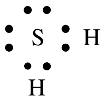
f 4 bonding electrons, 16 non-bonding electrons



g 8 bonding electrons



h 4 bonding electrons, 4 non-bonding electrons



16 **a** The bonds are similar in that they all involve the sharing of electron pairs between two atoms; that is, they are covalent bonds.

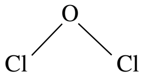
b They differ in the number of electron pairs shared: one pair (fluorine), two pairs (oxygen) and three pairs (nitrogen).

17 PCl_3—pyramidal, HOCl—V-shaped, $CHCl_3$—tetrahedral, HF—linear

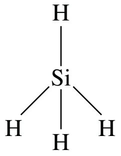
18 **a** non-polar

$S{=}C{=}S$

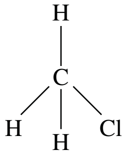
b polar



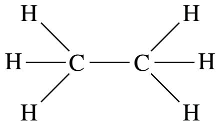
c non-polar



d polar



e non-polar



19 Si–O, H–Br, N–O, O–Cl, F–F

20

a SO_3	**i** non-polar	**ii** dispersion forces
b $SiCl_4$	**i** non-polar	**ii** dispersion forces
c CF_4	**i** non-polar	**ii** dispersion forces
d NF_3	**i** polar	**ii** dipole–dipole attraction
e CH_3NH_2	**i** polar	**ii** hydrogen bonding

21 **a** CCl_4

b CH_4 and CCl_4 are both non-polar, so their molecules are held together only by dispersion forces. CCl_4 is the larger of these two molecules and has more electrons, so the dispersion forces between CCl_4 molecules will be greater than those between CH_4 molecules. As there are stronger dispersion forces between molecules of CCl_4 than between molecules of CH_4, it takes more energy to vaporise CCl_4.

22 When sugar turns to a liquid, it is melting; the intermolecular bonds are being broken. When the liquid turns black and a gas is produced, the intramolecular bonds are being broken, allowing new substances to be produced.

23 Refer to Figure 5.5.5 on page 189 of the Student Book. Each carbon atom bonds covalently to three other carbon atoms in graphite to form layers. These layers consist of hexagonal rings of atoms, connected to one another. The fourth electron in each carbon atom is delocalised which explains its electrical conductivity. There are weak dispersion forces between the layers of graphite, allowing the layers to slide over each other, thus enabling it to act as a lubricant.

24 **a** Methane is an example of a molecular substance. It has strong, covalent intramolecular bonds and very weak intermolecular bonds. Diamond is an example of a covalent network structure. It has strong covalent bonds throughout its structure.

b Due to the weak dispersion forces between its molecules, methane will have extremely low boiling and melting points. If it were a solid it would be crystalline, brittle and soft. Due to its extended covalent network, diamond is extremely hard, does not exist as a liquid and has a very high sublimation point.

25 Diamond and silicon dioxide have a three-dimensional covalent network structure with strong covalent bonding. Each carbon and silicon atom is at the centre of a tetrahedron. In diamond, each carbon atom is bonded to four other carbon atoms. In silicon dioxide, each silicon atom is bonded to four oxygen atoms by covalent bonds. Each oxygen atom is bonded to two silicon atoms. In both substances, the strong covalent bonds extend throughout the network and account for its hardness and high melting points. All four valence electrons in each carbon and silicon atom are involved in the formation of covalent bonds, and there are no free electrons to carry an electric current.

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26 Metallic. Ag and Cu. Ag and Cu are both transition metals. Since metals have low electronegativities and ionisation energies, they easily lose their valence electrons. A metal lattice consists of metal cations surrounded by a sea of delocalised valence electrons.

Ionic. $CuCl_2$ and CaS. Ionic compound form by a transfer of electrons between metals and non-metals. Metals have low electronegativities and non-metals have high electronegativities. Metals are mainly found in groups 1 and 2 and in the transition block in the periodic table. Some metals are also found at the lower ends of the other groups. Non-metals are found in groups 15, 16 and 17 of the periodic table.

Polar covalent. HCl, H_2O and NH_3 have polar covalent bonds between their constituent atoms. Oxygen and nitrogen are more electronegative than hydrogen. The bonding electrons are not shared equally because they are attracted to the more electronegative atom. Polar covalent bonds are formed between the atoms of different non-metals found in groups 15, 16, and 17 of the periodic table. The asymmetrical nature of both molecules means they have molecular dipoles making them polar molecules.

Non-polar covalent. Cl_2. Since both atoms in Cl_2 are identical, there is no electronegativity difference between them and the bonding electrons are shared equally. Non-polar covalent bonds are formed between identical atoms of non-metals found in groups 15, 16 and 17 in the periodic table.

27 **a** both metallic and ionic lattices

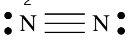
b both metallic and ionic lattices

c both metallic and ionic lattices

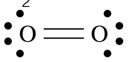
d ionic lattices only

e both metallic and ionic lattices

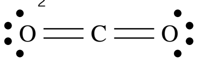
28 **a** **i** N_2



ii O_2



iii CO_2



b **i** NH_3, HCl, H_2O, $CHCl_3$

ii N_2, Cl_2, O_2, CH_4, CO_2, CCl_4

iii NH_3, H_2O

29 The intermolecular bonds that hold molecules together in covalent molecular substances are much weaker (100 times weaker) than the chemical bonds holding the atoms together in ionic, metallic and covalent network substances. As a result, it takes much less heat energy to break the intermolecular bonds holding covalent molecular solids and liquids together, and these substances have relatively low melting and boiling points.

30 **a**

Name	Nitrogen	Hydrogen bromide	Ammonia	Iron	Diamond	Potassium fluoride
formula	N_2	HBr	NH_3	Fe	C	KF
melting point (°C)	–210	–87	–78	1535	>3550 (sublimes)	858
bond type	dispersion forces	dipole–dipole forces (and dispersion forces)	hydrogen bonding (and dispersion forces)	metallic bonding	covalent bonding	ionic bonding

b Nitrogen (N_2). In solid nitrogen, weak dispersion forces exist between nitrogen molecules. These dispersion forces are a result of electrostatic attraction between the positive and negative ends of instantaneous dipoles in neighbouring molecules. The instantaneous dipoles are a result of random fluctuations in electron density around the two nitrogen atoms.

Hydrogen bromide (HBr). Dipole–dipole forces are the main cause of attraction between molecules when HBr is in the solid state. HBr molecules are polar because there is an electronegativity difference between the hydrogen and bromine atoms. Dipole–dipole attractions occur between polar HBr molecules: the partially positive charged hydrogen end of one HBr molecule is attracted to the partially negatively charged bromine end of another HBr molecule.

Ammonia (NH_3). Ammonia molecules are polar because nitrogen is more electronegative than hydrogen. There is an electrostatic attraction between the partially negative nitrogen atom on one ammonia molecule and the partially positive hydrogen atom on a nearby ammonia molecule. This special form of dipole-dipole bonding is called hydrogen bonding when hydrogen is bonded to fluorine, oxygen or nitrogen.

Iron (Fe). Metallic bonding is an electrostatic attraction between positively charged cations in the metal lattice and negatively charged delocalised electrons.

Diamond (C). The bonding in covalent network structures such as diamond is covalent. In diamond each carbon atom shares one valence electron with each of its four neighbouring carbon atoms. There is an electrostatic attraction between the shared electrons and the positive nuclei of the two carbon atoms involved in the formation of each bond.

Potassium fluoride (KF). An ionic bond is formed by the electrostatic attraction between the positively charged potassium cations and the negatively charged fluoride anions.

31 Student answers will vary, but will need to discuss how the salt and bicarbonate of soda are ionic substances that dissolve in water due to ions breaking away from the ionic lattice.

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Module 1 Review answers

Properties and structure of matter

Multiple choice

1 A. A gas has no fixed volume and no fixed shape. A liquid has a fixed volume and a solid has a fixed volume and shape.

2 B. Volume, elasticity and melting point are physical properties, because changes in volume, stretching the material or melting the material do not change a material's chemical properties. Corrosion resistance is a chemical property, because after corrosion a material has been chemically changed.

3 C. All atoms of the same element have the same number of protons in their nuclei.

4 B. All atoms of copper have the same atomic number. However, isotopes differ in the number of neutrons in the nucleus. The number of neutrons is the mass number minus the atomic number, so an atom of copper could have either 34 neutrons (63 – 29) or 36 neutrons (65 – 29).

5 D. In general, for the first series of transition metals, the 4*s*-subshell fills before the 3*d*-subshell. However, Cu and Cr are exceptions. Cr is more stable with a half-filled 4*s*-subshell and five electrons in its 3*d*-subshell.

6 C. A sulfur atom has 16 electrons so an S^{2-} ion has 18 electrons. D is incorrect because the 3*p*-subshell contains six electrons in the ground state.

7 C. Any *d*-subshell is made up of five orbitals. Each orbital can contain a maximum of two electrons. So a *d*-subshell can contain a maximum of 10 electrons in total.

8 B. Isotopes have the same atomic number but different mass numbers.

9 B. X has an atomic number of 19 so it has 19 protons in the nucleus. Its electronic configuration is $1s^22s^22p^63s^23p^6$ so it has 18 electrons. Therefore, as it has one more proton than electrons, it will have a positive charge.

10 A. Emission spectra involve the release of energy. When electrons at higher energy levels fall back to a lower energy level, the excess energy is emitted and can be observed as lines of light of particular frequencies on an emission spectrum.

11 B. Equating the atomic numbers on both sides of the nuclear equation gives $92 = x + 2$. Therefore, $x = 90$

12 C. Elements in the same group of the periodic table have similar properties because they have the same number of valence electrons. Only N and P belong to the same group.

13 D. Mass number is the sum of the number of protons and neutrons. So for this ion the mass number is 11 + 12 = 23. The ion has one less electrons than protons, so it has a charge of +1.

14 D. For main-group elements, the group number is related to the number of valence electrons. Elements in groups 1 and 2 have 1 and 2 outer-shell electrons, respectively. Elements in groups 13–18 have (group number – 10) outer-shell electrons.

15 B. The reactivity of metals increases down a group and decreases across a period.

16 A. The electrostatic forces of attraction between ions will determine the melting point of an ionic solid. These electrostatic forces of attraction are proportional to the charge on the ions. MgO is the only compound listed in which both ions are doubly charged (Mg^{2+} and O^{2-}) and these will exert the strongest attraction. So MgO will be expected to have the highest melting point.

17 C. Ionic bonding involves the transfer of electrons, not the sharing of electrons, from a metal to a non-metal.

18 C. The structure of HCN is H–C≡N. There is a triple bond between C and N. Therefore six electrons are shared between these two atoms.

19 A. Diatomic molecules, e.g. HF, can only be linear. CO_2 is also linear because it has two bonding regions (each a double bond). These repel and so are oriented as far from one another as possible, i.e. at 180°.

20 A. F_2 is a non-polar molecule and so only has dispersion forces between molecules. CH_3F has a polar C–F bond, so it has intermolecular dipole–dipole interactions. However, these are not as strong as hydrogen bonds in this case because the F is not bonded directly to the H atom as in HF.

Short answer

1 69.0% is potassium chlorate, so 31.0% must be aluminium. Of the 69.0% that composes potassium chlorate, 31.9% is potassium to contribute 22.0% (= 69.0% × 31.9%) to the total mass, 28.9% is chlorine to contribute 20.0% (= 69.0% × 29.0%) to the total mass, and 39.2% is oxygen to contribute 27.0% (= 69.0% × 39.2%) to the total mass. The total mass composition must be 100%: 22.0 + 20.0 + 27.0 + 31.0 = 100%.

2 **a** mass spectrometer

b $A_r = \frac{(37.30\times190.97)+(62.70\times192.97)}{100} = 192.2$

3 **a** Sodium has 11 electrons, which are distributed in three main shells or energy levels: two electrons in the first shell closest to the nucleus, eight electrons in the second shell and one in the third shell.

Shells are divided into subshells labelled *s*, *p*, *d* and *f*. Two of sodium's second-shell electrons are in the *s*-subshell and the remaining six are in the *p*-subshell.

s-, *p*-, *d*- and *f*-subshells are further divided into orbitals, which are regions of space in which electrons may be found. The *s*-subshell consists of only one orbital, the *p*-subshell has three. The six electrons of sodium in the 2*p*-subshell are distributed in pairs in the three *p*-orbitals.

b It represents an excited atom. A ground state electronic configuration would contain an electron in the 4*s*-subshell, because the 4*s*-subshell has a lower energy than the 3*d*-subshell.

4 **a** Chlorine is on the right side of the periodic table and sodium is on the left. Atomic radius decreases across a period because the increasing core charge pulls the outer-shell electrons closer to the nucleus, causing the volume of the atom to decrease.

b Fluorine is further to the right on the periodic table than lithium, and core charge increases from left to right across the periodic table. As core charge increases, the electrons are held closer to the nucleus, and more energy is required to remove the first one.

c Ba and Be are in the same group, with Be higher than Ba. Going down a group, atomic size increases, meaning the outer-shell electrons are further from the nucleus. Therefore, the outer electrons of Be are held more tightly and are less readily lost.

d The *s*-block elements have an *s*-subshell as their outer occupied electron subshell. The *s*-subshell can take one or two electrons, so the block only has two groups.

5 **a** K_3PO_4

b Al_2O_3

c $NaNO_3$

d Fe_2S_3

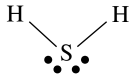
6 **a** Student answers will vary, for example: NaCl preserves food, $NaHCO_3$ is baking soda, NaF is in toothpaste to harden tooth enamel.

b **i** Although charged particles are present, they are held in fixed positions in the lattice and so cannot move to carry a current.

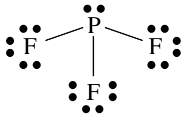
ii If a strong force is applied to a crystal of the compound, the layers of ions will move relative to one another, causing ions of like charge to be adjacent and hence repel. The crystal thus shatters.

iii A large amount of energy is needed to overcome the electrostatic attraction between oppositely charged ions.

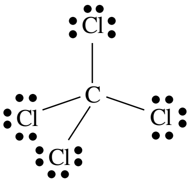
7 **a** V-shaped/bent and polar



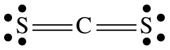
b triangular pyramid and polar



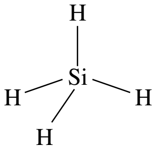
c Tetrahedral and non-polar.



d linear and non-polar



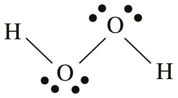
e tetrahedral and non-polar



8 **a** The bonds between the molecules of ice are hydrogen bonds. The bonds between the atoms within the water molecules are covalent bonds. Covalent bonds are much stronger than hydrogen bonds and so require much more energy, and thus a higher temperature, to break.

b In ice, each water molecule forms hydrogen bonds with four other water molecules. This arrangement is very open, causing ice to have a lower density than liquid water.

c Ethyne has the structure CHCH. The C atoms have a triple bond between them, each using three of their four valence electrons to form the triple bond. The fourth valence electron of each carbon atom forms a covalent bond with a hydrogen atom. There are no lone pairs and only two bonding regions on the carbon atoms. So these adopt a linear arrangement. Hydrogen peroxide has the following structure:



Each oxygen has six valence electrons. Two are involved in bonding, leaving two pairs of non-bonding electrons. So there are four regions where electrons are present. These assume a tetrahedral arrangement to minimise repulsion, so the molecule is not linear.

9 **a** HCl, HF, H_2O, H_2S

b HCl, H_2S

c HCl, HF, F_2, H_2, O_2

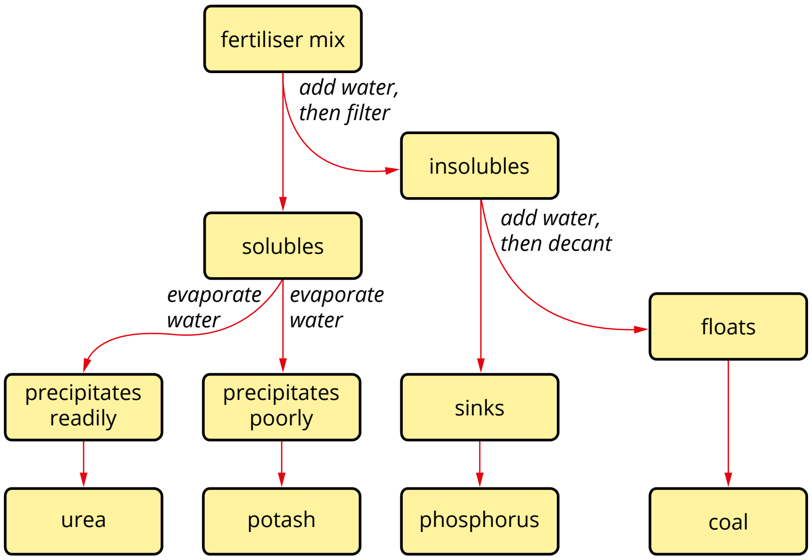
d H_2O, HF

e O_2

f H_2, as it is the smallest non-polar molecule in the list

Extended response

1 The fertiliser is a mixture of insoluble (rock, coal) and soluble (urea, potash) materials, so the first step is to filter out the insolubles. Mixing the fertiliser with water and passing it through a filter will separate the coal and phosphorus rock. Coal is less dense than water, so mixing the insolubles with water will make the coal float and the rock settle, which will allow the floating coal to be decanted. The coal and rock are now separated but are still mixed with water. A final filtration step isolates these from the water. Of the solubles, the urea derivatives are less soluble than potash. By evaporating the water, the urea derivatives will precipitate first, and can be separated by decanting or filtration. The final component, potash, can be separated from water by evaporating to dryness.



2 **a** $1s^2 2s^2 2p^6 3s^2$

b **i** 1 pm = 10^{-12} m, 1 nm = 10^{-9} m, so 140 pm = 0.140 nm

ii The Na atom would have a larger radius because there are fewer protons in the nucleus, so the attraction for the outer-shell electron would be weaker, pulling it less strongly towards the nucleus.

c **i** It is a lattice of positively charged magnesium ions surrounded by a 'sea' of valence electrons. The lattice is held together by the electrostatic attraction between electrons and cations.

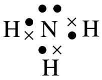
ii The electrons are not localised, but are free to move, so metals can conduct an electric current.

iii The model cannot explain the differences in melting points/densities/electrical conductivities/magnetism between metals.

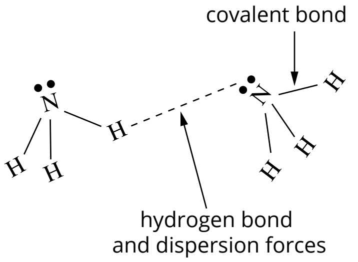
d any suitable example such as K, Na, Ca

e Transition elements have an incomplete *d*-subshell of electrons.

3 **a**



b The four electron pairs form a tetrahedral arrangement around the atom due to the repulsion of the electron pairs. The result is a pyramid-shaped molecule.



c **i** $N\equiv N$

ii $O=C=O$

d **i** N_2 has no polar bonds, so the intermolecular bonds are weak dispersion forces. Although the bonds within CO_2 are polar, the molecule overall is symmetrical, so the bond dipoles cancel and the overall molecule is non-polar. Thus, the only intermolecular forces are dispersion forces.

ii Nitrogen is a highly electronegative element, and consequently the bonds between the atoms of nitrogen and hydrogen are highly polarised. The ammonia molecule has a dipole because its shape is not symmetrical. There is an electrostatic attraction between the nitrogen atom of one ammonia molecule and the hydrogen atom of a nearby ammonia molecule. The attractions between these dipoles are known as hydrogen bonds.

4 **a** Allotropes are different physical forms of the same element.

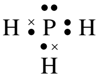
b Both consist of carbon atoms covalently bonded to other carbon atoms.

c Diamond is a three-dimensional lattice in which each carbon atom is covalently bonded to four other carbon atoms in a tetrahedral configuration. Thus, strong bonding extends throughout the lattice. Graphite consists of layers of carbon atoms in which each atom is covalently bonded to three other carbon atoms, making strong layers. There are weaker dispersion forces between the layers. The one electron not involved in bonding is delocalised.

d The delocalised electrons in graphite are free to move and serve as current so, graphite is an excellent conductor. In diamond, each carbon atom is bonded with four other carbon atoms, so there are no free electrons.

e Because of the weak bonding between the layers of graphite, the layers can slide over one another and thus can slide onto a page.

5 **a**



b PH_3 is a polar molecule. It is asymmetrical, so it has an overall molecular dipole. (The P–H bond is not particularly polar, but the overall asymmetry of the molecule results in an overall molecular dipole.)

c In a PH_3 molecule, there are four pairs of electrons around the central P atom. These electron pairs adopt a tetrahedral geometry. Since there is one lone pair, the molecular geometry is trigonal pyramidal.

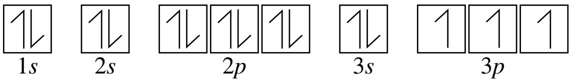
d $M_r(PH_3) = 30.97 + 3 \times 1.008 = 33.99$. $M_r(NH_3) = 14.01 + 3 \times 1.008 = 17.03$.

e Since H is bonded to N in ammonia, NH_3 molecules are able to form hydrogen bonds between molecules. Between PH_3 molecules there are dipole–dipole forces. Hydrogen bonding is stronger than dipole–dipole forces, so NH_3 has the higher melting point.

f $^{32}_{15}P$

g $^{32}_{15}P \rightarrow {}^{32}_{16}S + {}^{0}_{-1}e$

h



Chapter 6 Chemical reactions and stoichiometry

6.1 Writing chemical equations

Worked example: Try yourself 6.1.1

BALANCING A SIMPLE CHEMICAL EQUATION

Balance the following chemical equation: $H_2(g) + N_2(g) \rightarrow NH_3(g)$

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Underneath the arrow, write the letters
R | E | P
(Reactants, Elements, Products)</td><td>$H_2(g) + N_2(g) \rightarrow NH_3(g)$
R | E | P</td></tr>
<tr><td>List the elements under ‘E’ in the order in which they appear on the reactant side.</td><td>$H_2(g) + N_2(g) \rightarrow NH_3(g)$
R | E | P
| H |
| N |</td></tr>
<tr><td>Write the total number of each listed element on both the reactant and product sides.</td><td>$H_2(g) + N_2(g) \rightarrow NH_3(g)$
R | E | P
2 | H | 3
2 | N | 1
Neither the hydrogen nor the nitrogen is balanced. It is easier to balance the nitrogen first on the product side as it requires a simple doubling.</td></tr>
<tr><td>Balance by multiplying the coefficients until you have an equal number of each element on each side of the equation.</td><td>$H_2(g) + N_2(g) \rightarrow 2NH_3(g)$
R | E | P
2 | H | 3 × 2 = 6
2 | N | 1 × 2 = 2
On the products side:
Multiplying the number of nitrogen atoms by 2 requires doubling the number of NH_3 molecules. This means the number of hydrogen atoms increases to 6.
$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$
R | E | P
6 = 3 × 2 | H | 6
2 | N | 2
On the reactants side:
Multiplying the number of hydrogen atoms by 3 requires tripling the number of H_2 molecules.</td></tr>
<tr><td>Write the equation with the correct coefficients.</td><td>$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$</td></tr>
</table>

Worked example: Try yourself 6.1.2

BALANCING A CHEMICAL EQUATION WITH TRANSFER OF POLYATOMIC IONS

Balance the following chemical equation:

$AgNO_3(aq) + K_3PO_4\ (aq) \rightarrow Ag_3PO_4(s) + KNO_3(aq)$

Thinking	Working
Underneath the arrow, write the letters R | E | P (Reactants, Elements, Products) Note: Polyatomic ions are treated like ‘elements’ when balancing chemical equations using this method.	$AgNO_3(aq) + K_3PO_4(aq) \rightarrow Ag_3PO_4(s) + KNO_3(aq)$ R | E | P
List the elements and polyatomic ions that do not change from reactant to product side under ‘E’ in the order in which they appear on the reactant side.	$AgNO_3(aq) + K_3PO_4(aq) \rightarrow Ag_3PO_4(s) + KNO_3(aq)$ R | E | P | Ag^+ | | NO_3^- | | K^+ | | PO_4^{3-} |
Write the total number of each listed element/polyatomic ion on both the reactant and product sides.	$AgNO_3(aq) + K_3PO_4(aq) \rightarrow Ag_3PO_4(s) + KNO_3(aq)$ R | E | P 1 | Ag^+ | 3 1 | NO_3^- | 1 3 | K^+ | 1 1 | PO_4^{3-} | 1
Balance by multiplying the coefficients until you have an equal number of each element/polyatomic ion on each side of the equation.	$AgNO_3(aq) + K_3PO_4 \rightarrow Ag_3PO_4(s) + 3KNO_3(aq)$ R | E | P 1 | Ag^+ | 3 1 | NO_3^- | 1 × 3 = 3 3 | K^+ | 1 × 3 = 3 1 | PO_4^{3-} | 1 On the products side: Multiplying the number of K^+ by 3 requires tripling the number of KNO_3. This means the number of NO_3^- increases to 3. $3AgNO_3(aq) + K_3PO_4 \rightarrow Ag_3PO_4(s) + 3KNO_3(aq)$ R | E | P 3 = 3 × 1 | Ag^+ | 3 3 = 3 × 1 | NO_3^- | 3 3 | K^+ | 3 1 | PO_4^{3-} | 1 On the reactants side: Multiplying the number of Ag^+ by 3 requires tripling the number of $AgNO_3$. This means the number of NO_3^- increases to 3.
Write the equation with the correct coefficients.	$3AgNO_3(aq) + K_3PO_4(aq) \rightarrow Ag_3PO_4(s) + 3KNO_3(aq)$

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Worked example: Try yourself 6.1.3

BALANCING A MORE COMPLEX CHEMICAL EQUATION

Balance the following chemical equation:
$Ca_3(PO_4)_2(s) + SiO_2(s) + C(s) \rightarrow CaSiO_3(s) + CO(g) + P(s)$

Thinking	**Working**
In front of each chemical species, left to right, place the letters *a*, *b*, *c*, etc. until all chemical species have a letter.	$aCa_3(PO_4)_2(s) + bSiO_2(s) + cC(s) \rightarrow dCaSiO_3(s) + eCO(g) + fP(s)$
List the elements as they appear on the reactant side. Only list each element once, even if it appears a second (or third) time.	$aCa_3(PO_4)_2(s) + bSiO_2(s) + cC(s) \rightarrow dCaSiO_3(s) + eCO(g) + fP(s)$ Ca: P: Si: O: C:
Look at each element and determine an algebraic expression to balance each side.	$aCa_3(PO_4)_2(s) + bSiO_2(s) + cC(s) \rightarrow dCaSiO_3(s) + eCO(g) + fP(s)$ Ca: $3a = d$ P: $2a = f$ Si: $b = d$ O: $8a + 2b = 3d + e$ C: $c = e$
As there is more Ca on the reactant side, consider initially setting $a = 1$. Using this, solve for all values.	$aCa_3(PO_4)_2(s) + bSiO_2(s) + cC(s) \rightarrow dCaSiO_3(s) + eCO(g) + fP(s)$ Ca: $3(1) = d = 3$ P: $2(1) = f = 2$ Si: $b = d = 3$ O: $8(1) + 2(3) = 3(3) + e$; $e = 14 - 9 = 5$ C: $c = e = 5$
The example already has a whole number solution, so no multiplication is required.	$a = 1$ $b = 3$ $c = 5$ $d = 3$ $e = 5$ $f = 2$
Write down the balanced equation. (Remember that coefficients of ‘1’ do not need to be written.)	$Ca_3(PO_4)_2(s) + 3SiO_2(s) + 5C(s) \rightarrow 3CaSiO_3(s) + 5CO(g) + 2P(s)$

6.1 KEY QUESTIONS

1 C. The mass of a closed system cannot change over time. Mass cannot be created nor destroyed. Option A is incorrect as mass is always conserved in a closed system. Option B is incorrect as it is only true for nuclear fusion, whereby mass is converted directly into energy. For chemical reactions, mass is always conserved. Option D is incorrect because mass is conserved in a closed system. It does not gain or lose mass to the surroundings.

2 **a** $2NO(g) + 5H_2(g) \rightarrow 2NH_3(g) + 2H_2O(l)$
b $2C(s) + O_2(g) \rightarrow 2CO(g)$
c $5CuO(s) + 2NH_3(g) \rightarrow 5Cu(s) + 2NO(g) + 3H_2O(l)$

3 **a** $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$
b $2C_2H_2(g) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(g)$
c $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

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4 **a** Ag: 8 × 2 = 16; S: 8 × 1 = 8

b Mg: 2 × 3 = 6; P: 2 × 2 = 4; O: 2 × 4 × 2 = 16

c Na: 3 × 1 = 3; Cl: 3 × 1 = 3; O: 3 × 3 = 9

5 **a** Zinc metal is added to copper(II) sulphate solution. The zinc becomes coated with copper and colourless zinc sulphate solution is produced.

b Magnesium ribbon reacts with carbon dioxide gas. Black carbon (soot) and some magnesium oxide solid is produced.

c A potassium hydroxide solution reacts with a solution of carbonic acid (dihydrogen carbonate) to form liquid water and a potassium carbonate solution.

6.2 Problems involving conservation of mass

Worked example: Try yourself 6.2.1

CALCULATING THE MASS OF REACTANT CONSUMED OR PRODUCT FORMED

Assuming that all reactants are consumed in the following reaction, calculate the mass of reactant consumed. When 84.6 g of solid iodine (I_2) reacts with sufficient sodium (Na) metal to produce 99.9 g of sodium iodide (NaI), what mass of sodium is consumed?

Thinking	Working
Write a balanced chemical equation for the process.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$
List the initial amounts of all chemical species before the reaction takes place.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$ Initial: x g 84.6 g 0 g
List the changes in mass of the chemical reactant species. As all reactants are consumed in this reaction, all of the mass is converted into the product.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$ Initial: x g 84.6 g 0 g Change: $-x$ –84.6 g +99.9
The change in mass of the sodium will be equal to the difference between the mass of sodium iodide and the mass of the iodine.	$x = 99.9 - 84.6$ $= 15.3$ g

Worked example: Try yourself 6.2.2

CALCULATING THE MASS OF EXCESS REACTANT

One student measures 11.5 g of sodium (Na) and reacts it with 126.9 g of iodine (I_2), which is all consumed, producing 138.4 g of sodium iodide (NaI).

A second student measures 15.0 g of sodium and reacts it with the same amount of iodine and still produces 138.4 g of sodium iodide. What mass of sodium is in excess?

Thinking	Working
Write a balanced chemical equation for the process.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$
List the initial amounts of all chemical species before the reaction takes place.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$ Initial: 15.0 g 126.9 g 0 g
List the changes in mass of the chemical reactant species. All reactants are consumed in the first student's reaction. The masses of iodine in each student's reaction are the same and the amount of sodium is more than what was required initially.	$2Na(s) + I_2(s) \rightarrow 2NaI(s)$ Initial: 15.0 g 126.9 g 0 g Change: –11.5 –126.9 g +138.4
The mass of sodium remaining at the end of the second student's reaction will be the difference between the initial mass and the mass consumed in the reaction.	Excess sodium = 15.0 – 11.5 = 3.50 g

Worked example: Try yourself 6.2.3

CALCULATING MASS FROM VOLUME OF LIQUIDS

Hydrogen peroxide (H_2O_2) is a strong disinfectant and is also used in hair bleaching products. The production of hydrogen peroxide can be simplified to:

$$H_2(g) + O_2(g) \rightarrow H_2O_2(l)$$

0.04533 L of hydrogen peroxide is produced from the reaction between 4.0 g of hydrogen gas (H_2) and sufficient oxygen (O_2) to allow the reaction to be complete. Determine the mass of hydrogen peroxide produced, and hence, using the law of conservation of mass, determine the mass of oxygen gas consumed.

Take the density of hydrogen peroxide to be 1.50 g mL^{-1} at 25°C.

Thinking	**Working**
Write a balanced chemical equation for the process.	$H_2(g) + O_2(g) \rightarrow H_2O_2(l)$
Convert the volume of hydrogen peroxide from L to mL.	$0.04533 \times 1000 = 45.33$ mL
Using the density formula, determine the mass of hydrogen peroxide produced.	$d = \frac{m}{V}$ $m = d \times V$ $m(H_2O_2) = 1.50 \times 45.33$ $= 68.0$ g
List the initial amounts of all chemical species before the reaction takes place.	$H_2(g) + O_2(g) \rightarrow H_2O_2(l)$ Initial: 4.0 g x g 0 g
List the changes in mass of the chemical reactant species. As all reactants are consumed in this reaction, all of the mass is converted into the product.	Initial: 4.0 g x g 0 g Change: –4.0 $-x$ +68.0
The change in mass of the oxygen will be equal to the difference between the mass of the hydrogen peroxide and the mass of the hydrogen gas.	$x = 68.0 - 4.0$ $= 64.0$ g $m(H_2) = x = 64.0$ g

Worked example: Try yourself 6.2.4

CALCULATING MASS FROM VOLUME OF GASES

At 15°C and 1 atm of pressure, the density of butane (C_4H_{10}) is $2.48\,g\,L^{-1}$, the density of oxygen gas (O_2) is $1.36\,g\,L^{-1}$ and the density of carbon dioxide (CO_2) is $1.87\,g\,L^{-1}$.

Butane reacts with oxygen to form carbon dioxide and water according to the following equation:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$$

If 23.4 L of butane reacts completely with 152.1 L of oxygen to produce 90.0 g of water, what mass of carbon dioxide is produced?

Thinking	**Working**
Write a balanced chemical equation for the process.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$
Using the density formula, calculate the mass of the reactant species.	$d = \frac{m}{V}$ $m = d \times V$ $m(C_4H_{10}) = 2.48 \times 23.4$ $= 58.0\,g$ $m(O_2) = 1.36 \times 152.1$ $= 207\,g$
List the initial amounts of all chemical species before the reaction takes place and list the changes in mass of the chemical reactant species. As all reactants are consumed, all of the mass is converted into the products.	$\begin{array}{lllll} & 2C_4H_{10}(g) + & 13O_2(g) \rightarrow & 8CO_2(g) + & 10H_2O(g) \\ \text{Initial:} & 58.0\,g & 207\,g & 0 & 0 \\ \text{Change:} & -58.0 & -207 & +x & +90.0 \end{array}$
The change in mass of the carbon dioxide will equal the difference between the total mass of the reactants and the mass of the water vapour produced, and thus will equal the mass of carbon dioxide produced.	$x = (58.0 + 207) - 90.0$ $= 175\,g$ $m(CO_2) = x = 175\,g$

6.2 KEY QUESTIONS

1 **a** true

b False. Density is a measure of the mass of a substance divided by the volume that it occupies.

c true

d False. Volume is not always conserved in a chemical reaction.

2 **a** $m(MgO) = 5.00 + 3.29 = 8.29\,g$

b $m(O_2) = 40.3 - 24.3 = 16.0\,g$

c $m(Mg) = 30.35 - 12.05 = 18.30\,g$

3 **a** Reactants: 200.0 + 67.6 = 267.6 g. Products: 127.8 + 139.8 g = 267.6 g. As the mass of the reactants and products is equal, this satisfies the law of conservation of mass.

b 100.0 – 67.6 = 32.4 g excess Al

c 200.0 – 165.0 = 35.0 g more Fe_2O_3 would be required.

d Following the law of conservation of mass, if all reactant quantities are doubled and all reactants are consumed in the chemical reaction, then the total mass of the products will also double.

4 **a** $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$m(CO_2) = 10.00 - 5.72 = 4.28\,g$

b $V = \frac{m}{d} = \frac{4.28}{1.81} = 2.36\,L$

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CHAPTER 6 REVIEW

1 D. The total mass before and after a reaction takes place in a closed system is the same

Options A, B and C are incorrect as mass is constant in a closed system.

2 D. The other equations would be balanced as follows:

A. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$

B. $3F_2(g) + 2P(s) \rightarrow 2PF_3(g)$

C. $2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$

3 B.

A. $CuNO_3(aq) + Ag(s) \rightarrow Cu(s) + AgNO_3(aq)$. Not copper metal reacting.

C. $2Cu(s) + AgNO_3(aq) \rightarrow 2CuNO_3(aq) + Ag(s)$. Nitrate not balanced and incorrect empirical formula of copper (II) nitrate.

D. $Cu(s) + Ag(NO_3)_2(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$. Silver not balanced. Incorrect formula for silver nitrate.

4 **a** $4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$

b $2Al(OH)_3(s) \rightarrow Al_2O_3(s) + 3H_2O(l)$

c $10Rb(s) + 2RbNO_3(aq) \rightarrow 6Rb_2O(s) + N_2(g)$

d $4Fe(OH)_2(aq) + O_2(g) \rightarrow 2Fe_2O_3(s) + 4H_2O(l)$

e $2NaOCl(aq) \rightarrow 2NaCl(aq) + O_2(g)$

f $4H_2SiCl_2(s) + 4H_2O(l) \rightarrow H_8Si_4O_4(aq) + 8HCl(aq)$

5 **a** $2HNO_3(aq) + Na_2CO_3(aq) \rightarrow 2NaNO_3(aq) + H_2O(l) + CO_2(g)$

b $H_2SO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$

c $Ba(NO_3)_2(aq) + H_2CO_3(aq) \rightarrow BaCO_3(s) + 2HNO_3(aq)$

6 **a** $3Cu(s) + 8HNO_3(aq) \rightarrow 3Cu(NO_3)_2(aq) + 2NO(g) + 4H_2O(l)$

b $8CO(g) + 17H_2(g) \rightarrow C_8H_{18}(l) + 8H_2O(l)$

c $(NH_4)_2Cr_2O_7(s) \rightarrow N_2(g) + 4H_2O(g) + Cr_2O_3(s)$

d $10P_2I_4(s) + 13P_4(s) + 128H_2O(l) \rightarrow 40PH_4I(aq) + 32H_3PO_4(aq)$

Full working for part **d**

$a\ P_2I_4(s) + b\ P_4(s) + c\ H_2O(l) \rightarrow d\ PH_4I(aq) + e\ H_3PO_4(aq)$
$a\ P_2I_4(s) + b\ P_4(s) + c\ H_2O(l) \rightarrow d\ PH_4I(aq) + e\ H_3PO_4(aq)$ P: $2a + 4b = d + e$ I: $4a = d$ H: $2c = 4d + 3e$ O: $c = 4e$
$a\ P_2I_4(s) + b\ P_4(s) + c\ H_2O(l) \rightarrow d\ PH_4I(aq) + e\ H_3PO_4(aq)$ $a = 1$ P: $2(1) + 4b = 4 + e;\ 2 + 4b = 4 + e;\ b = \frac{2+e}{4} = \frac{13}{10}$ I: $4(1) = d;\ d = 4$ H: $2c = 4(4) + 3e;\ 2c = 16 + 3e;\ 2(4e) = 16 + 3e;\ 5e = 16;\ e = \frac{16}{5}$ O: $c = 4e = \frac{16 \times 4}{5} = \frac{64}{5}$
10 is lowest common factor to make all coefficients whole numbers $a = 1 \times 10 = 10$ $b = \frac{13}{10} \times 10 = 13$ $c = \frac{64}{5} \times 10 = 128$ $d = 4 \times 10 = 40$ $e = \frac{16}{5} \times 10 = 32$

7

Reaction	Reactant(s)			Product(s)	
a	$2H_2(g)$ +	$O_2(g)$	→	$H_2O(l)$	
	3.4 g	27.2 g		30.6 g	
b	$CH_4(g)$ +	$2O_2(g)$	→	$CO_2(g)$ +	$2H_2O(g)$
	12.2 g	48.8 g		41.0 g	20.0 g
c	$2HgO(s)$		→	$2Hg(l)$ +	$O_2(g)$
	23.6 g			21.9 g	1.7 g

8 **a** 0.4 g

b 0.34 g

c 0.022 g

9 **a** A chemical change is said to have occurred when a new substance has been formed. The evidence to suggest this includes a colour change; heat given off or absorbed; light emitted; a smell, odour or gas produced; or a precipitate formed. In the case of vinegar and milk, a precipitate is formed.

b The atoms from the two substances are rearranged to form the new products. The number and type of atoms in a chemical reaction remains the same.

c Due to the conservation of mass, the mass of a closed system must remain constant, i.e. 30.0 + 25.0 = 55.0 g products.

10 **a** $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

b 10.0 + 4.0 = 14.0 g

c 10.0 – 4.0 = 6.0 g

11 2.00 – 1.84 = 0.16 g

12 **a** $m = d \text{ x } V = 1.05 \times 30 = 31.5$ g

b $V = \frac{m}{d} = \frac{10.0}{1.05} = 9.52$ mL

c $m = d \text{ x } V = 1.05 \times 2000 = 2100$ g or 2.10 kg

Hint: You need to convert L to mL, as density is given in g mL^{-1}.

13 **a** $V(NH_3) = \frac{m}{d} = \frac{50.0}{0.860} = 58.1$ L

b $m(NH_3) = d \times V = 0.860 \times 15.0 = 12.9$ g

c $V(NH_3) = \frac{m}{d} = \frac{7500}{0.860} = 8.72 \times 10^3$ L

Hint: You need to convert kg to g, as density is given in g mL^{-1}.

14 **a** As temperature increases, particles have more kinetic energy and spread further apart, thus reducing the density of that substance.

b As pressure increases, particles are forced closer together, thus the density increases.

15 **a** $C_2H_5OH(l) + C_3H_7COOH(l) \rightarrow C_6H_{12}O_2(l) + H_2O(l)$

b $m(C_2H_5OH) = d \times V = 0.789 \times 10.0 = 7.89$ g

c $m(C_3H_7COOH) = d \times V = 0.953 \times 15.8 = 15.1$ g

d mass of reactants = 7.89 + 15.1 = 23.0 g

$m(C_6H_{12}O_2) – m(H_2O) = 23.0 – 3.09 = 19.9$ g

e $V(C_6H_{12}O_2) = \frac{m}{d} = \frac{19.86}{0.879} = 22.6$ mL

16 **a** $2NH_4NO_3(s) \rightarrow 2N_2(g) + O_2(g) + 4H_2O(g)$

Note: As the question stipulates water vapour, it must be listed in gaseous state.

b $40.0 - 14.0 - 8.0 = 18.0\,g\ H_2O$

c $V(N_2) = \frac{m}{d} = \frac{14.0}{1.25} = 11.2\,L$

d $V(O_2) = \frac{m}{d} = \frac{8.0}{1.43} = 5.6\,L$

e $V(\text{total}) = 11.2 + 5.6 + 24.7 = 41.5\,L$

f As the volume goes from 23.1 mL to 41.5 L, it can be seen that the volume increases by more than 1000 times. The rapid expansion and high kinetic energy of the particles during the reaction accounts for the explosive properties of ammonium nitrate.

17 **a** Though the total number of particles has decreased in the reaction, the total mass of the products must equal the total mass of the reactants according to the law of conservation of mass.

b Gases do have mass, but it can be difficult to measure.

c Density is a measure of how close particles are, whereas viscosity is a measure of how easily a substance flows or how ‘thick’ it appears.

18 $2HgO(s) \rightarrow 2Hg(l) + O_2(g)$

a $m(HgO) = 2.00 + 0.16 = 2.16\,g$

b $m(Hg) = 4.60 - 0.37 = 4.23\,g$

19 **a** $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$

b The mass of the product (sodium oxide) produced is equal to the sum of the mass of the sodium and the mass of the oxygen reacting.

20 **a** reactant: copper(II) nitrate; products: oxygen gas, copper(II) oxide, nitrogen dioxide gas

b $2Cu(NO_3)_2(s) \rightarrow 2CuO(s) + 4NO_2(g) + O_2(g)$

21 It is not a closed system and it is likely that some of the carbon dioxide gas evolved from the reaction has escaped from the beaker, resulting in the apparent mass loss in the reaction.

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Chapter 7 The mole concept

7.1 Introducing the mole

Worked example: Try yourself 7.1.1

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of molecules in 1.6 mol of carbon dioxide (CO_2).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of carbon dioxide molecules is the unknown, so: $N(CO_2) = ?$ $n(CO_2) = 1.6$ mol $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(CO_2) = n \times N_A$
Substitute in data and solve.	$N(CO_2) = n \times N_A$ $= 1.6 \times 6.022 \times 10^{23}$ $= 9.6 \times 10^{23}$ molecules

Worked example: Try yourself 7.1.2

CALCULATING THE NUMBER OF ATOMS

Calculate the number of hydrogen atoms in 0.35 mol of methane (CH_4).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of hydrogen atoms is the unknown, so: $N(H) = ?$ $n(CH_4) = 0.35$ mol $N_A = 6.022 \times 10^{23}$
Calculate the amount, in mol, of hydrogen atoms from the amount of methane and the molecular formula.	$n(H) = n(CH_4) \times 4$ $= 0.35 \times 4$ $= 1.4$ mol
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ so $N(H) = n \times N_A$
Substitute in data and solve.	$N(H) = n \times N_A$ $= 1.4 \times 6.022 \times 10^{23}$ $= 8.4 \times 10^{23}$ atoms

Worked example: Try yourself 7.1.3

CALCULATING THE NUMBER OF MOLES, GIVEN THE NUMBER OF PARTICLES

Calculate the amount, in mol, of magnesium atoms represented by 8.1×10^{20} magnesium atoms.

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of mol of magnesium atoms is the unknown, so: $n(\text{Mg}) = ?$ $N(\text{Mg}) = 8.1 \times 10^{20}$ atoms $N_A = 6.022 \times 10^{23}$
Rearrange the formula to make the unknown the subject.	$n = \frac{N}{N_A}$ n is the unknown, so rearrangement is not required
Substitute in data and solve.	$n(\text{Mg}) = \frac{N}{N_A}$ $= \frac{8.1 \times 10^{20}}{6.022 \times 10^{23}}$ $= 0.0013$ mol

Worked example: Try yourself 7.1.4

CALCULATING THE NUMBER OF MOLES OF ATOMS, GIVEN THE NUMBER OF MOLES OF MOLECULES

Calculate the amount, in mol, of hydrogen atoms in 0.75 mol of water (H_2O).

Thinking	Working
List the data given in the question next to the appropriate symbol. Include units.	The number of mol of hydrogen atoms is the unknown, so: $n(\text{H}) = ?$ $n(H_2O) = 0.75$ mol
Calculate the amount, in mol, of hydrogen atoms from the amount of water and the molecular formula.	$n(\text{H}) = n(H_2O) \times 2$ $= 0.75 \times 2$ $= 1.5$ mol

7.1 KEY QUESTIONS

1 Number of particles = amount (mol) × N_A

- **a** $N(\text{Na}) = 2.0 \times 6.022 \times 10^{23}$
  $= 1.2 \times 10^{24}$ atoms
- **b** $N(N_2) = 0.10 \times 6.022 \times 10^{23}$
  $= 6.0 \times 10^{22}$ molecules
- **c** $N(\text{C}) = 20.0 \times 6.022 \times 10^{23}$
  $= 1.20 \times 10^{25}$ atoms
- **d** $N(H_2O) = 4.2 \times 6.022 \times 10^{23}$
  $= 2.5 \times 10^{24}$ molecules
- **e** $N(\text{Fe}) = 1.0 \times 10^{-2} \times 6.022 \times 10^{23}$
  $= 6.0 \times 10^{21}$ atoms
- **f** $N(CO_2) = 4.62 \times 10^{-5} \times 6.022 \times 10^{23}$
  $= 2.78 \times 10^{19}$ molecules

2 Remember: Avogadro's constant, N_A, is 6.022×10^{23}.

Amount (mol) = $\frac{\text{number of particles}}{N_A}$

a $n(H_2O \text{ molecules}) = \frac{3.0 \times 10^{23}}{N_A}$

$= 0.50 \text{ mol}$

b $n(\text{Ne atoms}) = \frac{1.5 \times 10^{23}}{N_A}$

$= 0.25 \text{ mol}$

c $n(\text{Fe atoms}) = \frac{4.2 \times 10^{25}}{N_A}$

$= 70 \text{ mol}$

d $n(C_2H_5OH \text{ molecules}) = \frac{4.2 \times 10^{25}}{N_A}$

$= 70 \text{ mol}$

3 Remember: Avogadro's constant, N_A, is 6.022×10^{23}.

Amount (mol) = $\frac{\text{number of particles}}{N_A}$

a $n(\text{Na atoms}) = \frac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4} \text{ mol}$

b $n(\text{Al atoms}) = \frac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4} \text{ mol}$

c $n(Cl_2 \text{ molecules}) = \frac{1.0 \times 10^{20}}{N_A}$

$= 1.7 \times 10^{-4} \text{ mol}$

4 **a** Each Cl_2 molecule has 2 Cl atoms.

$\therefore n(\text{Cl atoms}) = 0.4 \times 2$

$= 0.8 \text{ mol}$

b Each methane molecule has 4 H atoms.

$\therefore n(\text{H atoms}) = 4 \times 1.2$

$= 4.8 \text{ mol}$

c Each ethane molecule has 6 H atoms.

$\therefore n(\text{H atoms}) = 6 \times 0.12$

$= 0.72 \text{ mol}$

d Each sulfate ion has 4 O atoms.

$\therefore n(\text{O atoms}) = 4 \times 1.5$

$= 6.0 \text{ mol}$

7.2 Molar mass

Worked example: Try yourself 7.2.1

CALCULATING THE MASS OF A SUBSTANCE

Calculate the mass of 4.68 mol of sodium carbonate (Na_2CO_3).

Thinking	Working
List the data given to you in the question. Remember that whenever you are given a molecular formula, you can calculate the molar mass.	$m(Na_2CO_3) = ?$ $n(Na_2CO_3) = 4.68\,mol$ $M(Na_2CO_3) = (2 \times 22.99) + 12.01 + (3 \times 16.00)$ $= 105.99\,g\,mol^{-1}$
Calculate the mass of sodium carbonate using: $n = \frac{m}{M}$	$n = \frac{m}{M}$, so $m = n \times M$ $m(Na_2CO_3) = 4.68 \times 105.99$ $= 496\,g$

Worked example: Try yourself 7.2.2

CALCULATING THE NUMBER OF MOLECULES

Calculate the number of sucrose molecules in a teaspoon (4.2 g) of sucrose ($C_{12}H_{22}O_{11}$).

Thinking	Working
List the data given to you in the question. Convert mass to grams if required. Remember that whenever you are given a molecular formula, you can calculate the molar mass.	The number of sucrose molecules is the unknown, so: $N(C_{12}H_{22}O_{11}) = ?$ $M(C_{12}H_{22}O_{11}) = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00)$ $= 342.30\,g\,mol^{-1}$ $m(C_{12}H_{22}O_{11}) = 4.2\,g$
Calculate the amount, in mol, of $C_{12}H_{22}O_{11}$, using: $n = \frac{m}{M}$	$n = \frac{N}{N_A}$ $= \frac{4.2}{342.30}$ $= 0.012\,mol$
Calculate the number of $C_{12}H_{22}O_{11}$ molecules using: $n = \frac{N}{N_A}$	$n = \frac{N}{N_A}$, so $N = n \times N_A$ $N(C_{12}H_{22}O_{11}) = 0.012 \times 6.022 \times 10^{23}$ $= 7.4 \times 10^{21}$

7.2 KEY QUESTIONS

1 **a** $M = 2 \times A_r(N)$
$= 2 \times 14.01$
$= 28.02\,g\,mol^{-1}$

b $M = A_r(N) + 3 \times A_r(H)$
$= 14.01 + 3 \times 1.008$
$= 17.03\,g\,mol^{-1}$

c $M = 2 \times A_r(H) + A_r(S) + 4 \times A_r(O)$
$= 2 \times 1.008 + 32.07 + 4 \times 16.00$
$= 98.09\,g\,mol^{-1}$

d $M = A_r(Fe) + 3 \times A_r(N) + 9 \times A_r(O)$
$= 55.85 + 3 \times 14.01 + 9 \times 16.00$
$= 241.88\,g\,mol^{-1}$

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e $M = 2 \times A_r(\text{C}) + 4 \times A_r(\text{H}) + 2 \times A_r(\text{O})$
$= 2 \times 12.01 + 4 \times 1.008 + 2 \times 16.00$
$= 60.05\,\text{g}\,\text{mol}^{-1}$

f $M = A_r(\text{S})$
$= 32.07\,\text{g}\,\text{mol}^{-1}$

g $M = 6 \times A_r(\text{C}) + 8 \times A_r(\text{H}) + 6 \times A_r(\text{O})$
$= 6 \times 12.01 + 8 \times 1.008 + 6 \times 16.00$
$= 176.12\,\text{g}\,\text{mol}^{-1}$

h $M = A_r(\text{Cu}) + A_r(\text{S}) + 4 \times A_r(\text{O}) + 5 \times M_r(\text{H}_2\text{O})$
$= 63.55 + 32.07 + 4 \times 16.00 + 5 \times 18.02$
$= 249.72\,\text{g}\,\text{mol}^{-1}$

2 It is useful to remember the formula $m = n \times M$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Use a periodic table to work out the molar masses.

a $m(\text{Na atoms}) = 1.0 \times 22.99$
$= 23\,\text{g}$

b $m(\text{O}_2) = 2.0 \times 32.00$
$= 64\,\text{g}$

c $m(\text{CH}_4) = 0.10 \times 16.04$
$= 1.6\,\text{g}$

d $m(\text{Al}_2\text{O}_3) = 0.25 \times 101.96$
$= 25\,\text{g}$

3 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Use a periodic table to work out the molar masses.

a $n(\text{H}_2) = \frac{5.0}{2.016} = 2.5\,\text{mol}$

b $n(\text{H atoms}) = \frac{5.0}{1.008} = 5.0\,\text{mol}$

c $n(\text{Al atoms}) = \frac{2.7}{26.98} = 0.10\,\text{mol}$

d $n(\text{CH}_4) = \frac{0.40}{16.04} = 0.025\,\text{mol}$

e $n(\text{O}_2) = \frac{0.10}{32.00} = 0.0031\,\text{mol}$

f $n(\text{O atoms}) = \frac{0.10}{16.00} = 0.0063\,\text{mol}$

g $n(\text{P}_4) = \frac{1.2 \times 10^{-3}}{123.88} = 9.7 \times 10^{-6}\,\text{mol}$

h $n(\text{P atoms}) = \frac{1.2 \times 10^{-3}}{30.97} = 3.9 \times 10^{-5}\,\text{mol}$

4 Use the formulae: Number of particles $= n \times N_A$, where $N_A = 6.022 \times 10^{23}$ and $n = \frac{m}{M}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Use a periodic table to work out the molar masses.

a $n(\text{Na}) = \frac{23}{22.99} = 1.0\,\text{mol}$

Number of Na atoms $= 1.0 \times 6.022 \times 10^{23} = 6.0 \times 10^{23}$ atoms

b $n(\text{Ar}) = \frac{4.0}{39.95} = 0.10\,\text{mol}$

Number of Ar atoms $= 0.10 \times 6.022 \times 10^{23} = 6.0 \times 10^{22}$ atoms

c $n(\text{Mg}) = \frac{0.243}{24.31} = 0.0100\,\text{mol}$

Number of Mg atoms $= 0.0100 \times 6.022 \times 10^{23} = 6.02 \times 10^{21}$ atoms

d $n(\text{Au}) = \frac{10.0}{197.0} = 0.0508\,\text{mol}$

Number of Au atoms $= 0.0508 \times 6.022 \times 10^{23} = 3.06 \times 10^{22}$ atoms

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5 Use the formulae: Number of particles = $n \times N_A$, where $N_A = 6.022 \times 10^{23}$ and $n = \frac{m}{M}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Use a periodic table to work out the molar masses.

a **i** $n(O_2) = \frac{16}{32.00} = 0.50\,\text{mol}$

number of O_2 molecules = $0.50 \times 6.022 \times 10^{23} = 3.0 \times 10^{23}$ molecules

ii $n(N_2) = \frac{2.8}{28.02} = 0.10\,\text{mol}$

number of N_2 molecules = $0.10 \times 6.022 \times 10^{23} = 6.0 \times 10^{22}$ molecules

b $n(SO_2) = \frac{3.2}{64.07} = 0.050\,\text{mol}$

number of SO_2 molecules = $0.050 \times 6.022 \times 10^{23} = 3.0 \times 10^{22}$ molecules
Each molecule contains 2 oxygen atoms.
So, number of oxygen atoms = 6.0×10^{22} atoms.

c $n(NH_3) = \frac{288}{17.03} = 16.9\,\text{mol}$

number of NH_3 molecules = $16.9 \times 6.022 \times 10^{23} = 1.02 \times 10^{25}$ molecules
Each molecule contains 4 atoms (1 of N and 3 of H).
So, total number of atoms = $4 \times 1.02 \times 10^{25} = 4.08 \times 10^{25}$ atoms.

7.3 Percentage composition and empirical formula

Worked example: Try yourself 7.3.1

CALCULATING PERCENTAGE COMPOSITION

Calculate the percentage by mass of nitrogen in ammonium nitrate (NH_4NO_3).

Thinking	Working
Find the molar mass of the compound.	$M(NH_4NO_3) = (2 \times 14.01) + (4 \times 1.008) + (3 \times 16.00)$ $= 80.05\,\text{g mol}^{-1}$
Find the total mass of the element in 1 mol of the compound.	mass of N in 1 mol = $2 \times M(N)$ $= 2 \times 14.01$ $= 28.02\,\text{g}$
Find the percentage by mass of the element in the compound.	% by mass of N in NH_4NO_3 $= \frac{\text{mass of N in mol of } NH_4NO_3}{\text{molar mass of } NH_4NO_3} \times 100$ $= \frac{28.02}{80.05} \times 100$ $= 35.00\%$

Worked example: Try yourself 7.3.2

DETERMINING THE EMPIRICAL FORMULA

0.50 g of magnesium is heated and allowed to completely react with chlorine. 1.96 g of white powder is formed. Determine the empirical formula of the compound.

Thinking	Working
Write down the mass, in g, of all elements present in the compound. If masses are given as percentages, assume that the sample weighs 100 g. The percentages then become masses in grams.	$m(\text{Mg}) = 0.50\,\text{g}$ and $m(\text{Cl}) = 1.96 - 0.50 = 1.46\,\text{g}$
Calculate the amount, in mol, of each element in the compound using: $n = \frac{m}{M}$	$n(\text{Mg}) = \frac{0.50}{24.31}$ $= 0.0206\,\text{mol}$ $n(\text{Cl}) = \frac{1.46}{35.45}$ $= 0.0411\,\text{mol}$
Simplify the ratio by dividing each number of mol by the smallest number of mol calculated in the previous step. This gives you a ratio of the number of atoms of each element.	$\frac{0.0206}{0.0206}$ $= 1$ $\frac{0.0411}{0.0206}$ $= 2$
Find the simplest whole number ratio.	1:2
Write the empirical formula.	MgCl_2

Worked example: Try yourself 7.3.3

DETERMINING THE MOLECULAR FORMULA

A compound has the empirical formula C_2H_5. The molar mass of this compound was determined to be 58.12 g mol^{-1}. What is the molecular formula of the compound?

Thinking	Working
Calculate the molar mass of one unit of the empirical formula.	Molar mass of a C_2H_5 unit $= 2 \times 12.01 + 5 \times 1.008$ $= 29.06\,\text{g mol}^{-1}$
Determine the number of empirical formula units in the molecular formula.	Number of C_2H_5 units $= \frac{58.12}{29.06}$ $= 2$
Determine the molecular formula of the compound.	Molecular formula $= 2 \times C_2H_5$ $= C_4H_{10}$

7.3 KEY QUESTIONS

1 **a** Percentage by mass of an element

$$= \frac{\text{mass of the element in 1 mol of compound}}{\text{molar mass of the compound}} \times 100$$

Use a periodic table to work out the molar masses. For example, $M(\text{Fe}_2\text{O}_3) = 159.70\,\text{g mol}^{-1}$.

$$\therefore \%(\text{Fe}) = \frac{2 \times 55.85}{159.70} \times 100$$

$$= 69.94\%$$

b $\%(\text{U}) = \frac{3 \times 238.0}{842.0} \times 100$

$$= 84.80\%$$

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c $\%(\text{N}) = \frac{14.01}{53.49} \times 100$

$= 26.19\%$

d $\%(\text{O}) = \frac{6 \times 16.00}{187.57} \times 100$

$= 51.18\%$

2 **a** The empirical formula provides the simplest whole number ratio of atoms in a compound. The number of mol of each atom is found by using $n = \frac{m}{M}$, where m is the mass in g, and M is the molar mass in g mol^{-1}.

	H	Cl
mass	2.74 g	97.26 g
molar mass	1.008 g mol^{-1}	35.45 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{2.74}{1.008} = 2.72$ mol	$n = \frac{97.26}{35.45} = 2.74$ mol
Divide all by the smallest amount.	$\frac{2.72}{2.72} = 1$	$\frac{2.74}{2.72} = 1.01$
Round off to whole numbers.	1	1

∴ empirical formula is HCl

b

	C	O
mass	42.9 g	57.1 g
molar mass	12.01 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{42.9}{12.01} = 3.57$ mol	$n = \frac{57.1}{16.00} = 3.57$ mol
Divide all by the smallest amount.	$\frac{3.57}{3.57} = 1$	$\frac{3.57}{3.57} = 1$
Round off to whole numbers.	1	1

∴ empirical formula is CO

c When 6.03 g is the mass of magnesium in 10.0 g of the compound, the mass of oxygen is (10.0 – 6.03) = 3.97 g.

	Mg	O
mass	6.03 g	3.97 g
molar mass	24.31 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{6.03}{24.31} = 0.248$ mol	$n = \frac{3.97}{16.00} = 0.248$ mol
Divide all by the smallest amount.	$\frac{0.248}{0.248} = 1$	$\frac{0.248}{0.248} = 1$
Round off to whole numbers.	1	1

∴ empirical formula is MgO

d When 2.4 g is the mass of carbon in 3.2 g of the hydrocarbon, the mass of hydrogen is (3.2 – 2.4) = 0.8 g.

	C	H
mass	2.4 g	0.8 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{2.4}{12.01} = 0.2$ mol	$n = \frac{0.8}{1.008} = 0.8$ mol
Divide all by the smallest amount.	$\frac{0.2}{0.2} = 1$	$\frac{0.8}{0.2} = 4$
Round off to whole numbers.	1	4

∴ empirical formula is CH_4

3 **a** The molecular formula is always a whole number multiple of the empirical formula. The empirical formula provides the simplest whole number ratio of atoms in a compound.

Molar mass of a CH unit (empirical formula) = 12.01 + 1.008 = 13.02 g mol^{-1}

Molar mass of the compound (molecular formula) = 78.11 g mol^{-1}

∴ number of CH units in one molecule = $\frac{78.11}{13.02} = 6$

∴ molecular formula is C_6H_6

b Molar mass of an HO unit (empirical formula) = 1.008 + 16.00 = 17.01 g mol^{-1}
Molar mass of the compound (molecular formula) = 34.02 g mol^{-1}

∴ number of OH units in one molecule = $\frac{34.02}{17.01} = 2$

∴ molecular formula is H_2O_2

c Molar mass of a CH_2O unit (empirical formula) = 12.01 + (2 × 1.008) + 16.00 = 30.03 g mol^{-1}

Molar mass of the compound (molecular formula) = 90.09 g mol^{-1}

∴ number of CH_2O units in one molecule = $\frac{90.09}{30.03} = 3$

∴ molecular formula is $C_3H_6O_3$

d Molar mass of an NO_2 unit (empirical formula) = 14.01 + (16.00 × 2) = 46.01 g mol^{-1}
Molar mass of the compound (molecular formula) = 46.01 g mol^{-1}

∴ number of NO_2 units in one molecule = $\frac{46.01}{46.01} = 1$

∴ molecular formula is NO_2

e Molar mass of a CH_2 unit (empirical formula) = 12.01 + (1.008 × 2) = 14.03 g mol^{-1}
Molar mass of the compound (molecular formula) = 154.29 g mol^{-1}

∴ number of CH_2 units in one molecule = $\frac{154.29}{14.03} = 11$

∴ molecular formula is $C_{11}H_{22}$

4 **a** A hydrocarbon contains only carbon and hydrogen.

∴ %H = 100 – 85.7 = 14.3%

	H	**C**
mass	14.3 g	85.7 g
molar mass	1.008 g mol^{-1}	12.01 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{14.3}{1.008} = 14.2$ mol	$n = \frac{85.7}{12.01} = 7.14$ mol
Divide all by the smallest amount.	$\frac{14.2}{7.14} = 1.98$	$\frac{7.14}{7.14} = 1$
Round off to whole numbers.	2	1

∴ empirical formula is CH_2

b Molar mass of a CH_2 unit (empirical formula) = 12.01 + (2 × 1.008) = 14.03 g mol^{-1}
Molar mass of the compound (molecular formula) = 70 g mol^{-1}

∴ number of CH_2 units in one molecule = $\frac{70}{14.03} = 5$
∴ molecular formula is C_5H_{10}

5 **a**

	C	H	O
mass	1.8 g	0.3 g	2.4 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{1.8}{12.01} = 0.15$ mol	$n = \frac{0.3}{1.008} = 0.29$ mol	$n = \frac{2.4}{16.00} = 0.15$ mol
Divide all by the smallest amount.	$\frac{0.15}{0.15} = 1$	$\frac{0.29}{0.15} = 1.9$	$\frac{0.15}{0.15} = 1$
Round off to whole numbers	1	2	1

∴ empirical formula is CH_2O

b Molar mass of a CH_2O unit (empirical formula) = 12.01 + (2 × 1.008) + 16.00 = 30.03 g mol^{-1}

Molar mass of the compound (molecular formula) = 180 g mol^{-1}

∴ number of CH_2O units in one molecule = $\frac{180}{30.03} = 6$

∴ molecular formula is $C_6H_{12}O_6$

7.4 Calculations based on the amount of a reactant or product

Worked example: Try yourself 7.4.1

USING MOLE RATIOS

In the breadmaking reaction, baker's yeast ferments table sugar (sucrose, $C_{12}H_{22}O_{11}$) to form carbon dioxide and ethanol (C_2H_5OH). 0.050 mol of CO_2 is needed to make a loaf of bread rise properly. If the chemical reaction is $C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 4CO_2(g) + 4CH_3CH_2OH(l)$, how much sucrose, in mol, is needed to form this amount of CO_2?

Thinking	**Working**
Write a balanced equation for the reaction.	$C_{12}H_{22}O_{11}(aq) + H_2O(l) \rightarrow 4CO_2(g) + 4CH_3CH_2OH(l)$,
Determine the number of mol of the 'known substance'. The known substance is the one you are provided information about in the question.	$n(CO_2) = 0.050$ mol
Find the mole ratio: mole ratio $= \frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$ The 'unknown substance' is the one for which you need to calculate the number of mol.	$\frac{\text{coefficient of } C_{12}H_{22}O_{11}}{\text{coefficient of } CO_2} = \frac{1}{4}$
Calculate the number of mol of the unknown substance using: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(C_{12}H_{22}O_{11}) = 0.050 \times \frac{1}{4}$ $= 0.013$ mol

Worked example: Try yourself 7.4.2

SOLVING MASS–MASS STOICHIOMETRIC PROBLEMS

A reaction between solutions of sodium sulfate and barium nitrate produces solid barium sulfate with a mass of 2.440 g. Aqueous sodium nitrate is also formed. Calculate the mass of sodium sulfate required to produce this amount of barium sulfate. The equation for this reaction is:

$$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$$

Thinking	**Working**
Write a balanced equation for the reaction.	$Na_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)$
Calculate the number of mol of the known substance (the precipitate): $n = \frac{m}{M}$	$n(BaSO_4) = \frac{2.440}{233.37}$ $= 0.01045$ mol
Calculate the mole ratio: mole ratio = $\frac{\text{coefficient of unknown chemical}}{\text{coefficient of known chemical}}$	mole ratio = $\frac{\text{coefficient of } Na_2SO_4}{\text{coefficient of } BaSO_4} = \frac{1}{1}$
Calculate the number of mol of the unknown substance: $n(\text{unknown}) = n(\text{known}) \times (\text{mole ratio})$	$n(Na_2SO_4) = 0.01045 \times \frac{1}{1}$ $= 0.01045$ mol
Calculate the mass of the unknown substance: $m = n(\text{unknown}) \times M$	$m(Na_2SO_4) = 0.01045 \times 142.05$ $= 1.486$ g

7.4 KEY QUESTIONS

1 $m(CO_2) = 120$ g

Working:

$n(CO_2) = \frac{120}{44.02} = 2.73$ mol

$n(C_6H_{12}O_6) = 2.73 \times \frac{1}{6} = 0.454$ mol

$m(C_6H_{12}O_6) = 0.454 \times 180.16 = 8.95$ g

2 **a** $n(Al_2(CO_3)_3) = \frac{4.68}{233.99} = 0.0200$ mol

b mole ratio = $\frac{n(Al(NO_3)_3)}{n(Al_2(CO_3)_3)} = \frac{2}{1}$

c $n(Al(NO_3)_3)$ required $= 0.0200 \times \frac{2}{1} = 0.0400$ mol

$m(Al(NO_3)_3) = 0.0400 \times 213.01 = 8.52$ g

3 $m(Si) = 45.9$ g

$n(Si) = \frac{45.9}{28.09} = 1.63$ mol

$\frac{n(SiCl_4)}{n(Si)} = \frac{1}{1}$

$n(SiCl_4) = 1.63 \times \frac{1}{1} = 1.63$ mol

$m(SiCl_4) = 1.63 \times 169.89 = 278$ g

4 $m(Hg(CH_3COO)_2) = 3.38$ g

Working:

$n(HgI_2) = \frac{4.82}{454.4} = 0.0106$ mol

mole ratio = $\frac{n(Hg(CH_3COO)_2)}{n(HgI_2)} = \frac{1}{1}$

$n(Hg(CH_3COO)_2)$ required $= 0.0106 \times \frac{1}{1} = 0.0106$ mol

$m(Hg(CH_3COO)_2) = 0.0106 \times 318.69 = 3.38$ g

7.5 Calculations based on the amounts of two reactants

Worked example: Try yourself 7.5.1

SOLVING MASS–MASS STOICHIOMETRY PROBLEMS WITH ONE REACTANT IN EXCESS

A space shuttle is loaded with 1.06×10^8 g of liquefied H_2 and 6.29×10^8 g of liquefied O_2, which are burnt as fuels during take-off. The reaction is:

$$2H_2(l) + O_2(l) \rightarrow 2H_2O(g)$$

(a) Which reactant is the limiting reactant?

Thinking	**Working**
Calculate the number of mol of each of the reactants in the equation using: $n = \frac{m}{M}$	Use the equation $n = \frac{m}{M}$. For O_2 : $n(O_2) = \frac{6.29 \times 10^8}{32.00} = 1.97 \times 10^7$ mol For H_2: $n(H_2) = \frac{1.06 \times 10^8}{2.016} = 5.26 \times 10^7$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows that 2 mol of H_2 react with 1 mol of O_2. So, to react all of the O_2 you will require: $\frac{2}{1} \times n(O_2)$ of H_2 $\frac{2}{1} \times 1.97 \times 10^7 = 3.93 \times 10^7$ mol As there is 5.26×10^7 mol available, the H_2 is in excess. So, O_2 is the limiting reactant. (It will be completely consumed.)

(b) What mass of water will be formed?

Thinking	**Working**
Find the mole ratio of the unknown substance to the limiting reactant from the equation coefficients: mole ratio = $\frac{\text{coefficient of unknown}}{\text{coefficient of limiting reactant}}$	From the equation coefficients: $\frac{\text{coefficient of } (H_2O)}{\text{coefficient of } O_2} = \frac{2}{1}$
Calculate the number of mol of the unknown substance using the number of mol of the limiting reactant: $n(\text{unknown}) = n(\text{limiting reactant}) \times (\text{mole ratio})$	$n(H_2O) = n(O_2) \times \frac{2}{1}$ $= 1.97 \times 10^7 \times \frac{2}{1}$ $= 3.93 \times 10^7$ mol
Calculate the mass of the unknown substance using: $m(\text{unknown}) = n(\text{unknown}) \times M$	Molar mass of H_2O = 18.02 g mol^{-1} $m(H_2O) = 3.93 \times 10^7 \times 18.02$ $= 7.08 \times 10^8$ g

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7.5 KEY QUESTIONS

1 **B** Calculate the number of mol of aluminium and oxygen.
D Refer to the balanced equation.
A Use mole ratios to determine which reactant is limiting.
E Calculate the number of mol of aluminium oxide that forms.
C Calculate the mass of aluminium oxide that forms.

2

Nitrogen molecules available	**Hydrogen molecules available**	**Ammonia molecules produced**	**Nitrogen molecules in excess**	**Hydrogen molecules in excess**
2	10	4	0	4
879	477	318	720	0
9 mol	6 mol	4 mol	7 mol	0 mol

Explanation

- Two nitrogen molecules will react with 2 × 3 = 6 hydrogen molecules to produce 2 × 2 = 4 ammonia molecules, so 10 − 6 = 4 hydrogen molecules will be in excess.
- All 477 hydrogen molecules will react with $477 \times \frac{1}{3} = 159$ nitrogen molecules to produce $477 \times \frac{2}{3} = 318$ ammonia molecules. 879 – 159 = 720 nitrogen molecules will be in excess.
- All 6 mol hydrogen molecules will react with $6 \times \frac{1}{3} = 2$ mol nitrogen molecules to produce $6 \times \frac{2}{3} = 4$ mol ammonia molecules. 9 – 2 = 7 mol nitrogen molecules will be in excess.

3 **a** $2Na(s) + Cl_2 \rightarrow 2NaCl(s)$

b $n(Na) = \frac{m}{M} = \frac{25.0}{22.99} = 1.09$ mol

$n(Cl_2) = \frac{m}{M} = \frac{50.0}{70.90} = 0.704$ mol

$\frac{n(Na)}{n(Cl_2)} = \frac{2}{1}$

$n(Na)$ required $= 0.704 \times \frac{2}{1} = 1.41$ mol
Sodium is the limiting reagent; use the number of mol of sodium to calculate $n(NaCl)$:

$\frac{n(NaCl)}{n(Na)} = \frac{1}{1}$

$n(NaCl) = n(Na) = 1.09$ mol
$m(NaCl) = n \times M = 1.09 \times 58.54 = 63.8$ g

4 **a** From the equation for the reaction:
$\frac{\text{coefficient of KI}}{\text{coefficient of } Pb(NO_3)_2} = \frac{2}{1} = 2$

2 mol of KI reacts with 1 mol of $Pb(NO_3)_2$
In this case, not all the $Pb(NO_3)_2$ can react and it will be in excess.
$n(Pb(NO_3)_2)$ reacted $= \frac{1}{2} \times n(KI) = \frac{1}{2} \times 1.0 = 0.50$ mol

So, $Pb(NO_3)_2$ is in excess by (1.0 – 0.50) = 0.5 mol.

b From the equation for the reaction:

$\frac{\text{coefficient of KI}}{\text{coefficient of } Pb(NO_3)_2} = \frac{2}{1} = 2$

2 mol of KI reacts with 1 mol of $Pb(NO_3)_2$.

In this case, not all the $Pb(NO_3)_2$ can react and it will be in excess.

$n(Pb(NO_3)_2)$ reacted $= \frac{1}{2} \times n(KI)$

$= \frac{1}{2} \times 0.5 = 0.25\,mol$

So, $Pb(NO_3)_2$ is in excess by $(2.0 - 0.25) = 1.8\,mol$

c $n(Pb(NO_3)_2) = \frac{1.00}{331.22} = 0.00302\,mol$

$n(KI) = \frac{1.50}{166.0} = 0.00904\,mol$

From the equation for the reaction:

$\frac{\text{coefficient of KI}}{\text{coefficient of }Pb(NO_3)_2} = \frac{2}{1} = 2$

2 mol of KI reacts with 1 mol of $Pb(NO_3)_2$.

In this case, not all of the KI can react and some of it will be in excess. This means that $Pb(NO_3)_2$ is the limiting reactant.

The equation shows that 1 mol of PbI_2 precipitate will form for every 1 mol of $Pb(NO_3)_2$ that reacts.

$n(PbI_2) = n(Pb(NO_3)_2) = 0.00302\,mol$

$m(PbI_2) = 0.00302 \times 461.0 = 1.39\,g$

CHAPTER 7 **REVIEW**

1 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Remember also that the number of particles in 1 mol, Avogadro's constant, $N_A = 6.022 \times 10^{23}$.

Use the formula: $n = \frac{\text{number of particles}}{N_A}$. Use a periodic table to work out the molar masses.

a $n(H_2O) = \frac{4.50 \times 10^{23}}{6.022 \times 10^{23}} = 0.75\,mol$

b $n(CH_4) = \frac{9.00 \times 10^{24}}{6.022 \times 10^{23}} = 15.0\,mol$

c $n(Cl_2) = \frac{2.3 \times 10^{28}}{6.022 \times 10^{23}} = 3.8 \times 10^4\,mol$

d $n(C_{12}H_{22}O_{11}) = \frac{1}{6.022 \times 10^{23}} = 1.7 \times 10^{-24}\,mol$

2 **a** **i** $N(NH_3) = n \times N_A = 1.45 \times 6.022 \times 10^{23} = 8.73 \times 10^{23}$ molecules

ii $N(\text{atoms}) = N(NH_3) \times 4 = 8.73 \times 10^{23} \times 4 = 3.49 \times 10^{24}$ atoms

iii 3 significant figures

b **i** $N(H_2S) = n \times N_A = 0.576 \times 6.022 \times 10^{23} = 3.47 \times 10^{23}$ molecules

ii $N(\text{atoms}) = N(H_2S) \times 3 = 3.47 \times 10^{23} \times 3 = 1.04 \times 10^{24}$ atoms

iii 3 significant figures

c **i** $N(HNO_3) = n \times N_A = 0.0153 \times 6.022 \times 10^{23} = 9.21 \times 10^{21}$ molecules

ii $N(\text{atoms}) = N(HNO_3) \times 5 = 9.21 \times 10^{21} \times 5 = 4.61 \times 10^{22}$ atoms

iii 3 significant figures

d **i** $N(C_{12}H_{22}O_{11}) = n \times N_A = 2.5 \times 6.022 \times 10^{23} = 1.5 \times 10^{24}$ molecules

ii $N(\text{atoms}) = N(C_{12}H_{22}O_{11}) \times 45 = 1.5 \times 10^{24} \times 45 = 6.8 \times 10^{25}$ atoms

iii 2 significant figures

3 The molar mass, M, has the same numerical value as the relative molecular mass, M_r, which is the sum of the relative atomic masses, A_r, of the elements in the compound. The molar mass, M, is the actual mass of 1 mol and so has the unit $g\,mol^{-1}$.

4 **a** $M(Fe) = 55.85\,g\,mol^{-1}$

b $M_r(H_2SO_4) = (2 \times 1.008) + 32.07 + (4 \times 16.00) = 98.09$
$\therefore M(H_2SO_4) = 98.09\,g\,mol^{-1}$

c $M_r(Na_2O) = (2 \times 22.99) + 16.00 = 61.98$
$\therefore M(Na_2O) = 61.98\,g\,mol^{-1}$

d $M_r(Zn(NO_3)_2) = 65.38 + (2 \times 14.01) + (6 \times 16.00) = 189.40$
$\therefore M(Zn(NO_3)_2) = 189.40\,g\,mol^{-1}$

e $M_r(H_2NCH_2COOH) = (5 \times 1.008) + 14.01 + (2 \times 12.01) + (2 \times 16.00) = 75.07$
$\therefore M(H_2NCH_2COOH) = 75.07\,g\,mol^{-1}$

f $M_r(Al_2(SO_4)_3) = (2 \times 26.98) + (3 \times 32.07) + (12 \times 16.00) = 342.17$
$\therefore M(Al_2(SO_4)_3) = 342.17\,g\,mol^{-1}$

g $M_r(FeCl_3.6H_2O) = 55.85 + (3 \times 35.45) + (12 \times 1.008) + (6 \times 16.00) = 270.30$
$\therefore M(FeCl_3.6H_2O) = 270.30\,g\,mol^{-1}$

5 **a** Mass of one atom = $\frac{\text{mass of 1 mol}}{\text{number of particles in 1 mol}} = \frac{\text{molar mass}}{N_A}$

Mass of one calcium atom = $\frac{40.08}{6.022 \times 10^{23}} = 6.656 \times 10^{-23}\,g$

b Mass of one water molecule = $\frac{18.02}{6.022 \times 10^{23}} = 2.992 \times 10^{-23}\,g$

c Mass of one CO_2 molecule = $\frac{44.02}{6.022 \times 10^{23}} = 7.310 \times 10^{-23}\,g$

6 **a** **i** $n(P_4) = \frac{m}{M} = \frac{4.2}{4 \times 30.97} = 0.032\,mol$
ii $N(P_4) = n \times N_A = 0.032 \times 6.022 \times 10^{23} = 1.9 \times 10^{22}$ molecules
iii Total number of atoms = $4 \times 1.9 \times 10^{22} = 7.8 \times 10^{22}$ atoms

b **i** $n(S_8) = \frac{m}{M} = \frac{75.0}{8 \times 32.07} = 0.292\,mol$
ii $N(S_8) = n \times N_A = 0.292 \times 6.022 \times 10^{23} = 1.76 \times 10^{23}$ molecules
iii Total number of atoms = $0.292 \times 8 \times 6.022 \times 10^{23} = 1.41 \times 10^{24}$ atoms

c **i** $n(HCl) = \frac{m}{M} = \frac{0.32}{1.008 + 35.45} = 0.0088\,mol$
ii $N(HCl) = n \times N_A = 0.0088 \times 6.022 \times 10^{23} = 5.3 \times 10^{21}$ molecules
iii Total number of atoms = $0.0088 \times 2 \times 6.022 \times 10^{23} = 1.1 \times 10^{22}$ atoms

d **i** $n(C_6H_{12}O_6) = \frac{m}{M} = \frac{2.2 \times 10^{-2}}{(6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00)} = 1.22 \times 10^{-4}\,mol$
ii $n(C_6H_{12}O_6) = n \times 6.022 \times 10^{23} = 7.3 \times 10^{19}$ molecules
iii Total number of atoms = $(1.22 \times 10^{-4}) \times 24 \times (6.022 \times 10^{23}) = 1.8 \times 10^{21}$ atoms

7 It is useful to remember the formula $n = \frac{m}{M}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass in $g\,mol^{-1}$. Use a periodic table to work out the molar masses.

a **i** $n(NaCl) = \frac{5.85}{58.44} = 0.100\,mol$

ii $n(Na^+) = n(NaCl) = 0.100\,mol$
$n(Cl^-) = n(NaCl) = 0.100\,mol$

b **i** $n(CaCl_2) = \frac{45.0}{110.98} = 0.405\,mol$

ii $n(Ca^{2+}) = n(CaCl_2) = 0.405\,mol$
$n(Cl^-) = 2 \times n(CaCl_2) = 2 \times 0.405 = 0.810\,mol$

c **i** $n(Fe_2(SO_4)_3) = \frac{1.68}{399.91} = 0.00420\,mol$

ii $n(Fe^{3+}) = 2 \times n(Fe_2(SO_4)_3) = 0.00840\,mol$
$n(SO_4^{2-}) = 3 \times n(Fe_2(SO_4)_3) = 0.0126\,mol$

8 It is useful to remember the formula $M = \frac{m}{n}$, where m is the mass in g, n the amount of substance in mol, and M the molar mass in g mol^{-1}.

a $M(\text{substance}) = \frac{80}{2.0} = 40\,\text{g mol}^{-1}$

All other parts follow the same process.

b $98\,\text{g mol}^{-1}$

c $44\,\text{g mol}^{-1}$

d $106\,\text{g mol}^{-1}$

9 It is useful to remember the formula $m = n \times M$, where m is the mass in g, n the amount of substance in mol, and M the molar mass. Use a periodic table to find the molar masses of iron and silver: $M = 55.85\,\text{g mol}^{-1}$ and $107.9\,\text{g mol}^{-1}$, respectively.

Remember also that the number of particles in 1 mol is N_A.

$m(\text{Cu}) = 100\,\text{g}$

$m(\text{Fe}) = 4.0 \times 55.85 = 223\,\text{g}$

$n(\text{Ag}) = \frac{1.2 \times 10^{24}}{N_A} = 2.0\,\text{mol}$

$\therefore m(\text{Ag}) = 2.0 \times 107.9 = 216\,\text{g}$

$\therefore$ B is correct: 4.0 mol of iron atoms has the greatest mass

10 It is useful to remember the formula $m = n \times M$, where m is the mass in g, n the amount of substance in mol, and M the molar mass in g mol^{-1}. Remember also that the number of particles in 1 mol is Avogadro's constant, $N_A = 6.022 \times 10^{23}$.

a $M(\text{antibiotic}) = 12\,500\,\text{g mol}^{-1} = 1.25 \times 10^{4}\,\text{g mol}^{-1}$

b $n(\text{antibiotic}) = \frac{2.0 \times 10^{-3}}{1.25 \times 10^{-4}} = 1.6 \times 10^{-7}\,\text{mol}$

c Number of molecules $= n \times N_A$

$= 1.6 \times 10^{-7} \times 6.022 \times 10^{23} = 9.6 \times 10^{16}$ molecules

11 Percentage by mass of an element $= \frac{\text{mass of the element in 1 mol of compound}}{\text{mass of 1 mol of the compound}} \times 100$

Use a periodic table to work out the molar masses. A useful check of these answers is provided by seeing that they add up to 100%, or somewhere close to that value.

a $M(\text{Al}) = 26.98\,\text{g mol}^{-1}$, $M(\text{O}) = 16.00\,\text{g mol}^{-1}$, $M(\text{Al}_2\text{O}_3) = 101.96\,\text{g mol}^{-1}$

$\%(\text{Al}) = \frac{2 \times 26.98}{101.96} \times 100 = 52.92\%$

$\%(\text{O}) = \frac{3 \times 16.00}{101.96} \times 100 = 47.07\%$

b Cu 65.1%; O 32.8%; H 2.1%

c Mg 12.0%; Cl 34.9%; H 5.9%; O 47.2%

d Fe 27.9%; S 24.1%; O 48.0%

e H 1.0%; Cl 35.3%; O 63.7%

12 The empirical formula provides the simplest whole number ratio of atoms in a compound. The amount of each atom is found by using $n = \frac{m}{M}$, where m is the mass in g and M is the molar mass in g mol^{-1}. Use a periodic table to find the molar masses of C, H, N and O.

	C	**H**	**N**	**O**
mass	32 g	6.7 g	18.7 g	42.6 g
molar mass	$12.01\ \text{g mol}^{-1}$	$1.008\ \text{g mol}^{-1}$	$14.01\ \text{g mol}^{-1}$	$16.00\ \text{g mol}^{-1}$
amount using $n = \frac{m}{M}$	$n = \frac{32}{12.01}$ $= 2.67\ \text{mol}$	$n = \frac{6.7}{1.008}$ $= 6.7\ \text{mol}$	$n = \frac{18.7}{14.01}$ $= 1.34\ \text{mol}$	$n = \frac{42.6}{16.00}$ $= 2.66\ \text{mol}$
Divide all by the smallest amount.	$\frac{2.67}{1.34} = 1.99$	$\frac{6.7}{1.34} = 5$	$\frac{1.34}{1.34} = 1$	$\frac{2.66}{1.34} = 1.99$
Round off to whole numbers.	2	5	1	2

∴ empirical formula is $C_2H_5NO_2$

13 **a** Step 1: Calculate mass of each element present in 100 g.

$m(\text{C}) = 52.2\ \text{g}$

$m(\text{O}) = 34.8\ \text{g}$

$m(\text{H}) = 100 - 52.2 - 34.8 = 13.0\ \text{g}$

Step 2: Calculate amount, in mol, of each element present.

$n(\text{C}) = \frac{m}{M} = \frac{52.2}{12.01} = 4.35\ \text{mol}$

$n(\text{O}) = \frac{34.8}{16.00} = 2.175\ \text{mol}$

$n(\text{H}) = \frac{13.0}{1.008} = 12.9\ \text{mol}$

Step 3: Convert to whole number ratios.

$n(\text{C}) = \frac{4.35}{2.175} = 2$

$n(\text{O}) = \frac{2.175}{2.175} = 1$

$n(\text{H}) = \frac{12.9}{2.175} = 6$

Step 4: Write as empirical formula: C_2H_6O

b Step 1: Calculate molar mass of one empirical formula unit.

$M = 2 \times 12.01 + 6 \times 1.008 + 16.00$

$= 46.07\ \text{g mol}^{-1}$

Step 2: Calculate molar mass of compound.

$M = \frac{m}{M} = \frac{100}{2.17} = 46.07\ \text{g mol}^{-1}$

Step 3: Calculate number of empirical formula units in compound.

no. of units $= \frac{46.07}{46.07} = 1$

Step 4: Write molecular formula: C_2H_6O

14

	C	**H**
mass	90.0 g	100 – 90.0 = 10.0 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{90.0}{12.01}$ = 7.49 mol	$n = \frac{10.0}{1.008}$ = 9.92 mol
Divide all by the smallest amount.	$\frac{7.49}{7.49}$ = 1	$\frac{9.92}{7.49}$ = 1.3
Round off to whole numbers.	1 × 3 = 3	1.3 × 3 = 4

∴ empirical formula is C_3H_4

15

	Ni	**O**
mass	3.370 g	4.286 – 3.370 = 0.916 g
molar mass	?	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{3.370}{M(\text{Ni})}$	$n = \frac{0.916}{16.00}$ = 0.0573 mol

$\frac{n(\text{Ni})}{n(\text{O})} = \frac{0.370}{M(\text{Ni})} = \frac{1}{0.0573}$

As the empirical formula is NiO, $\frac{0.370}{M(\text{Ni})} \times \frac{1}{0.0573} = \frac{1}{1}$

∴ $M(\text{Ni}) = \frac{0.370}{0.0573}$ = 58.9 g mol^{-1}

16

	W	**Cl**
mass	4.150 g	8.950 – 4.150 = 4.800 g
molar mass	?	35.45 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{4.150}{M(\text{W})}$	$n = \frac{4.800}{35.45}$ = 0.1354 mol

$\frac{n(\text{W})}{n(\text{Cl})} = \frac{4.150}{M(\text{W}) \times 0.1354}$

As the empirical formula is WCl_6, $\frac{4.150}{M(\text{W}) \times 0.1354} = \frac{1}{6}$

∴ $M(\text{W}) = 4.150 \times \frac{6}{0.1354}$ = 183.9 g mol^{-1}

17 a

	C	**H**
mass	82.75 g	17.25 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{82.75}{12.01}$ = 6.890 mol	$n = \frac{17.25}{1.008}$ = 17.25 mol
Divide all by the smallest amount.	$\frac{6.890}{6.890}$ = 1	$\frac{17.25}{6.890}$ = 2.5
Round off to whole numbers.	1 × 2 = 2	2.5 × 2 = 5

∴ empirical formula is C_2H_5
Molar mass of a C_2H_5 unit (empirical formula) = 29.06 g mol^{-1}.
Molar mass of the compound (molecular formula) = 58.12 g mol^{-1}.
∴ number of C_2H_5 units in one molecule = $\frac{58.12}{29.06}$ = 2
∴ molecular formula is C_4H_{10}

b

	P	**O**
mass	43.66 g	56.34 g
molar mass	30.97 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n(\text{P}) = \frac{43.66}{30.97} = 1.410$ mol	$n(\text{O}) = \frac{56.34}{16.00} = 3.521$ mol
Divide all by the smallest amount.	$\frac{1.410}{1.410} = 1$	$\frac{3.521}{1.410} = 2.497 \approx 2.5$
Round off to whole numbers.	1 × 2 = 2	2.5 × 2 = 5

∴ empirical formula is P_2O_5

Molar mass of a P_2O_5 unit (empirical formula) = 141.94 g mol^{-1}.

Molar mass of the compound (molecular formula) = 283.88 g mol^{-1}.

∴ number of P_2O_5 units in one molecule = $\frac{283.88}{141.94} = 2$

∴ molecular formula is P_4O_{10}

c

	C	**H**	**O**
mass	40.0 g	6.7 g	53.3 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n(\text{C}) = \frac{40.0}{12.017} = 3.3$ mol	$n(\text{H}) = \frac{6.7}{1.008} = 6.6$ mol	$n(\text{O}) = \frac{53.3}{16.00} = 3.3$ mol
Divide all by the smallest amount.	$\frac{3.3}{3.3} = 1$	$\frac{6.6}{3.3} = 1$	$\frac{3.3}{3.3} = 1.0$
Round off to whole numbers.	1	2	1

∴ empirical formula is CH_2O

Molar mass of a CH_2O unit (empirical formula) = 12.01 + (1.008 × 2) + 16.00 = 30.03 g mol^{-1}.

Molar mass of the compound (molecular formula) = 180.16 g mol^{-1}.

∴ number of CH_2O units in one molecule = $\frac{180.16}{30.03} = 6.0$

∴ molecular formula is $C_6H_{12}O_6$

d

	H	**S**	**O**
mass	0.164 g	5.25 g	9.18 g
molar mass	1.008 g mol^{-1}	32.07 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{0.164}{1.008} = 0.164$ mol	$n = \frac{5.25}{32.07} = 0.164$ mol	$n = \frac{9.18}{16.00} = 0.574$ mol
Divide all by the smallest amount.	$\frac{0.164}{0.164} = 1$	$\frac{0.164}{0.164} = 1$	$\frac{0.574}{0.164} = 3.5$
Round off to whole numbers.	1 × 2 = 2	1 × 2 = 2	3.5 × 2 = 7

∴ empirical formula is $H_2S_2O_7$

Molar mass of a $H_2S_2O_7$ unit (empirical formula) = (2 × 1.008) + (2 × 32.07) + (7 × 16.00) = 178.16 g mol^{-1}.

Molar mass of the compound (molecular formula) = 178.16 g mol^{-1}.

∴ number of $H_2S_2O_7$ units in one molecule = $\frac{178.16}{178.16} = 1$

∴ molecular formula is $H_2S_2O_7$

18 **a** The relative isotopic mass of an isotope is the mass of an atom of that isotope relative to the mass of an atom of ^{12}C, taken as 12 units exactly. For example, the relative isotopic mass of the lighter of the two chlorine isotopes is 34.969.

b The relative atomic mass of an element is the weighted average of the relative masses of the isotopes of the element on the ^{12}C scale. For example, the relative atomic mass of boron is 10.81.

c The relative molecular mass (M_r) of a compound is the mass of one molecule of that substance relative to the mass of a ^{12}C atom, which is 12 exactly. For example, the relative molecular mass of carbon dioxide is 44.01.

d Relative formula mass is calculated by taking the sum of the relative atomic masses of the elements in the formula. Relative formula mass (rather than relative molecular mass) is the appropriate term to use for ionic compounds, as they do not contain molecules. For example, the relative formula mass of sodium chloride is 58.44.

e The molar mass of an element is the mass of 1 mol of the element. It is equal to the relative atomic mass of the element expressed in grams. For example, the molar mass of magnesium is 24.31 g mol^{-1}. Note that relative atomic mass and molar mass of an element are numerically equal. However, relative atomic mass has no units because it is the mass of one atom of the element compared with the mass of one atom of the carbon-12 isotope. The molar mass of a compound is the mass of 1 mol of the compound. It is equal to the relative molecular or relative formula mass of the compound expressed in grams. For example, the molar mass of sodium chloride is 58.44 g mol^{-1}.

19 **a** Take 100 g of caffeine. The mass of oxygen can be found by subtraction

$= 100 - (49.48 + 5.15 + 28.87)$
$= 16.50$ g

	C	**H**	**N**	**O**
mass	49.48 g	5.15 g	28.87 g	16.50 g
molar mass	12.01 g mol^{-1}	1.008 g mol^{-1}	14.01 g mol^{-1}	16.00 g mol^{-1}
amount using $n = \frac{m}{M}$	$n = \frac{49.48}{12.01}$ $= 4.17$ mol	$n = \frac{5.15}{1.008}$ $= 5.10$ mol	$n = \frac{28.87}{14.01}$ $= 2.06$ mol	$n = \frac{16.50}{16.00}$ $= 1.03$ mol
Divide all by the smallest amount.	$\frac{4.17}{1.03} = 4.04$	$\frac{5.10}{1.03} = 5.0$	$\frac{2.06}{1.03} = 2.0$	$\frac{1.03}{1.03} = 1.0$
Round off to whole numbers.	4	5	2	1

$\therefore$ empirical formula is $C_4H_5N_2O$

b Molar mass of caffeine $= \frac{m}{n} = \frac{38.8}{0.200} = 194$ g mol^{-1}

c The molecule must contain a whole number of ($C_4H_5N_2O$) units.
The molar mass of a $C_4H_5N_2O$ unit is
$(4 \times 12.01) + (5 \times 1.008) + (2 \times 14.01) + 16.00 = 97.46$ g mol^{-1}.
If the compound has a molar mass of 194 g mol^{-1}, then the number of ($C_4H_5N_2O$) units in a molecule $= \frac{\text{molar mass of compound}}{\text{molar mass of one unit}}$

$= \frac{194}{97.46} = 2$

The molecular formula of caffeine is therefore 2 × ($C_4H_5N_2O$); that is, $C_8H_{10}N_4O_2$.

d $n(\text{caffeine}) = \frac{m}{M}$

$= \frac{1.00}{194.92}$

$= 5.15 \times 10^{-3}$ mol

e $N(\text{caffeine}) = n(\text{caffeine}) \times N_A$
$= 5.15 \times 10^{-3} \times 6.022 \times 10^{23} = 3.10 \times 10^{21}$ molecules of caffeine

f $N(\text{atoms in caffeine}) = N(\text{molecules}) \times 24 = 7.44 \times 10^{22}$ atoms

20 **a** D, F, E, A, C, B

b Step 1: Calculate mass of compound after reaction.
$m = 21.068 - 20.310 = 0.758$ g
Step 2: Calculate mass of oxygen.
$m = 0.758 - 0.542 = 0.216$ g
Step 3: Calculate amount, in mol, of oxygen.
$n = \frac{m}{M}$
$= \frac{0.126}{16.00} = 0.0135$ mol

Step 4: Use mole ratios to determine amount of metal.
Ratio is 1:1, so $n = 0.0135$ mol
Step 5: Calculate molar mass of metal.

$M = \frac{m}{n} = \frac{0.542}{0.135} = 40.08$ g mol^{-1}

Step 6: Fill in table.

	Metal	**Oxygen**
mass (g)	0.542	0.216
relative atomic mass	40.08	16.00
moles	0.0135 mol	0.0135 mol
ratio	1	1

c Refer to periodic table to identify metal as calcium.

21

$Ca(NO_3)_2$	**Na_3PO_4**	**$Ca_3(PO_4)_2$**	**$NaNO_3$**
27 mol	18	9.0	54
0.72	0.48 mol	0.24	0.44
0.54	0.36	0.18 mol	1.08
1.2	0.8	0.4	2.4 mol

22 $n(KOH) = \frac{2}{1} \times n(Fe(OH)_2)$

$n(FeSO_4) = \frac{2}{1} \times n(KOH)$

$n(KOH) = \frac{2}{1} \times n(K_2SO_4)$

$n(Fe(OH)_2) = \frac{1}{1} \times n(FeSO_4)$

23 $m(C_3H_8) = 50.2\text{ g}$

$n(C_3H_8) = \frac{50.2}{44.09}$

$= 1.14\text{ mol}$

$\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$

$n(CO_2) = \frac{3}{1} \times 1.14$

$= 3.42\text{ mol}$

$m(CO_2) = n \times M$

$= 3.42 \times 44.01$

$= 150\text{ g}$

24 $m(Li) = 8.22\text{ g}$

$n(Li) = \frac{8.22}{6.941}$

$= 1.18\text{ mol}$

$\frac{n(LiCl)}{n(Li)} = \frac{2}{2} = \frac{1}{1}$

$n(LiCl) = \frac{1}{1} \times 1.18$

$= 1.18\text{ mol}$

$m(LiCl) = n \times M$

$= 1.18 \times 42.39$

$= 50.0\text{ g}$

25

Carbon atoms available	**Oxygen molecules available**	**Carbon dioxide molecules produced**	**Carbon atoms in excess**	**Oxygen molecules in excess**
8	20	8	0	12
1000	3000	1000	0	2000
9 mol	6 mol	6 mol	3 mol	0 mol

26 **a** Oxygen is in excess.

b 1.5 mol of sodium oxide is produced.

$\frac{n(O_2)}{n(Na)} = \frac{1}{4}$

$n(O_2)$ required $= \frac{1}{4} \times 3.0 = 0.75\text{ mol}$

Since the $n(O_2)$ required is 0.75 mol, oxygen is in excess and sodium is the limiting reagent.

Next, calculate the number of mol of sodium oxide from the number of mol of sodium.

$\frac{n(Na_2O)}{n(Na)} = \frac{2}{4} = \frac{1}{2}$

$n(Na_2O) = 3.0 \times \frac{1}{2} = 1.5\,mol$

27 **a** $m(S) = 22.6\,g$

$n(H_2S) = \frac{16.0}{34.09} = 0.471\,mol$

$n(SO_2) = \frac{20.0}{64.07} = 0.313\,mol$

From the balanced equation, the number of mol of H_2S needs to be double that of SO_2. It is not; therefore, H_2S is the limiting reactant.

$n(S) = \frac{3}{2} \times 0.471 = 0.707\,mol$

$\therefore\ m(S) = n \times M = 0.707 \times 32.07 = 22.6\,g$

b $m(SO_2) = 4.93\,g$

This is calculated by following these steps:

$n(SO_2\ \text{reacting}) = \frac{1}{2} \times 0.471 = 0.236\,mol$

$n(SO_2\ \text{remaining}) = 0.313 - 0.236 = 0.077\,mol$

$\therefore\ m(SO_2) = n \times M = 0.077 \times 64.07 = 4.93\,g$

28 **a** $n(P_4O_6) = \frac{4.40}{219.88} = 0.0200\,mol$

$n(I_2) = \frac{3.00}{253.8} = 0.0118\,mol$

Use $n(P_4O_6)$ provided to find $n(I_2)$ needed to react completely with the P_4O_6. From the equation, 8 mol of I_2 react with 5 mol of P_4O_6.

$\frac{n(I_2)}{n(P_4O_6)} = \frac{8}{5}$

$n(I_2)\ \text{required} = \frac{8}{5} \times n(P_4O_6) = \frac{8}{5} \times 0.0200 = 0.0320\,mol$

$n(I_2)\ \text{available} = 0.0118\,mol$

So, there is insufficient I_2 to react with all the P_4O_6, so I_2 is the limiting reagent and P_4O_6 is in excess.

Use $n(I_2)$ provided to find $n(P_4O_6)$ needed to react completely. From the equation, 8 mol of I_2 react with 5 mol of P_4O_6.

$n(P_4O_6)\ \text{required} = \frac{8}{5} \times n(I_2) = \frac{8}{5} \times 0.0118 = 0.00738\,mol$

$n(P_4O_6)\ \text{available} = 0.0200\,mol$

So, P_4O_6 is in excess by 0.0200 – 0.00738 = 0.0126 mol

$\therefore\ m(P_4O_6)\ \text{in excess} = 0.0126 \times 219.88 = 2.77\,g$

b Use the amount of the limiting reagent, I_2, to determine the amount of P_2I_4 formed. From the equation, 8 mol of I_2 react to form 4 mol of P_2I_4.

$\frac{n(P_2I_4)}{n(I_2)} = \frac{4}{8} = \frac{1}{2}$

$n\left(P_2I_4\right) = \frac{1}{2} \times n\left(I_2\right) = \frac{1}{2} \times 0.0118 = 0.00591\,mol$

$\therefore\ m(P_2I_4) = n \times M = 0.00591 \times 569.6 = 3.37\,g$

c Use the amount of the limiting reagent, I_2, to determine the amount of P_4O_{10} formed. From the equation, 8 mol of I_2 react to form 3 mol of P_4O_{10}.

$\frac{n(P_4O_{10})}{n(I_2)} = \frac{3}{8}$

$n\left(P_4O_{10}\right) = \frac{3}{8} \times n\left(I_2\right) = \frac{3}{8} \times 0.0118 = 0.00443\,mol$

$\therefore\ m(P_4O_{10}) = n \times M = 0.00443 \times 283.88 = 1.26\,g$

d Total mass of products = 3.37 + 1.26 = 4.63 g. 4.63 g of products formed plus 2.78 g unreacted P_4O_6 = 7.40 g, which is consistent with the total mass of reactants used initially.

29 Student answers will vary. Possible answers include: a large molar mass, like that of sucrose, gives a smaller number of mol; there is no connection between the type of bonding and the number of mol in half a cup of a substance; low density and high molar mass substances have a low number of mol in half a cup of the substance.

Chapter 8 Concentration and molarity

8.1 Concentration of solutions

Worked example: Try yourself 8.1.1

CALCULATING CONCENTRATION IN GRAMS PER LITRE ($g\,L^{-1}$)

What is the concentration, in $g\,L^{-1}$, of a solution containing 5.00 g of glucose in 250 mL of solution?

Thinking	Working
Change the volume of solution so that it is expressed in litres.	$250\text{ mL} = \frac{250}{1000}$ $= 0.250\text{ L}$
Calculate the concentration (c) in $g\,L^{-1}$.	$c = \frac{\text{mass of solute (in g)}}{\text{volume of solution (in L)}}$ $= \frac{5.00}{0.250}$ $= 20.0\text{ g L}^{-1}$

Worked example: Try yourself 8.1.2

CALCULATING CONCENTRATION IN PARTS PER MILLION (PPM)

A sample of tap water was found to contain 0.0537 g of sodium chloride (NaCl) per 250.0 g of solution. Calculate the concentration of NaCl in parts per million (ppm).

Thinking	Working
Calculate the mass of solute in mg. Remember: mass (in mg) = mass (in g) × 1000	Mass of solute (NaCl) in mg $= 0.0537 \times 1000$ $= 53.7\text{ mg}$
Calculate the mass of solution in kg. Remember: mass (in kg) = $\frac{\text{mass (in g)}}{1000}$	Mass of solution in kg $= \frac{250.0}{1000}$ $= 0.2500\text{ kg}$
Calculate the concentration of the solution in $mg\,kg^{-1}$. This is the same as concentration in ppm.	Concentration of NaCl in ppm $= \frac{\text{mass of solute (in g)}}{\text{mass of solution (in kg)}}$ $= \frac{53.7}{0.2500}$ $= 215\text{ mg kg}^{-1}$ $= 215\text{ ppm}$

8.1 KEY QUESTIONS

1 D. Volume in litres = $\frac{60}{1000} = 0.060\text{ L}$. Concentration of sugar = $\frac{5.0}{0.060} = 83\text{ g L}^{-1}$

2 **a** $\frac{4.22 \times 100}{380} = 1.11\%(\text{w/v})$

b $\frac{4.22 \times 1000}{380} = 11.1\text{ g L}^{-1}$

3 **a** $\frac{244 \times 100}{750} = 32.5\%(\text{v/v})$

b $\frac{350 \times 4.9}{100} = 17\text{ mL}$

4 **a** $c(\text{CaCl}_2) = 5.0\,\text{ppm}$

b $c(\text{Pb(NO}_3)_2) = \frac{m(\text{Pb(NO}_3)_2)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(\text{Pb(NO}_3)_2) = 1.25\,\text{g} \times 1000 = 1250\,\text{mg}$

$m(\text{solution}) = \frac{2000}{1000} = 2.000\,\text{kg}$

$c(\text{Pb(NO}_3)_2) = \frac{1250}{2.000} = 625\,\text{ppm}$

c $c(\text{MgSO}_4) = \frac{m(\text{MgSO}_4)\text{ in mg}}{m(\text{solution})\text{ in kg}}$

$m(\text{MgSO}_4) = 4.0 \times 10^{-3}\,\text{g} \times 1000 = 4.0\,\text{mg}$

$m(\text{solution}) = \frac{150}{1000} = 0.150\,\text{kg}$

$c(\text{MgSO}_4) = \frac{4.0}{0.150} = 27\,\text{ppm}$

5 To find the %(w/v), the mass of each nutrient in grams is divided by the volume in millilitres and multiplied by 100.

Sugar = $\frac{35.0}{250} \times 100 = 14.0\%\text{(w/v)}$

Fat = $\frac{7.5}{250} \times 100 = 3.0\%\text{(w/v)}$

6 Parts per billion is the same as micrograms per kilogram.

Mass of ciguatoxin in µg = $15 \times 1000 = 15 \times 10^3\,\mu\text{g}$

The lethal concentration = $\frac{15 \times 10^3}{70} = 2.1 \times 10^2\,\text{ppb}$

7 $\frac{0.42}{175} \times 100 = 0.24\%\text{(w/w)}$

Therefore, the fluoride concentration is above the legal limit.

8.2 Molar concentration

Worked example: Try yourself 8.2.1

CALCULATING MOLAR CONCENTRATIONS (MOLARITY)

Calculate the molar concentration of a solution that contains 0.24 mol of glucose dissolved in 500 mL of solution.

Thinking	**Working**
Convert the given volume to litres.	$V(\text{glucose}) = \frac{500}{1000}$ $= 0.500\,\text{L}$
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(\text{glucose}) = \frac{n}{V}$ $= \frac{0.24}{0.500}$ $= 0.48\,\text{mol L}^{-1}$

Worked example: Try yourself 8.2.2

CALCULATING THE NUMBER OF MOLES OF SOLUTE IN A SOLUTION

Calculate the amount, in mol, of potassium permanganate ($KMnO_4$) in 100 mL of a 0.0250 mol L^{-1} solution of the compound.

Thinking	**Working**
Convert the given volume to litres.	$V(KMnO_4) = \frac{100}{1000}$ $= 0.100\,L$
Calculate the amount of compound, in mol, using the formula: $n = c \times V$	$n(KMnO_4) = c \times V$ $= 0.0250 \times 0.100$ $= 0.00250$ mol

Worked example: Try yourself 8.2.3

CALCULATING MOLARITY GIVEN THE MASS OF SOLUTE

Calculate the concentration, mol L^{-1}, of a solution that contains 4000 mg of ethanoic acid (CH_3COOH) dissolved in 100 mL of solution.

Thinking	**Working**
Convert the volume to litres.	$V(CH_3COOH) = \frac{100}{1000}$ $= 0.100\,L$
Convert the mass to grams.	$m(CH_3COOH) = \frac{4000}{1000}$ $= 4.000\,g$
Calculate the molar mass of the solute.	$M(CH_3COOH) = (12.01 \times 2) + (1.008 \times 4) + (16.00 \times 2)$ $= 60.05\,g\,mol^{-1}$
Calculate the number of moles of solute using the formula: $n = \frac{m}{M}$	$n(CH_3COOH) = \frac{m}{M}$ $= \frac{4.000}{60.05}$ $= 0.06661$ mol
Calculate the molar concentration using the formula: $c = \frac{n}{V}$	$c(CH_3COOH) = \frac{n}{V}$ $= \frac{0.06661}{0.100}$ $= 0.666\,mol\,L^{-1}$

8.2 KEY QUESTIONS

1 B. Convert the mass of H_2O_2 to moles:

$M(H_2O_2) = (1.008 \times 2) + (16.00 \times 2) = 34.02\,g\,mol^{-1}$

$n(H_2O_2) = \frac{m}{M} = \frac{5.00}{34.02} = 0.147$ mol

Volume $= \frac{120}{1000} = 0.120\,L$

Concentration $= \frac{0.147}{0.120} = 1.23\,mol\,L^{-1}$

2 **a** Volume = $\frac{25}{1000} = 0.025\,\text{L}$

Concentration = $\frac{2.0 \times 10^{-3}}{0.025} = 0.080\,\text{mol L}^{-1}$

b Concentration = $\frac{1.23}{4.1} = 0.30\,\text{mol L}^{-1}$

c Concentration = $\frac{1.8 \times 10^{3}}{9.3 \times 10^{3}} = 0.19\,\text{mol L}^{-1}$

3 **a** Concentration = $\frac{2.0}{8.0} = 0.25\,\text{mol L}^{-1}$

b Volume = $\frac{500}{1000} = 0.500\,\text{L}$, concentration = $\frac{0.25}{0.500} = 0.50\,\text{mol L}^{-1}$

c Volume = $\frac{200}{1000} = 0.200\,\text{L}$, concentration = $\frac{0.0876}{0.200} = 0.438\,\text{mol L}^{-1}$

4 Concentration = $\frac{0.98}{0.750} = 1.3\,\text{mol L}^{-1}$

5 **a** Amount = $c \times V = 0.22 \times 0.10 = 2.2 \times 10^{-2}\,\text{mol}$

b Volume = $\frac{10}{1000} = 0.010\,\text{L}$, amount = $c \times V = 0.64 \times 0.010 = 6.4 \times 10^{-3}\,\text{mol}$

c Volume = $\frac{15.6}{1000} = 0.0156\,\text{L}$, amount = $c \times V = 0.0150 \times 0.0156 = 2.34 \times 10^{-4}\,\text{mol}$

d Volume = $\frac{1.5 \times 10^{-1}}{1000} = 1.5 \times 10^{-4}\,\text{L}$, amount = $c \times V = 5.2 \times 1.5 \times 10^{-4} = 7.8 \times 10^{-4}\,\text{mol}$

6 Convert the mass of $AgNO_3$ to moles:

$M(AgNO_3) = 107.9 + 14.01 + (16.00 \times 3) = 169.9\,\text{g mol}^{-1}$

$n(AgNO_3) = \frac{m}{M} = \frac{5.09}{169.9} = 0.0300\,\text{mol}$

Volume = $\frac{250}{1000} = 0.250\,\text{L}$, concentration = $\frac{0.0300}{0.250} = 0.120\,\text{mol L}^{-1}$

7 Convert the mass of $CaCl_2$ to moles:

$M(CaCl_2) = 40.08 + (35.45 \times 2) = 110.98\,\text{g mol}^{-1}$

$n(CaCl_2) = \frac{m}{M} = \frac{1.223}{110.98} = 1.102 \times 10^{-2}\,\text{mol}$

Concentration = $\frac{1.102 \times 10^{-2}}{1.55} = 7.11 \times 10^{-3}\,\text{mol L}^{-1}$

8 $n(Na_2CO_3) = \frac{m}{M} = \frac{1.32}{105.99} = 0.0125\,\text{mol}$

$c = \frac{n}{V} = \frac{0.0125}{0.2500} = 0.0498\,\text{mol L}^{-1}$

8.3 Dilution

Worked example: Try yourself 8.3.1

CALCULATING THE CONCENTRATION OF A DILUTED SOLUTION

Calculate the concentration of the solution formed when 95.0 mL of water is added to 5.00 mL of 0.500 mol L^{-1} sulfuric acid (H_2SO_4).

Thinking	Working
Write down the value of c_1 and V_1. Note: c_1 and V_1 refer to the original solution, before water was added.	$c_1 = 0.500$ mol L^{-1} $V_1 = 5.00$ mL
Write down the value of V_2. Note: V_2 is the total volume of the original solution plus the added water.	$V_2 = 5.00 + 95.0$ $= 100.0$ mL
Transpose the equation $c_1V_1 = c_2V_2$ to calculate the concentration (c_2) of the new solution.	$c_1V_1 = c_2V_2$ $c_2 = \frac{c_1V_1}{V_2}$
Calculate the concentration of the diluted solution.	$c_2 = \frac{0.500 \times 5.00}{100}$ $= 0.0250$ mol L^{-1}

Worked example: Try yourself 8.3.2

CONVERTING CONCENTRATION UNITS

What is the concentration, in ppm, of a 0.0100 mol L^{-1} solution of NaOH? Remember that concentration in ppm is the same as mg L^{-1}.

Thinking	Working
Calculate the number of mol of solute in 1.00 L of the solution.	$n(\text{NaOH}) = c \times V$ $= 0.0100 \times 1.00$ $= 0.0100$ mol
Calculate the mass, in g, of solute in 1.00 L of the solution.	$M(\text{NaOH}) = 22.99 + 16.00 + 1.008$ $= 40.00$ g mol^{-1} $m(\text{NaOH}) = n \times M$ $= 0.0100 \times 40.00$ $= 0.400$ g
Calculate the mass, in mg, of solute in 1.00 L of the solution.	$m(\text{NaOH}) = 0.400 \times 1000$ $= 400$ mg
Express the concentration of the solute in ppm.	$c(\text{NaOH}) = 400$ ppm

8.3 KEY QUESTIONS

1 The volume of the solution is increased by a factor of 5, therefore the concentration will decrease by a factor of 5 to 0.40 mol L^{-1}.

Alternatively, $c_1 = \frac{c_2V_2}{V_1} = \frac{2.0 \times 50}{250} = 0.40$ mol L^{-1}

2 **a** $V_2 = 10.0 + 5.0 = 15.0$ mL, $c_2 = \frac{c_1V_1}{V_2} = \frac{1.2 \times 5.0}{15.0} = 0.40$ mol L^{-1}

b $V_2 = 1.0 + 3.0 = 4.0$ L, $c_2 = \frac{c_1V_1}{V_2} = \frac{0.1 \times 3.0}{4.0} = 0.075$ mol L^{-1}

c $V_2 = 5.0 + 95.0 = 100.0$ mL, $c_2 = \frac{c_1V_1}{V_2} = \frac{0.50 \times 5.0}{100.0} = 0.025$ mol L^{-1}

3 D. $V_2 = \frac{250}{1000} = 0.250$ L, $V_1 = \frac{c_1V_1}{V_2} = \frac{0.30 \times 0.250}{10} = 0.0075$ L $= 7.5 \times 10^{-3}$ L

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4 Mass of ammonia in 100 mL of original solution = 1.5 g

$M(NH_3) = 14.01 + (1.008 \times 3) = 17.03\,g\,mol^{-1}$

$n(NH_3) = \frac{m}{M} = \frac{1.5}{17.03} = 0.088\,mol$

$c_1 = \frac{0.088}{0.100} = 0.88\,mol\,L^{-1}$, $V_2 = 25 + 250 = 275\,mL$

$c_2 = \frac{c_1V_1}{V_2} = \frac{0.88 \times 25.0}{275} = 8.0 \times 10^{-2}\,mol\,L^{-1}$

5 In 100 mL of solution, there will be 17.0 g of NaOH.

$M(NaOH) = 22.99 + 1.008 + 16.00 = 40.00\,g\,mol^{-1}$

$n(NaOH) = \frac{m}{M} = \frac{17.0}{40.00} = 0.425\,mol$, $c(NaOH) = \frac{0.425}{0.100} = 4.25\,mol\,L^{-1}$

6 Calculate the volume of concentrated acid required using $c_1V_1 = c_2V_2$. $V_1 = 0.011\,L = 11\,mL$.

Add about a litre of water to a 2.0 L volumetric flask. Carefully add small amounts of concentrated sulfuric acid until you have added 11 mL. Make the volume up to the mark of 2.0 L.

8.4 Standard solutions

Worked example: Try yourself 8.4.1

CALCULATING THE CONCENTRATION OF A STANDARD SOLUTION PREPARED FROM A PRIMARY STANDARD

Calculate the concentration of a standard solution prepared from 117.0 g of sodium chloride (NaCl) dissolved in a 500.0 mL volumetric flask.

Thinking	Working
Use the chemical formula to determine the molar mass (*M*) of the compound.	$M(NaCl) = 22.99 + 35.45$ $= 58.44\,g\,mol^{-1}$
Use the mass (*m*) and molar mass (*M*) of the compound and the formula $n = \frac{m}{M}$ to determine the amount, in mol.	$n = \frac{117.0}{58.44}$ $= 2.002\,mol$
Use the amount, in mol, to determine the concentration of the solution using the formula $c = \frac{n}{V}$	$c = \frac{2.002}{0.5000}$ $= 4.004\,mol\,L^{-1}$

8.4 KEY QUESTIONS

1 C. Anhydrous Na_2CO_3 can be used as a primary standard. It is stable and has a known chemical formula.

2 Weigh the solid primary standard on an electronic balance.

Transfer the solid into the volumetric flask using a clean, dry funnel.

Rinse any remaining solid particles into the flask using deionised water.

Half fill the flask with deionised water and swirl vigorously to dissolve the solid.

Fill the flask with deionised water to just below the calibration mark.

Add deionised water drop by drop up to the calibration line on the flask until the bottom of the meniscus touches the line.

Stopper and shake the solution to ensure an even concentration throughout.

3 $n = \frac{m}{M} = \frac{2.042}{204.22} = 0.009\,999\,mol$

$c = \frac{n}{V} = \frac{0.009999}{0.05000} = 0.2000\,mol\,L^{-1}$

4 $c = \frac{n}{V} = 0.500 = \frac{n}{0.250}$

$n = 0.500 \times 0.2500 = 0.125\,mol$

$m = n \times M = 0.125 \times 105.99 = 13.2\,g$

5 $n(\text{hydrated oxalic acid}) = c \times V = 0.200 \times 0.500 = 0.100\,mol$

$m = n \times M = 0.100 \times 126.08 = 12.6\,g$

CHAPTER 8 REVIEW

1 B. Since the ethanol is a liquid and the solution is a liquid, %(v/v) is the unit to use.

2 **a** Concentration of lead (ppm) = $\frac{m(\text{Pb}) \text{ in mg}}{m(\text{solution}) \text{ in kg}} = \frac{12}{6.0} = 2.0$ ppm

b Percentage by mass = $\frac{m(\text{Pb}) \text{ in mg}}{m(\text{solution}) \text{ in mg}} \times 100$, mass of solution = 6.0 kg × $(1.0 \times 10^6) = 6.0 \times 10^6$ mg

%(w/w) = $\frac{12}{6.0 \times 10^6} \times 100 = 2.0 \times 10^{-4}$%(w/w)

3 %(w/w) = $\frac{15}{300} \times 100 = 5.0$%(w/w)

4 Concentration of active ingredient (mg L^{-1}) = $\frac{\text{mass (in mg)}}{\text{volume of solution (in L)}}$

Volume of solution = $\frac{5.0}{1000} = 0.0050$ L, concentration = $\frac{2.00}{0.0050} = 4.0 \times 10^2$ mg L^{-1}

Concentration of active ingredient (%(w/v)) = $\frac{\text{mass (in g)}}{\text{volume (in mL)}} \times 100$

Mass of active ingredient = $\frac{2.00}{1000} = 2.00 \times 10^{-3}$ g

Concentration %(w/v) = $\frac{2.00 \times 10^{-3}}{5.0} \times 100 = 4.0 \times 10^{-2}$%(w/v)

5 **a** Mass = $\frac{1.50 \times 1300}{100} = 19.5$ g

b Mass water = 1300 − 19.5 = 1281 g

6 **a** Solution contains 1.2 mg per 15 mL.

In 1 L, it will contain $\frac{1.2 \times 1000}{15} = 80$ mg L^{-1}

b %(w/v) = $\frac{1.2}{15} \times 100 = 8.0$%(w/v)

7 $c = \frac{n}{V}$, $V = 0.160$ L, $c(\text{NaBr}) = \frac{0.380}{0.160} = 2.38$ mol L^{-1}

8 $c = \frac{n}{V}$, $M(\text{NaOH}) = 22.99 + 16.00 + 1.008 = 40.00$ g mol^{-1},

$n(\text{NaOH}) = \frac{m}{M} = \frac{30}{40.00} = 0.75$ mol,

$c(\text{NaOH}) = \frac{n}{V} = \frac{0.75}{2.0} = 0.38$ mol L^{-1}

9 $c = \frac{n}{V}$, in 100 mL of solution there will be 4.26 g of solute,

$M(\text{Na}_2\text{SO}_4) = (22.99 \times 2) + 32.07 + (16.00 \times 4) = 142.05$ g mol^{-1}

$n(\text{Na}_2\text{SO}_4) = \frac{m}{M} = \frac{4.26}{142.05} = 0.0300$ mol

$c(\text{Na}_2\text{SO}_4) = \frac{n}{V} = \frac{0.0300}{0.100} = 0.300$ mol L^{-1}

10 **a** Volume = $\frac{12}{1000} = 0.012$ L, $n = cV = 0.22 \times 0.012 = 2.6 \times 10^{-3}$ mol

b Volume = $\frac{150}{1000} = 0.150$ L, $n = cV = 0.0250 \times 0.150 = 3.75 \times 10^{-3}$ mol

c $n = cV = 3.15 \times 10^{-3} \times 7.2 = 2.3 \times 10^{-2}$ mol

11 **a** Volume = $\frac{100}{1000}$ = 0.100 L

$n(NH_3) = cV = 1.2 \times 0.100 = 0.120$ mol

$M(NH_3) = 14.01 + (1.008 \times 3) = 17.03\ g\ mol^{-1}$

$m(NH_3) = n \times M$

$= 0.120 \times 17.03 = 2.04$ g

b Volume = $\frac{20}{1000}$ = 0.020 L

$n(AgNO_3) = cV = 0.50 \times 0.020 = 0.010$ mol

$M(AgNO_3) = 107.9 + 14.01 + (16.00 \times 3) = 169.9\ g\ mol^{-1}$

$m(AgNO_3) = n \times M = 0.010 \times 169.9 = 1.7$ g

12 **a** $n(HCl) = c \times V = 1.00 \times 2.50 = 2.50$ mol

$m(HCl) = n \times M = 2.50 \times 36.46 = 91.2$ g

b Concentration %(w/v) = $\frac{91.2}{2500} \times 100 = 3.65$%(w/v)

13 D. The addition of water will not change the amount of solute present.

14 $c_2 = \frac{c_1V_1}{V_2} = \frac{5.00 \times 50.0}{250} = 1.00\ mol\ L^{-1}$

15 $V_2 = \frac{c_1V_1}{c_2} = \frac{4.0 \times 25}{1.6} = 62.5$ mL, so volume to be added = 62.5 – 25 = 37.5 = 38 mL

16 $n(CaCl_2)$ in the original solution = $c \times V$, volume = $\frac{1.50}{1000}$ = 0.00150 L

$n(CaCl_2) = 0.0500 \times 0.00150 = 7.50 \times 10^{-5}$ mol of $CaCl_2$, and therefore $n(Cl^-) = 2 \times n(CaCl_2) = 2 \times 7.50 \times 10^{-5} = 1.50 \times 10^{-4}$ mol of Cl^- ions

$M(Cl^-) = 35.45\ g\ mol^{-1}$

$m(Cl^-) = n \times M = 1.50 \times 10^{-4} \times 35.45 = 5.32 \times 10^{-3}$ g

Concentration of final solution (ppm) = $\frac{m(Cl^-)\text{ in mg}}{V(Cl^-)\text{ in L}}$

$m(Cl^-) = 5.32 \times 10^{-3}\ g \times 1000 = 5.32$ mg

$c(Cl^-) = \frac{5.32}{10.0} = 0.532$ ppm

17 Calculate the volume of diluted solution that can be prepared from the concentrated solution.

$V_2 = \frac{c_1V_1}{c_2} = \frac{1 \times 18}{1} = 18$ L

The number of 2 L bottles that can be filled will be nine.

18 C. The colour of a compound is not usually considered when selecting a primary standard. Primary standards must be soluble because they react with the substance being analysed in solution. A high level of purity minimises errors in calculating the results of the analyses. A high molar mass minimises the percentage error in calculations.

19 $n = \frac{m}{M} = \frac{25.21}{126.08} = 0.2000$ mol

$c = \frac{n}{V} = \frac{0.2000}{0.2500} = 0.7998\ mol\ L^{-1}$

20 $c = \frac{n}{V}$, $0.400 = \frac{n}{0.500}$, $n = 0.400 \times 0.500 = 0.200$ mol

$n = \frac{m}{M}$, $m = n \times M$, $m = 0.200 \times 105.99 = 21.2$ g

21 A. Each answer needs to be tested to find the mass of H_2SO_4. In A, the $n(H_2SO_4)$ is 0.5 × 2 = 1 mol. This amount is greater than the other alternatives.

22 **a** The chemical formula of hydrated sodium carbonate ($Na_2CO_3 \cdot 10H_2O$) changes over time. It can lose water, changing the ratio of sodium carbonate to water.

b Sodium hydroxide absorbs moisture from the air, and it also reacts readily with carbon dioxide in the air. This means the sodium hydroxide is relatively unstable, and therefore is not a completely pure substance.

c Sodium carbonate is heated to ensure it is dry so that the mass of solid is accurate.

d $n(Na_2CO_3) = c \times V = 0.200 \times 0.250 = 0.0500$ mol

$m = n \times M = 0.0500 \times 105.99 = 5.30$ g

23 **a** Product A. mass of iodine = 7.0 g

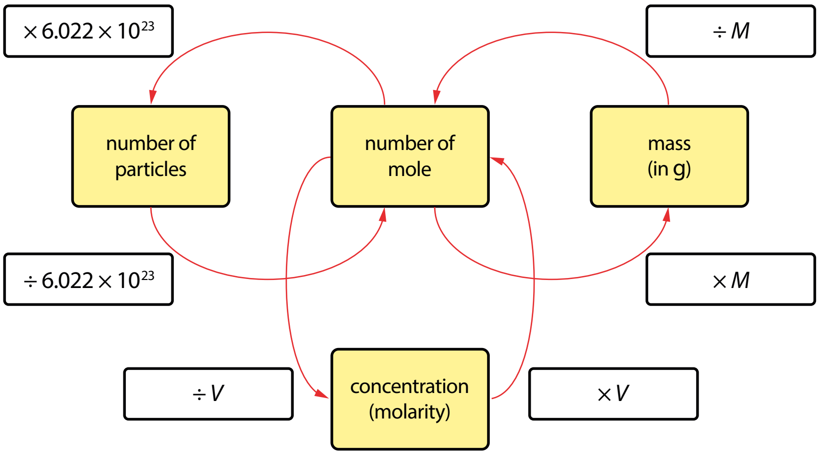
Product B. mass of iodine = 20 g

Product B supplies the greater mass of $\frac{500 \times 4.0}{100}$ iodine.

b Product A: $\frac{7.0}{4.10} = 1.7$ g per \$1

Product B: $\frac{20}{12.80} = 1.6$ g per \$1   Product A is better value.

24



25 The mass of solute in 100 mL of solution is 20 g.

$c = \frac{n}{V}$, $n(CuSO_4{\cdot}5H_2O) = \frac{m(CuSO_4\cdot5H_2O)}{M(CuSO_4\cdot5H_2O)}$

$M(CuSO_4{\cdot}5H_2O) = 63.55 + 32.07 + (16.00 \times 4) + 5 \times ((1.008 \times 2) + 16.00) = 249.70\ \text{g mol}^{-1}$

$n(CuSO_4{\cdot}5H_2O) = \frac{20}{249.70} = 0.080\ \text{mol}$

$c(CuSO_4{\cdot}5H_2O) = \frac{0.080}{0.100} = 0.80\ \text{mol L}^{-1}$

26 **a** $n(NaOH) = c \times V = 1.65 \times 25.0 = 41.3\ \text{mol}$

$m = n \times M = 41.3 \times 40.00 = 1650\ \text{g} = 1.65\ \text{kg}$

b Total mass of NaOH = mass in one container × number of containers = 1.65 × 180 = 297 kg

c Mass of NaOH in 25.0 L is 1650 g. In g L^{-1} = $\frac{1650}{25.0} = 66.0\ \text{g L}^{-1}$

d $c_2 = \frac{c_1V_1}{V_2} = \frac{1.65 \times 8.0}{25.0} = 0.53\ \text{mol L}^{-1}$

27 **a** The gas evolved is carbon dioxide, formed from the reaction of the carbonate ions.

b acetic acid + sodium hydrogen carbonate → sodium acetate + carbon dioxide + water

c Baking soda will release carbon dioxide gas during cooking. The gas is trapped in the cake, making the cake rise and the texture light.

Chapter 9 Gas laws

9.1 Introducing gases

Worked example: Try yourself 9.1.1

CONVERTING VOLUME UNITS

A gas has a volume of 700 mL.

a What is its volume in cubic centimetres (cm^3)?

b What is its volume in litres (L)?

c What is its volume in cubic metres (m^3)?

Thinking	Working
Recall the conversion factors for each of the units of volume. Apply the correct conversion to each situation.	**a** The units of mL and cm^3 are equivalent. $1\ mL = 1\ cm^3$ $700\ mL = 700\ cm^3$ **b** $1000\ mL = 1\ L$ Divide volume in mL by 1000 to convert to L. $700\ mL = \frac{700}{1000}$ $= 0.700\ L$ **c** $1 \times 10^6\ mL = 1\ m^3$ Divide volume in mL by 1×10^6 to convert to m^3. $700\ mL = \frac{700}{1 \times 10^6}$ $= 7.00 \times 10^{-4}\ m^3$

Worked example: Try yourself 9.1.2

CONVERTING PRESSURE UNITS

Cyclone Debbie was one of the biggest cyclones in Australian history.

a The atmospheric pressure in the eye of Cyclone Debbie was measured as 0.959 bar. What was the pressure in kilopascals (kPa)?

b The atmospheric pressure in the eye of Cyclone Debbie was 720 mmHg. What was the pressure in atmospheres (atm)?

c The atmospheric pressure in the eye of Cyclone Debbie was 0.947 atm. What was the pressure in kilopascals (kPa)?

d The atmospheric pressure in the eye of Cyclone Debbie was 720 mmHg. What was the pressure in bars?

Thinking	Working
a To convert bars to kilopascals, use the conversion relationship: 1 bar = 100 kPa To change bar to kPa, multiply the value by 100.	$0.959\ bar = 0.959 \times 100$ $= 95.9\ kPa$
b To convert millimetres of mercury to atmospheres, use the relationship: 1 atm = 760 mmHg To change mmHg to atm, divide the value by 760.	$720\ mmHg = \frac{720}{760}$ $= 0.947\ atm$

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c To convert atmospheres to kilopascals, use the conversion relationship: 1 atm = 101.3 kPa To change atm to kPa, multiply the value by 101.3.	0.947 atm = 0.947 × 101.3 = 95.9 kPa
d This can be done in two steps. First, convert millimetres of mercury to atmospheres. Use the conversion relationship: 760 mmHg = 1 atm To change mmHg to atm, divide the value by 760. Keep the answer in your calculator and proceed to the next step.	720 mmHg = $\frac{720}{760}$ atm
Convert atmospheres to bars. Use the conversion relationship: 1 atm = 1.013 bar To change atm to bar, multiply the quotient from the previous step by 1.013.	720 mmHg = $\frac{720}{760}$ × 1.013 = 0.959 bar

9.1 KEY QUESTIONS

1 **a** Molecules of gases are in constant, rapid random motion, and the forces between the molecules are negligible. They continue to move outwards until stopped by the walls of the container, filling all the space available.

b Most of the volume occupied by a gas is space, so compression can be achieved by reducing the space between the particles.

c The molecules in a gas are spread much further apart than those of a liquid. A given mass of gas would occupy a much greater volume than the same mass in the liquid phase. Therefore, the density of the gas is less.

d Gases mix easily together because of the large amount of space between the molecules.

e The pressure exerted by a gas depends on the number of collisions of the gas particles with the walls of the container. The pressure is independent of the type of gas involved. The total pressure exerted by a mixture of gases will depend on the total number of collisions each gas has with the container walls.

2 **a** Tyres have a recommended maximum pressure for giving a comfortable ride as well as good traction on the road. If the pressure in a tyre is too high, the gas inside cannot be compressed as easily and passengers will be more aware of bumps on the road.

b During a long journey on a hot day, the air in a tyre warms up. This means the air molecules have increased kinetic energy, and collisions with the walls of the tyres will increase in frequency, exerting more force. This will cause the pressure to increase.

c Particles from cooking food escape the pot and move randomly through the house. If the food has an odour, and if there are enough particles in the air, you will detect the odour as you enter the house.

d As air is pumped into a balloon, air molecules collide with the rubber of the balloon, forcing it to expand. If too much air is pumped in, the balloon reaches a stage where it cannot stretch any further. If the number of collisions with molecules per given surface area is increased still further, the rubber will break.

3 **a** As temperature increases, the average kinetic energy of the gas molecules in the can will increase. This will lead to an increase in the frequency and force of collisions of gas molecules with the inside wall of the aerosol can. This will cause an increase in pressure.

b As the syringe is compressed, the inside surface area of the syringe will decrease. The number of collisions of molecules per unit area per second with the inside walls of the syringe will increase. This will cause a pressure increase.

4 These conversions should be applied as needed: 1.00 atm = 760 mmHg = 101.3 kPa = 1.013 × 10^5 Pa = 1.013 bar

a 140 kPa = 140 × 1000 = 1.40 × 10^5 Pa

b 92 000 Pa = $\frac{92\,000}{1000}$ = 92 kPa

c 1.00 atm = 760 mmHg

So 4.24 atm = 760 × 4.24 = 3.22 × 10^3 mmHg

1.00 atm = 1.013 × 10^5 Pa

So 4.24 atm = 1.013 × 10^5 × 4.24 = 4.30 × 10^5 Pa

d $101.3\ \text{kPa} = 760\ \text{mmHg}$

So $120\ \text{kPa} = \frac{760}{101.3} \times 120 = 900\ \text{mmHg}$

$101.3\ \text{kPa} = 1.00\ \text{atm}$

So $120\ \text{kPa} = \frac{1.00}{101.3} \times 120 = 1.18\ \text{atm}$

$100\ \text{kPa} = 1\ \text{bar}$

So $120\ \text{kPa} = \frac{1}{100} \times 120 = 1.20\ \text{bar}$

e $760\ \text{mmHg} = 1.00\ \text{atm}$

So $1400\ \text{mmHg} = 1.00 \times \frac{1400}{760} = 1.84\ \text{atm}$

$760\ \text{mmHg} = 1.013 \times 10^5\ \text{Pa}$

So $1400\ \text{mmHg} = 1.013 \times 10^5 \times \frac{1400}{760} = 1.87 \times 10^5\ \text{Pa}$

$760\ \text{mmHg} = 1.013\ \text{bar}$

So $1400\ \text{mmHg} = 1.013 \times \frac{1400}{760} = 1.87\ \text{bar}$

f $1.013 \times 10^5\ \text{Pa} = 1.00\ \text{atm}$

So $80\,000\ \text{Pa} = 1.00 \times \frac{80\,000}{1.013 \times 10^5} = 0.790\ \text{atm}$

$1.013 \times 10^5\ \text{Pa} = 760\ \text{mmHg}$

So $80\,000\ \text{Pa} = 760 \times \frac{80\,000}{1.013 \times 10^5} = 600\ \text{mmHg}$

$1.013 \times 10^5\ \text{Pa} = 1.013\ \text{bar}$

So $80\,000\ \text{Pa} = 1.013 \times \frac{80\,000}{1.013 \times 10^5} = 0.800\ \text{bar}$

5 These conversions should be applied as needed: $1\ \text{mL} = 1\ \text{cm}^3$, $1\ \text{L} = 1000\ \text{mL} = 1000\ \text{cm}^3$, $1\ \text{m}^3 = 1 \times 10^6\ \text{cm}^3 = 1 \times 10^6\ \text{mL} = 1000\ \text{L}$

a $2\ \text{L} = 2 \times 1000 = 2 \times 10^3\ \text{mL}$

b $4.5\ \text{L} = \frac{4.5}{1000} = 4.5 \times 10^{-3}\ \text{m}^3$

c $2250\ \text{mL} = \frac{2250}{1000} = 2.250\ \text{L}$

d $120\ \text{mL} = \frac{120}{1000} = 0.120\ \text{L}$

e $5.6\ \text{mL} = \frac{5.6}{1000} = 5.6 \times 10^{-3}\ \text{L}$

f $3.7\ \text{m}^3 = 3.7 \times 1000 = 3.7 \times 10^3\ \text{L}$

g $285\ \text{mL} = 285\ \text{cm}^3 = \frac{285}{1 \times 10^6} = 2.85 \times 10^{-4}\ \text{m}^3$

h
$$\begin{aligned} 4.70 \times 10^{-3}\ \text{m}^3 &= 4.70 \times 10^{-3} \times 1000 \\ &= 4.70\ \text{L} \\ &= 4.70 \times 1000 \\ &= 4.70 \times 10^3\ \text{cm}^3 \end{aligned}$$

6 **a**
$$\begin{aligned} P_{\text{tot}} &= P_1 + P_2 \\ 100 &= 2.5 + P_{\text{dry air}} \\ P_{\text{dry air}} &= 100 - 2.5 \\ &= 97.5\ \text{kPa} \end{aligned}$$

b $\frac{P_{\text{gas}}}{P_{\text{tot}}} \times 100 = \frac{160}{760} \times 100$

$= 21.1\%$

9.2 The gas laws

Worked example: Try yourself 9.2.1

USING BOYLE'S LAW TO PREDICT CHANGES IN PRESSURE AND VOLUME

The swim bladder of a fish has a volume of 30 mL near the surface of the water, where the pressure acting on it is 1.00 bar. Assuming no changes to the gas composition of the bladder, and no change in the temperature of the water, what will be the volume of the bladder at a depth of 10 m, where the pressure on the swim bladder is 2.00 bar?

Thinking	**Working**
State Boyle's law.	$PV = k$
Determine the relationship between the two states.	As the amount of gas has not changed and the temperature has been held constant, the relationship between the two states is: $P_1V_1 = k = P_2V_2$
State the known and unknown variables.	P_1 = 1.00 bar V_1 = 30 mL P_2 = 2.00 bar V_2 = ?
Rearrange the equation to make V_2 the subject.	$P_1V_1 = P_2V_2$ $V_2 = \frac{P_1V_1}{P_2}$
Substitute in the known values and solve.	$V_2 = \frac{P_1V_1}{P_2}$ $V_2 = \frac{1.00 \times 30}{2.00}$ = 15 mL
State the new volume.	The new volume is 15 mL.

Worked example: Try yourself 9.2.2

CONVERTING TEMPERATURES FROM CELSIUS TO KELVIN

What is 100°C on the Kelvin temperature scale?

Thinking	**Working**
T (in K) = T (in °C) + 273	T (in K) = T (in °C) + 273 = 100 + 273 = 373 K

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Worked example: Try yourself 9.2.3

USING CHARLES' LAW TO PREDICT CHANGES IN TEMPERATURE AND VOLUME

A hot air balloon at 273 K has a volume of 355 L. The temperature is increased to 395 K. If the pressure remains constant, what will be the new volume of the balloon?

Thinking	**Working**
State Charles' law.	$V = kT$
Determine the relationship between the two states.	As the amount of gas has not changed and the pressure between the two states has been held constant, the relationship between the two states is: $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
State the known and unknown variables.	$V_1 = 355\text{ L}$ $T_1 = 273\text{ K}$ $V_2 = ?$ $T_2 = 395\text{ K}$
Rearrange the equation to make V_2 the subject.	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $V_2 = \frac{V_1 T_2}{T_1}$
Substitute in the known values and solve.	$V_2 = \frac{V_1 T_2}{T_1}$ $= \frac{355 \times 395}{273}$ $= 514\text{ L}$
State the new volume.	The new volume is 514 L.

Worked example: Try yourself 9.2.4

USING GAY-LUSSAC'S LAW TO PREDICT CHANGES IN TEMPERATURE AND PRESSURE

A gas in a sealed container at 25°C is heated in an oven. If the initial gas pressure was 1.0 bar and the final gas pressure is 1.6 bar, calculate the new temperature of the gas.

Thinking	**Working**
State Gay-Lussac's law.	$\frac{P}{T} = k$
Determine the relationship between the two states.	As the amount of gas has not changed and the volume between the two states has been held constant, the relationship between the two states is: $\frac{P_1}{T_1} = k = \frac{P_2}{T_2}$
Convert temperature from degrees Celsius to kelvin.	$T_1 = 273 + 25$ $= 298\text{ K}$
State the known and unknown variables.	$P_1 = 1.0\text{ bar}$ $T_1 = 298\text{ K}$ $P_2 = 1.6\text{ bar}$ $T_2 = ?$
Rearrange the equation to make T_2 the subject.	$T_2 = \frac{P_2 T_1}{P_1}$

Substitute in the known values.	$T_2 = \frac{P_2T_1}{P_1}$ $= \frac{1.6 \times 298}{1.00}$ $= 477\,\text{K}$
Convert temperature from kelvin to degrees Celsius.	$T_2 = 477 - 273$ $= 204°\text{C}$
State the new temperature.	The final temperature is 204°C.

Worked example: Try yourself 9.2.5

USING THE COMBINED GAS LAW

A student measured the pressure and volume of carbon dioxide at 25°C to be 96.0 kPa and 500 mL. The gas was then heated to 450°C. Calculate the new pressure if the volume increased to 338 mL.

Thinking	**Working**
Determine the relationship between the two states.	As the amount of gas has not changed, the relationship between the two states is: $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Convert temperature from degrees Celsius to kelvin.	$T_1 = 273 + 25 = 298\,\text{K}$ $T_2 = 273 + 450 = 723\,\text{K}$
State the known and unknown variables.	$P_1 = 96.0\,\text{kPa}$ $V_1 = 500\,\text{mL}$ $T_1 = 298\,\text{K}$ $P_2 = ?$ $V_2 = 338\,\text{mL}$ $T_2 = 723\,\text{K}$
Rearrange the equation to make P_2 the subject.	$P_2 = \frac{P_1V_1T_2}{T_1V_2}$
Substitute in the known values.	$P_2 = \frac{96.0 \times 500 \times 723}{298 \times 338}$ $= 345\,\text{kPa}$
State the new gas pressure.	The pressure of the gas at 450°C with a volume of 338 mL is 345 kPa.

9.2 KEY QUESTIONS

1 To convert from degrees Celsius to kelvin add 273.

- **a** 273 + 0 = 273 K
- **b** 273 + 25 = 298 K
- **c** 273 + 100 = 373 K
- **d** 273 + 175 = 448 K
- **e** 273 + (–145) = 128 K

2

	Initial condition		**Final condition**	
	Pressure	**Volume**	**Pressure**	**Volume**
a	1 atm	10 L	5 atm	$\frac{1\times10}{5}=2$ L
b	160 mmHg	235 mL	760 mmHg	$\frac{160\times235}{760}=49.5$ mL
c	101.3 kPa	1.6 L	693 kPa	$\frac{101.3\times1.6}{693}=0.23$ mL
d	1.00 bar	20 L	3.00 bar	$\frac{1.00\times20}{3.00}=6.7$ L
e	1.0 atm	100 L	200 atm	$\frac{1.0\times100}{200}=0.50$ L

3

	Initial condition		**Final condition**	
	Volume	**Temperature**	**Volume**	**Temperature**
a	100 mL	300 K	$\frac{100\times900}{300}=300$ mL	900 K
b	500 mL	500°C	$\frac{500\times(273+250)}{(273+500)}=338$ mL	250°C
c	10.0 L	25°C	$\frac{10.0\times(273+300)}{(273+25)}=19.2$ L	300°C
d	10.0 L	300 K	25.0 L	$\frac{25.0\times300}{10.0}=750$ K
e	3.2 m^3	27°C	6.9 m^3	$\frac{6.9\times(273+27)}{3.2}=647$ K $T_2 = 647 - 273 = 374$°C

4

	Initial condition		**Final condition**	
	Pressure	**Temperature**	**Pressure**	**Temperature**
a	5.0 atm	300 K	$P_2 = \frac{5.0 \times 900}{300} = 15$ atm	900 K
b	760 mmHg	500°C	Convert temperatures: $T_1 = 773$ K; $T_2 = 523$ K $P_2 = \frac{760 \times 523}{773} = 514$ mmHg	250°C
c	693 kPa	25°C	1582 kPa	Convert temperature: $T_1 = 298$ K $T_2 = \frac{1582 \times 298}{693}$ $= 680$ K $= 407$°C
d	300 bar	300 K	500 bar	$T_2 = \frac{500 \times 300}{300}$ $= 500$ K
e	200 atm	300 K	$P_2 = \frac{200 \times 650}{300}$ $= 433$ atm	650 K

5

	Initial condition			**Final condition**		
	Pressure	**Volume**	**Temperature**	**Pressure**	**Volume**	**Temperature**
a	101.3 kPa	10.0 L	25°C	53.2 kPa	Convert temperatures: $T_1 = 298$ K; $T_2 = 217$ K $V_2 = \frac{101.3 \times 10.0 \times 217}{298 \times 53.2} = 13.9$ L	–56°C
b	1.00 atm	3.2 L	3000 K	$P_2 = \frac{1.00 \times 3.2 \times 273}{3000 \times 1.0}$ $= 0.29$ atm	1.0 L	273 K
c	5.00 bar	116 mL	–25°C	Convert temperatures: $T_1 = 248$ K; $T_2 = 348$ K $P_2 = \frac{5.00 \times 116 \times 348}{248 \times 100}$ $= 8.14$ bar	100 mL	75°C
d	10 atm	375 mL	298 K	1 atm	7500 mL	$T_2 = \frac{1 \times 7500 \times 298}{10 \times 375} = 596$ K
e	760 mmHg	10.0 L	298 K	480 mmHg	$V_2 = \frac{760 \times 10.0 \times 273}{298 \times 480}$ $= 14.5$L	273 K

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6 $T_1 = 25 + 273 = 298\,\text{K}$

$T_2 = -50 + 273 = 223\,\text{K}$

$P_1 = 1\,\text{atm}$

$P_2 = ?$

$V_1 = 5.0\,\text{L}$

$V_2 = 10.0\,\text{L}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

$$P_2 = \frac{1 \times 5.0 \times (273 - 50)}{(273 + 25) \times 10.0}$$

$$= 0.37\,\text{atm}$$

The pressure on the balloon at the higher altitude is 0.37 atm.

7 $P_1 = 30\,\text{kPa}$

$P_2 = ?$

$T_1 = 15 + 273 = 283\,\text{K}$

$T_2 = 47 + 273 = 320\,\text{K}$

$V = \text{constant}$

$$P_2 = \frac{P_1T_2}{T_1}$$

$$P_2 = \frac{30 \times 320}{283}$$

$$= 34\,\text{kPa}$$

The new air pressure in the tyre is 34 kPa.

8 **a** $V_1 = 10\,000\,\text{m}^3$

$V_2 = ?$

$T_1 = 20 + 273 = 298\,\text{K}$

$T_2 = 5 + 273 = 278\,\text{K}$

$P_1 = 1.00\,\text{bar}$

P_2 is calculated from the height. A 1500 m altitude represents a drop in pressure of:

$$\frac{1500}{1000} \times 0.10 = 0.15\,\text{bar}$$

$$\therefore\ P_2 = 1.00 - 0.15 = 0.85\,\text{bar}$$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$= \frac{1.0 \times 10\,000 \times 278}{293 \times 0.85}$$

$$= 11\,162\,\text{m}^3$$

$$= 1.1 \times 10^4\,\text{m}^3$$

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b **i** $T_1 = 20 + 273 = 293\,\text{K}$

$T_2 = 0 + 273 = 273\,\text{K}$

$V_1 = 10\,000\,\text{m}^3$

$V_2 = 10\,500\,\text{m}^3$

$P_1 = 1.00\,\text{bar}$

$P_2 = ?$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$= \frac{1.00 \times 10\,000 \times 273}{293 \times 10\,500}$$

$$= 0.887\,\text{bar}$$

ii The pressure drops at a rate of 0.10 bar per 1000 m rise in altitude.

Drop in pressure = 1.00 – 0.887 = 0.11 bar.

Therefore, the height of the airship = $\frac{1000}{0.10} \times 0.11 = 1100\,\text{m}$.

9.3 The ideal gas law

Worked example: Try yourself 9.3.1

CALCULATING THE VOLUME OF A GAS FROM ITS AMOUNT (IN MOL)

Calculate the volume, in L, occupied by 3.5 mol of oxygen gas at SLC. Assume that oxygen behaves as an ideal gas.

Thinking	Working
Rearrange $n = \frac{V}{V_m}$ to make volume the subject.	$n = \frac{V}{V_m}$ $V = n \times V_m$
Substitute in the known values, including $V_m = 24.79\,\text{L mol}^{-1}$ (at SLC), and solve.	$V = n \times V_m$ $= 3.5 \times 24.79$ $= 86.765\,\text{L}$
Consider the units and significant figures.	$V = 87\,\text{L}$

Worked example: Try yourself 9.3.2

USING THE IDEAL GAS LAW TO CALCULATE THE VOLUME OF A GAS

Calculate the volume, in L, occupied by 13.0 mol of carbon dioxide gas (CO_2) if the pressure is 250 kPa at 75.0°C.

Thinking	Working
Convert units, if necessary. Pressure is in kPa and temperature in K.	$P = 250\,\text{kPa}$ (no conversion required) $T = 75.0 + 273$ $= 348\,\text{K}$
Rearrange the ideal gas law so that volume, V, is the subject.	$PV = nRT$ $V = \frac{nRT}{P}$
Substitute in values for pressure, amount, temperature and the gas constant, R, then solve for V.	$V = \frac{13.0 \times 8.314 \times 348}{250}$ $= 150\,\text{L}$

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9.3 KEY QUESTIONS

1 Remember that, under standard laboratory conditions (SLC), 1 mol of any gas has a volume of 24.79 L.

Use the formula $n = \frac{V}{V_m}$, where n is the amount, in mol, V is the volume, in L, and V_m is the molar volume in, L mol^{-1}.

To calculate V, the formula is rearranged to $V = n \times V_m$.

a $V(Cl_2) = 24.79 \times 1.4 = 34.706\,L = 35\,L$

b $V(H_2) = 24.79 \times 1.0 \times 10^{-3}\,L = 0.024\,79\,L = 25\,mL$

c Since $n = \frac{m}{M}$ and the molar mass of nitrogen is 28.02 g mol^{-1}:

$n = \frac{m}{M} = \frac{1.4}{28.02} = 0.050$ mol

$V(N_2) = n \times V_m = 0.050 \times 24.79 = 1.239\,L = 1.2\,L$

2 Remember, that under standard laboratory conditions (SLC), 1 mol of any gas has a volume of 24.79 L.

These questions involve three steps:

1 Calculate the amount, in mol, of gas at SLC.

2 Find the molar mass (M) of the gas.

3 Use $n = \frac{m}{M}$ to find the mass by reorganising as $m = n \times M$.

a $n(Ne) = \frac{2.80}{24.79} = 0.1129\,mol$

$m(Ne) = 0.1129 \times 20.18 = 2.28\,g$

b $n(O_2) = \frac{50.0}{24.79} = 2.017\,mol$

$m(O_2) = 2.017 \times 32.00 = 64.5\,g$

c $V = 140\,mL = 0.140\,L$

$n(CO_2) = \frac{0.140}{24.79} = 0.005\,647\,mol$

$m(CO_2) = 0.005\,647 \times 44.01 = 0.249\,g$

3 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L.
$R = 8.314\,J\,mol^{-1}\,K^{-1}$.

$P = \frac{nRT}{V} = \frac{0.25 \times 8.314 \times (5+273)}{5.0}$

$= 1.2 \times 10^3\,kPa$

4 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L.
$R = 8.314\,J\,mol^{-1}\,K^{-1}$.

a $V = \frac{nRT}{P} = \frac{0.20 \times 8.314 \times (40+273)}{115} = 4.5\,L$

b $V = \frac{nRT}{P} = \frac{12.5 \times 8.314 \times (150+273)}{5.00 \times 101.3} = 86.8\,L$

c Since $n = \frac{m}{M}$ and the molar mass of hydrogen sulfide is 34.09 g mol^{-1}:

$V = \frac{nRT}{P} = \frac{8.50 \times 8.314 \times (27+273)}{34.09 \times 100} = 6.22\,L$

5 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L.
$R = 8.314\,J\,mol^{-1}\,K^{-1}$.

The second step involves use of $n = \frac{m}{M}$.

The molar mass of helium is 4.003 g mol^{-1}.

$n(He) = \frac{PV}{RT} = \frac{95.000 \times 1\,00}{8.314 \times (273+0)}$

$= 4.19\,mol$

$m(He) = 4.19 \times 4.003$

$= 16.8\,g$

6 The ideal gas is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L.
$R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$

Since $n = \frac{m}{M}$ and the molar mass of nitrogen is 28.02 g mol⁻¹:

$$T = \frac{PV}{nR} = \frac{102 \times 10.0 \times 28.02}{11.3 \times 8.314}$$

$$= 304\,\text{K}$$

$$= 31°\text{C}$$

7 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L.
$R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

These conversions should be applied as needed: 1.00 atm = 101.3 kPa = 1.013 bar.

$$n(N_2) = \frac{PV}{RT} = \frac{1.2 \times 10^2 \times 3.2}{8.314 \times (25+273)} = 0.15\,\text{mol}$$

$$n(\text{He}) = \frac{PV}{RT} = \frac{1.2 \times 101.3 \times 2.5}{8.314 \times (25+273)} = 0.12\,\text{mol}$$

∴ There is a greater amount of nitrogen.

9.4 Stoichiometric calculations involving gases

Worked example: Try yourself 9.4.1

SOLVING MASS–VOLUME STOICHIOMETRY PROBLEMS AT STANDARD LABORATORY CONDITIONS

Calculate the volume of carbon dioxide, in L, produced when 300 g of butane is burned completely in oxygen to produce carbon dioxide and water. The gas volume is measured at SLC.

Thinking	Working
Write a balanced equation for the reaction.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$
Calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_4H_{10}) = \frac{300}{58.12}$ $= 5.16\,\text{mol}$
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{\text{coefficient of } CO_2}{\text{coefficient of } C_4H_{10}} = \frac{8}{2}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{8}{2} \times 5.16$ $= 20.6\,\text{mol}$
Calculate the volume of the unknown substance using: $V = n \times V_m$	$V(CO_2) = 20.6 \times 24.79$ $= 506\,\text{L}$

Worked example: Try yourself 9.4.2

SOLVING MASS–VOLUME STOICHIOMETRY PROBLEMS AT NON-STANDARD CONDITIONS

Calculate the volume of carbon dioxide, in L, produced when 5.00 kg of butane is burned completely in oxygen to produce carbon dioxide and water. The gas volume is measured at 40°C and 400 kPa.

Thinking	**Working**
Write a balanced equation for the reaction.	$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$
Calculate the number of moles of the known substance using: $n = \frac{m}{M}$	$n(C_4H_{10}) = \frac{5.0\times10^3}{58.12}$ $= 86.0$ mol
Find the mole ratio: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$	$\frac{\text{coefficient of } CO_2}{\text{coefficient of } C_4H_{10}} = \frac{8}{2}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{8}{2} \times 86.0$ $= 344$ mol
Express the pressure and temperature in the required units.	$P = 400$ kPa $T = 40 + 273 = 313$ K
Calculate the volume of the unknown substance using: $V = \frac{nRT}{P}$	$V(CO_2) = \frac{344 \times 8.314 \times 313}{400}$ $= 2238$ L $= 2.24 \times 10^3$ L

Worked example: Try yourself 9.4.3

SOLVING GAS VOLUME–VOLUME STOICHIOMETRY PROBLEMS

Methane gas (CH_4) is burned in a gas stove according to the following equation:

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

If 50 mL of methane is burned in air, calculate the volume of CO_2 gas produced under constant temperature and pressure conditions.

Thinking	**Working**
Use the balanced equation to find the mole ratio of the two gases involved.	1 mol of CH_4 gas produces 1 mol of CO_2 gas.
The temperature and pressure are constant, so volume ratios are the same as mole ratios.	1 volume of CH_4 produces 1 volume of CO_2 gas, so 50 mL of CH_4 produces 50 mL of CO_2.

Worked example: Try yourself 9.4.4

SOLVING EXCESS REACTANT STOICHIOMETRY PROBLEMS

65.0 g of butane is burned completely in 200 L of oxygen. The gas volume is measured at SLC. The equation for the reaction is:

$$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$$

a Which reactant is the limiting reactant?

b What is the volume of carbon dioxide formed?

Thinking	**Working**
a Calculate the number of moles of each reactant from $n = \frac{m}{M}$, $n = \frac{V}{V_m}$, or $n = \frac{PV}{RT}$, as appropriate.	$n(C_4H_{10}) = \frac{m}{M}$ $= \frac{65.0}{58.12}$ $= 1.12$ mol $n(O_2) = \frac{V}{V_m}$ $= \frac{200}{24.79}$ $= 8.07$ mol
Use the coefficients of the equation to find the limiting reactant.	The equation shows 2 mol of C_4H_{10} reacts with 13 mol of O_2. Amount of C_4H_{10} required to react all of the O_2: $n(C_4H_{10}) = \frac{2}{13} \times n(O_2)$ $= \frac{2}{13} \times 8.07$ $= 1.24$ mol As there is 1.12 mol of C_4H_{10}, C_4H_{10} is the limiting reactant. It will be completely consumed.
b Find the mole ratio using: $\frac{\text{coefficient of unknown}}{\text{coefficient of known}}$ The limiting reactant is the known substance.	$\frac{\text{coefficient of } CO_2}{\text{coefficient of } C_4H_{10}} = \frac{8}{2}$
Calculate the number of moles of the unknown substance using: $n(\text{unknown}) = \text{mole ratio} \times n(\text{known})$	$n(CO_2) = \frac{8}{2} \times 1.12$ $= 4.47$ mol
Calculate the volume of the unknown using: $V = n \times V_m$	$V(CO_2) = 4.47 \times 24.79$ $= 111$ L

9.4 KEY QUESTIONS

1 The balanced equation shows that 1 mol of propane reacts with 5 mol of oxygen, and produces 3 mol of carbon dioxide.

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

Remember that $n = \frac{m}{M}$ (The molar mass of propane is 44.09 g mol^{-1}.)

Remember that under SLC, 1 mol of any gas has a volume of 24.79 L.

Use the formula $n = \frac{V}{V_m}$ rearranged as $V = n \times V_m$.

a i $n(C_3H_8) = \frac{22}{44.09} = 0.50$ mol

$\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1}$ $\therefore V(O_2) = 5 \times 0.50 \times 24.79 = 62$ L

ii $\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1}$ $\therefore V(CO_2) = 3 \times 0.50 \times 24.79 = 37$ L

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b **i** $n(C_3H_8) = \frac{16.5}{44.09} = 0.374\,\text{mol}$

$\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1} \therefore V(O_2) = 5 \times 0.374 \times 24.79 = 46.4\,\text{L}$

ii $\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1} \therefore V(CO_2) = 3 \times 0.374 \times 24.79 = 27.8\,\text{L}$

c **i** $m(C_3H_8) = 3.40\,\text{kg} = 3.40 \times 10^3\,\text{g}$

$n(C_3H_8) = \frac{3.40 \times 10^3}{44.09} = 77.1\,\text{mol}$

$\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1} \quad \therefore V(O_2) = 5 \times 77.1 \times 24.79 = 9.56 \times 10^3\,\text{L}$

ii $\frac{n(CO_2)}{n(C_3H_8)} = \frac{3}{1} \quad \therefore V(CO_2) = 3 \times 77.1 \times 24.79 = 5.74 \times 10^3\,\text{L}$

2 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

The second step involves using $n = \frac{m}{M}$. The molar mass of octane is $114.22\,\text{g mol}^{-1}$.

The balanced equation shows that 2 mol of octane produces 16 mol of carbon dioxide.

$n(CO_2) = \frac{PV}{RT} = \frac{1.10 \times 101.3 \times 50.0}{8.314 \times (120 + 273)} = 1.71\,\text{mol}$

$\frac{n(C_8H_{18})}{n(CO_2)} = \frac{2}{16} \therefore n(C_8H_{18}) = \frac{2}{16} \times n(CO_2)$

$n(CO_2) = \frac{2}{16} \times 1.71 = 0.213\,\text{mol}$

$\therefore m(C_8H_{18}) = n \times M = 0.213 \times 114.22 = 24.4\,\text{g}$

3 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

The first step involves using $n = \frac{m}{M}$.

The molar mass of hydrogen peroxide is $34.02\,\text{g mol}^{-1}$.

The balanced equation shows that 2 mol of hydrogen peroxide produces 1 mol of oxygen.

$n(H_2O_2) = \frac{m}{M} = \frac{10.0}{34.02} = 0.294\,\text{mol}$

$n(O_2) = \frac{1}{2} \times 0.294 = 0.147\,\text{mol}$

$V(O_2) = \frac{nRT}{P} = \frac{0.147 \times 8.314 \times 303}{91.0} = 4.07\,\text{L}$

4 The balanced equation shows that 2 mol of NO_2 is produced from 2 mol of NO.

Therefore, the volumes of NO_2 produced will be equal to the volume of NO reacted (as long as the temperature and pressure are constant).

$\therefore V(NO) = V(NO_2) = 0.5\,\text{L}$

5 The balanced equation shows that 1 mol of O_2 reacts with 2 mol of CO. Therefore, the volumes of O_2 and CO used will also be in the ratio 1:2 (as long as the temperature and pressure are constant).

$\therefore V(CO) = 150\,\text{mL}$

$\therefore V(O_2) = 75\,\text{mL}$

6 Step 1: Calculate the amount of water, in mol. $n = \frac{m}{M} = \frac{1.00}{18.02} = 0.0555\,\text{mol}$

Step 2: Use the ideal gas law, $PV = nRT$, to calculate the pressure at 220°C. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

$PV = nRT$

$P = \frac{nRT}{V}$

$P = \frac{0.0555 \times 8.314 \times (273 + 220}{0.100}$

$P = 2.27 \times 10^3\ \text{kPa}$

7 Step 1: Determine which reactant is in excess.

$$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$$

The balanced equation shows that 1 mol of CH_4 reacts with 2 mol of O_2 to produce 1 mol of CO_2.

$n(CH_4) = \frac{m}{M} = \frac{23}{16.04} = 1.4\,\text{mol}$

$n(O_2) = \frac{V}{V_m} = \frac{10}{24.79} = 0.40\,\text{mol}$

$n(CH_4)$ required $= \frac{2}{1} \times 0.40 = 0.80\,\text{mol}$

Therefore, the limiting reagent is 0.40 mol of oxygen. That is, methane is in excess.

Step 2: Calculate the volume of CO_2 produced.

$n(CO_2) = \frac{0.40}{2} = 0.20\,\text{mol}$

$V = 0.20 \times 24.79 = 5.0\,\text{L}$

Alternatively, the answer can be derived from Avogadro's law: equal volumes of gas contain equal number of particles. Therefore, 10 L of oxygen will produce 5.0 L of CO_2.

A further alternative is to use the ideal gas law: $V = \frac{nRT}{P}$.

CHAPTER 9 **REVIEW**

1 B.

2 A. Temperature is a measure of the average kinetic energy of the particles. As the temperature increases, the average kinetic energy of the particles also increases. Average kinetic energy is proportional to the average speed of the particles. As average kinetic energy increases, the average speed of the particles also increases.

3 The volume occupied by the atoms or molecules in a gas is much *smaller* than the total volume occupied by the gas. The particles move in rapid, *straight-line* paths and collide with each other and with the walls of the container. The forces between particles are extremely *weak*. The collisions between particles are *elastic*. The average kinetic energy of the particles is *directly* proportional to the temperature of the gas, in units of *K*.

4 **a** As volume is reduced, there is an increase in the frequency of gas particle collisions per unit wall area. This is measured as an increase in pressure.

b When the temperature of a gas is lowered, the average kinetic energy of the particles decreases. The rate of collisions between particles and the walls of the container decreases and particles collide with less force. As pressure is a measure of the force of particle collisions per unit wall area of the container, pressure is found to decrease.

c In a mixture of gases, the particles of each gas are moving and colliding with the walls of the container, independently of one another. Each gas therefore exerts a pressure. As the gases behave independently of one another, total pressure is simply the sum of the individual gas (or partial) pressures.

d When more gas is added to a container, the total number of particles in the container increases. Provided that the volume of the container and the temperature do not change, the collisions of these additional particles means that the total pressure in the container will increase.

5 **a** The pressure inside the container is reduced when some of the gas escapes.

b There are fewer gas particles to collide with one another and with the walls of the container. Pressure is the force exerted by the particles over a defined area, so this will decrease.

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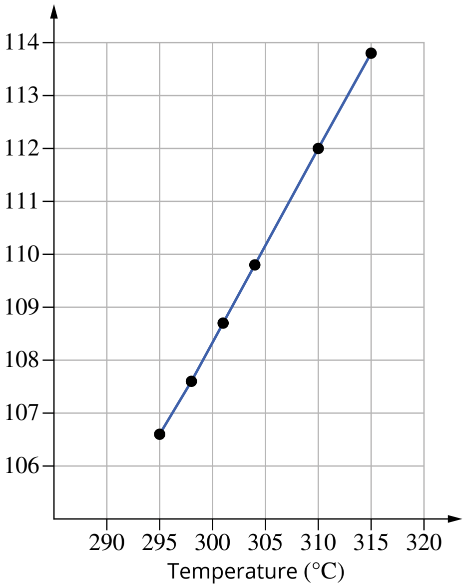
6 **a** A graph of the data reveals a straight-line relationship between P and T. That is, over this range the propane is acting as an ideal gas.

Propane gas pressure versus temperature



b Real gases will deviate from ideal gas behaviour because real gas particles have volume, which becomes significant at high pressures, when the gas volume is significantly compressed.

Alternate response: Real gases deviate from ideal gas behaviour at low temperatures because real gas particles have intermolecular forces of attraction between the particles, which become significant at low temperatures.

7 **a** $V = kT$. For an ideal gas, there is a positive linear relationship between the volume and absolute temperature of the gas.

b The gas particles have negligible volume, and there are no interactions between the particles.

c At 90 K, the oxygen is so compressed it forms a liquid. The particles are so close they begin to interact (dispersion forces become significant), causing oxygen to liquefy.

d Helium gas is monatomic. There are weak dispersion forces between helium atoms, so its behaviour is very close to that of an ideal gas. Oxygen molecules are larger than helium atoms. Oxygen molecules have more electrons and therefore form stronger instantaneous dipoles, resulting in stronger dispersion forces. This accounts for its deviation from ideal gas behaviour at extremely low temperatures.

8 $P_1 = 720\,\text{mmHg}$, $V_1 = 2.00\,\text{L}$, $T_1 = 273 + 65 = 338\text{ K}$; $P_2 = 760\,\text{mmHg}$, $V_2 = ?$, $T_2 = 273 + 25 = 298\text{ K}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$V_2 = \frac{P_1V_1T_2}{T_1P_2}$$

$$V_2 = \frac{720 \times 2.00 \times 298}{338 \times 760}$$

$$V_2 = 1.67\text{ L}$$

9 $P_1 = 1\text{ bar}$, $V_1 = 24\,\text{L}$, $T_1 = 298\,\text{K}$; $P_2 = ?$, $V_2 = 12\,\text{L}$, $T_2 = 318\,\text{K}$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$P_2 = \frac{P_1V_1T_2}{T_1V_2}$$

$$P_2 = \frac{1 \times 24 \times 318}{298 \times 12}$$

$$P_2 = 2.1\text{ bar}$$

10 $P_1 = 1\text{ bar}$, $V_1 = 500\,\text{mL}$, $T_1 = 298\,\text{K}$; $P_2 = 2.5\,\text{bar}$, $V_2 = 220\,\text{mL}$, $T_2 = ?$

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$T_2 = \frac{P_2V_2T_1}{P_1V_1}$$

$$T_2 = \frac{2.5 \times 220 \times 298}{1 \times 500}$$

$$T_2 = 328\text{ K} = 55°\text{C}$$

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11 This problem is solved in two steps.

1 Use Dalton’s law of partial pressures to determine the pressure of the ‘dry gas’.

2 Calculate the gas volume of the ‘dry gas’ at SLC.

Pressure of dry gas:

$P_{tot} = P_1 + P_2 \dots$

$P_{tot} = P(CO_2) + P(H_2O)$

$P(CO_2) = P_{tot} - P(H_2O)$

$= 98.0 - 4.0 = 94.0\,kPa$

Calculate volume of dry CO_2:

$P_1 = 94.0$, $V_1 = 300\,mL$, $T_1 = 273 + 29 = 302\,K$, $P_2 = 101.3\,kPa$, $V_2 = ?$, $T_2 = 273 + 25 = 298\,K$

$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

$V_2 = \frac{P_1V_1T_2}{T_1P_2}$

$V_2 = \frac{94.0 \times 300 \times 302}{298 \times 101.3}$

$V_2 = 282\,mL$

12 Under standard laboratory conditions (SLC), 1 mol of any gas has a volume of 24.79 L. This question involves three steps:

1 Calculate the amount (in mol) of gas at SLC.

2 Find the molar mass (*M*) of the gas.

3 Use $n = \frac{m}{M}$ to find the mass by rearranging the formula as $m = n \times M$.

$n = \frac{V}{V_m} = \frac{50.0}{24.79}$

$= 2.02\,mol$

$M(O_2) = 32.00\,g\,mol^{-1}$

$m(O_2) = 2.02 \times 32.00 = 64.6\,g$

13 This involves using $n = \frac{m}{M}$ and $V = n \times V_m$. You could combine these two formulae to give $V = \frac{m \times V_m}{M}$.

a The molar mass of oxygen is 32.00 g mol^{-1}.

$\therefore V(O_2) = \frac{8.0}{32.00} \times 24.79 = 6.2\,L$

b The molar mass of NO_2 is 46.01 g mol^{-1}.

$\therefore n(NO_2) = \frac{V}{V_m} = \frac{10}{24.79} = 0.40\,mol$

$\therefore m(NO_2) = 0.40 \times 46.01 = 19\,g$

14 B.

Calculate $n(O_2)$:

$n = \frac{m}{M} = \frac{64.0}{32.00}$

$= 2.00\,mol$

Convert given quantities to correct units.

$n = 2.00\,mol$

$R = 8.314\,J\,mol^{-1}\,K^{-1}$

$T = 303\,K$

$V = 25.0\,L$

Calculate $P(O_2)$.

$P = \frac{nRT}{V} = \frac{2.00 \times 8.314 \times 303}{25.0}$

$= 202\,kPa$

15 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

Remember that $n = \frac{m}{M}$.

The molar mass of carbon dioxide is $44.01\,\text{g mol}^{-1}$.

$$V = \frac{nRT}{P} = \frac{10.0 \times 8.314 \times (25 + 273)}{44.01 \times 101.3}$$

$$= 5.56\,\text{L}$$

16 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

Remember that $n = \frac{m}{M}$.

The molar mass of oxygen is $32.00\,\text{g mol}^{-1}$.

$$n(O_2) = \frac{PV}{RT} = \frac{105 \times 10.0}{8.314 \times (20 + 273)} = 0.431\,\text{mol}$$

$$m(O_2) = n \times M = 0.431 \times 32.00 = 13.8\,\text{g}$$

17 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

Remember that $n = \frac{m}{M}$

The molar mass of helium is $4.003\,\text{g mol}^{-1}$.

$$n(\text{He}) = \frac{0.20}{4.003} = 0.050\,\text{mol}$$

$$T = \frac{PV}{nR} = \frac{80 \times 4.0}{0.050 \times 8.314} = 770\,\text{K} = 497°\text{C}$$

18 **a** The molar mass of $CO_2 = 44.01\,\text{g mol}^{-1}$ and $m = n \times M$.

$m(CO_2) = 1.00 \times 44.01 = 44.0\,\text{g}$

b The volume of 1 mol of any gas at SLC = 24.79 L.

c $\text{Density}(CO_2) = \frac{44.0}{24.79}$

$= 1.78\,\text{g L}^{-1}$

d Greater. Gas volume decreases as temperature decreases. Since the mass of gas stays the same, density increases.

19 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

a Conversions are: $P = 100 \times 10^3\,\text{Pa} = 100\,\text{kPa}$

$T = 27 + 273 = 300\,\text{K}$

$$n = \frac{PV}{RT} = \frac{100 \times 5.4}{8.314 \times 300} = 0.22\,\text{mol}$$

b The second step involves using $n = \frac{m}{M}$.

$$M = \frac{m}{n} = \frac{10.0}{0.22} = 46\,\text{g mol}^{-1}$$

20 Container A:

$$n(O_2) = \frac{770 \times 101.3 \times 40.0}{760 \times 8.314 \times (25 + 273)} = 1.66\,\text{mol}$$

Container B:

$$n(O_2) = \frac{390 \times 101.3 \times 0.10}{760 \times 8.314 \times (45 + 273)} = 0.0020\,\text{mol}$$

∴ Container A has more oxygen.

21 The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$.

a Conversions are: $220\,\text{m}^3 = 220 \times 10^3\,\text{L}$

$$n(\text{air particles}) = \frac{PV}{RT} = \frac{100 \times 220 \times 10^3}{8.314 \times (23 + 273)} = 8940\,\text{mol} = 8.94 \times 10^3\,\text{mol}$$

b The second step involves using $n = \frac{m}{M}$.

The molar mass of oxygen is 32.00 g mol^{-1} and of nitrogen is 28.02 g mol^{-1}.

20% of the air in the room is oxygen.

$\therefore n(O_2) = \frac{20}{100} \times 8940 = 1788\,\text{mol}$

$\therefore m(O_2) = 1788 \times 32.00 = 57.2\,\text{kg}$

80% of the air in the room is nitrogen.

$\therefore n(N_2) = \frac{80}{100} \times 8940 = 7152\,\text{mol}$

$\therefore m(N_2) = 7152 \times 28.02 = 200\,\text{kg}$

$\therefore$ Total mass of gas = 57.2 + 200 = 257 kg

22 The general gas equation is required: $PV = nRT$. Temperature should be in kelvin, pressure in kPa and volume in L. $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$

The second step involves using $n = \frac{m}{M}$. The molar mass of octane is 114.22 g mol^{-1}.

The balanced equation shows that 2 mol of octane produces 16 mol of carbon dioxide.

$$n(CO_2) = \frac{PV}{RT} = \frac{1.10 \times 101.3 \times 50.0}{8.314 \times (120+273)} = 1.71\,\text{mol}$$

$$\frac{n(C_8H_{18})}{n(CO_2)} = \frac{2}{16}$$

$$n(C_8H_{18}) = \frac{2}{16} \times n(CO_2) = \frac{2}{16} \times 1.71\,\text{mol} = 0.214\,\text{mol}$$

$\therefore m(C_8H_{18}) = n \times M = 0.214 \times 114.22 = 24.4\,\text{g}$

23 The balanced equation for the reaction is:

$$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(g)$$

Since all measurements are made at the same temperature and pressure, the equation gives us not only the mole ratio with which the gases react with each other, but also the volume ratio. For 10.0 L of ethane to react completely, $\frac{7}{2} \times 10.0$ of oxygen is required.

$\therefore$ 35.0 L of oxygen is required for the complete combustion of 10.0 L of ethane.

24 The balanced equation shows that 1 mol of methane reacts with 2 mol of oxygen to give 1 mol of carbon dioxide and 2 mol of water. At constant temperature and pressure, mole and volume are directly proportional to each other.

a $\therefore V(CH_4)$ used = 5.0 L

b $\therefore V(O_2)$ used = 10 L

c Use the ideal gas law $PV = nRT$ and $n = \frac{m}{M}$. The molar mass of water is 18.02 g mol^{-1}.

$$\frac{n(H_2O)}{n(CO_2)} = \frac{2}{1}$$

$$n(H_2O) = 2 \times n(CO_2)$$

$$n(H_2O) = 2 \times \frac{PV}{RT} = \frac{100 \times 10}{8.314 \times (200+273)} = 0.25\,\text{mol}$$

$\therefore m(H_2O) = 0.25 \times 18.02 = 4.6\,\text{g}$

25 **a** The balanced equation shows that the mole ratio of propane to oxygen is 1 : 5. All temperatures and pressures are the same, therefore mole and volume are directly proportional to each other.

$\therefore$ 80 mL of propane would need 400 mL of oxygen to react completely, and 500 mL of oxygen would need 100 mL of propane to react completely.

$\therefore$ There is excess oxygen, which means that all of the propane will react, and there will be 100 mL of oxygen in excess.

b $V(CO_2) = 3 \times 80$

$= 240\,\text{mL}$

$V(H_2O) = 4 \times 80$

$= 320\,\text{mL}$

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c The original volume of gases = 80 + 500 = 580 mL

The final volume of gases = 100 + 240 + 320 = 660 mL

∴ There was an increase in volume of 80 mL.

26 The balanced equation is:

$$CaCO_3(s) + 4H_2(g) \rightarrow CH_4(g) + Ca(OH)_2(s) + H_2O(g)$$

It shows the mole ratio is 1:4:1:1:1.

a When temperature and pressure are constant, mole and volume are directly proportional to each other.

∴ The volume of methane produced will be $\frac{1}{4}$ the volume of hydrogen used.

∴ The volume of methane will be $100 \times \frac{1}{4} = 25\,L$.

b The ideal gas law is required: $PV = nRT$. Temperature should be in K, pressure in kPa and volume in L. $R = 8.314\,J\,mol^{-1}\,K^{-1}$.

$n(H_2) = \frac{PV}{RT} = \frac{100 \times 100}{8.314 \times (400 + 273)} = 1.79\,mol$

The mole ratio shows: $\frac{n(CaCO_3)}{n(H_2)} = \frac{1}{4}$

∴ $n(CaCO_3) = \frac{1.79}{4} = 0.447\,mol$

$M(CaCO_3) = 100.09\,g\,mol^{-1}$

Since $n = \frac{m}{M}$

$m(CaCO_3) = n \times M$

$= 0.447 \times 100.09 = 44.7\,g$

27 Write the balanced equation:

$$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$$

For each minute of operation:

$n(C_3H_8) = \frac{12.7}{44.09} = 0.288\,mol$

$\frac{n(O_2)}{n(C_3H_8)} = \frac{5}{1}$

$n(O_2)$ required $= 0.288 \times 5 = 1.44\,mol$

$V(O_2)$ required per minute $= 1.44 \times 24.79 = 35.7\,L$

28 Write the balanced equation.

$$C(s) + O_2(g) \rightarrow CO_2(g)$$

Calculate the mass of carbon in coal.

$m(C) = 4.9 \times 10^6 \times \frac{75}{100}$

$= 3.67 \times 10^6$ tonnes

$= 3.67 \times 10^{12}\,g$

Calculate the amount of carbon.

$n(C) = \frac{3.67 \times 10^{12}}{12.01}$

$= 3.06 \times 10^{11}\,mol$

From the equation, 1 mol CO_2 is produced from every 1 mol of C. Calculate the amount of carbon dioxide.

$\frac{n(CO_2)}{n(C)} = 1$

$n(CO_2) = 3.06 \times 10^{11}\,mol$

Calculate the volume of CO_2 at SLC.

$V(CO_2) = 3.06 \times 10^{11} \times 24.79$

$= 7.59 \times 10^{12}\,L$

29 **a** They have equal numbers of molecules. With pressure, volume and temperature the same, *n* will be the same.

b Carbon dioxide. Each CO_2 molecule contains 3 atoms and each O_2 molecule contains 2 atoms. As there are an equal number of molecules of each gas, there are more atoms in the CO_2 sample.

c Carbon dioxide. Density = mass ÷ volume. The volume is the same for each gas, but the mass of CO_2 is greater, so it has the greater density.

30 **a** $C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

b **i** $m(C_2H_5OH) = 1.00\,\text{kg} = 1.00 \times 10^3\,\text{g}$

$n(C_2H_5OH) = \frac{m}{M} = \frac{1.00\times10^3}{46.07} = 21.7\,\text{mol}$

$n(CO_2) = 2 \times n(C_2H_5OH\ ) = 2 \times 21.7 = 43.4\,\text{mol}$

$m(CO_2) = n \times M = 43.4 \times 44.01 = 1911\,\text{g} = 1.91\,\text{kg}$

ii $V(CO_2) = n \times V_\text{m} = 43.4 \times 24.79 = 1076\,\text{L} = 1.08 \times 10^3\,\text{L}$

c density $= \frac{m}{V}$ so $m(C_2H_5OH) = d \times V = 0.785 \times 50.0 \times 1000 = 39\,250\,\text{g} = 39.3\,\text{kg}$

31 Charles' law applies because the Inquiry activity investigated volume and temperature.

Module 2 Review answers

Introduction to quantitative chemistry

Multiple choice

1 C

2 B. 66.7 g. Total product – known reactant = unknown reactant = 78.9 – 12.2 = 66.7 g

3 D

4 B. A is calculated by incorrectly dividing the volume by mass. C is just the mass of the stone. Mass is not the same as density. D incorrectly multiplies the mass by the volume.

5 D. $m = 3.20 \times 1000 = 3200$ g. Volume = mass ÷ density = $\frac{3200}{0.535} = 5980$ L

A is missing unit conversion from kg to g. B is missing unit conversion from kg to g and incorrectly multiplying density by mass. C is incorrectly multiplying density by mass.

6 A. 1 mole of CO_2 has a mass of 44 g = $(12 + 2 \times 16)$

So 6.022×10^{23} molecules have a mass of 44 g

One molecule has a mass of $\frac{44}{6.022 \times 10^{23}} = 7.3 \times 10^{-23}$ g

7 A. m(urea) in 125 g fertiliser = $\frac{58.5}{100} \times 125 = 73.1$ g

% of N in urea = $\frac{M(\text{N})}{M(\text{urea})} \times 100 = \frac{28.02}{60.06} \times 100 = 46.65\%$

So m(N) in 125 g fertiliser = $\frac{46.65}{100} \times 73.1 = 34.1$ g

8 B. 6.022×10^{23} is equivalent to 1 mol of Cl^- ions. The formula for calcium chloride is $CaCl_2$, so 0.5 mol $CaCl_2$ contains 1 mol of Cl^- ions.

The molar mass of $CaCl_2$ is 110.98 g mol^{-1}. So 0.5 mol will have a mass of 55.49 g.

9 C. $n(\text{O}) = 4 \times 0.40 = 1.6$ mol

10 D. $n(\text{Fe}) = 2 \times n(\text{Fe}_2\text{O}_3) = 2 \times 2.0 = 4.0$ mol. $m(\text{Fe}) = n \times M = 4.0 \times 55.85 = 2.2 \times 10^2$ g

11 C. 0.10 mol HCl would react with 0.05 mol Na_2CO_3. Since 0.20 mol Na_2CO_3 is present, Na_2CO_3 is in excess ∴ HCl is limiting.

12 A.
$$\begin{aligned} c_1V_1 &= c_2V_2 \\ 3.0 \times V_1 &= 0.15 \times 750 \\ V_1 &= 0.15 \times \frac{750}{3.0} \\ &= 38 \text{ mL} \end{aligned}$$

13 B.
$$\begin{aligned} c_1V_1 &= c_2V_2 \\ 0.50 \times 20.0 &= 0.20 \times V_2 \\ V_2 &= 0.50 \times \frac{20.0}{0.20} \\ &= 50 \text{ mL} \\ 50 - 20.0 &= 30 \text{ mL of water needs to be added} \end{aligned}$$

14 A. ppb is equivalent to mg in 1000 L. 1.75 kg in 500 ML is equivalent to 1.75 mg in 500 L or 3.50 mg in 1000 L, i.e. 3.5 ppb.

15 C. $n(\text{AgCl}) = \frac{20.37}{143.4} = 0.1421 = n(\text{Cl}^-)$

$m(\text{Cl}^-)$ in 250 mL = $0.1421 \times 35.45 = 5.036$ g

$m(\text{Cl}^-)$ in 1.0 L = $\frac{5.036}{0.2500} = 20.14$ g

16 D. Because in a gas the particles are widely separated, gases are compressible. As the volume of the gas decreases, the particles are forced closer together, so there are more collisions of particles per second with the vessel wall, and so pressure will increase.

17 D. Volume is inversely proportional to pressure under these conditions. If the pressure is halved, the volume will double.

18 B. The volume of a gas is proportional to the temperature (in kelvin) but inversely proportional to the pressure. In this situation the two effects cancel, so the volume remains at 4.0 L.

19 D. $d = \frac{m}{V} = \frac{n \times M}{V}$

$PV = nRT \Rightarrow \frac{n}{V} = \frac{P}{R \times T}$ hence $d = \frac{n \times M}{V} = \frac{P \times M}{R \times T}$

20 C. Since all measurements are made at the same temperature and pressure, the equation gives not only the mole ratio with which the gases react with each other, but also the volume ratio. Oxygen is the limiting reagent, so after the reaction, there will be no oxygen remaining.

10 mL carbon dioxide is produced.

20 mL water vapour is produced.

10 mL methane is unreacted. So the total amount of gas after the reaction is 40 mL.

Short answer

1 **a** $2C_3H_6(g) + 9O_2(g) \rightarrow 6CO_2(g) + 6H_2O(g)$

b $Al(OH)_3(aq) + 3HBr(aq) \rightarrow AlBr_3(aq) + 3H_2O(l)$

c $MgCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$

d $2HNO_3(aq) + 6FeSO_4(aq) + 3H_2SO_4(aq) \rightarrow 3Fe_2(SO_4)_3(aq) + 4H_2O(l) + 2NO(g)$

2 **a** $3CO(g) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3CO_2(g)$

b $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

c $CaBr_2(l) \rightarrow Ca(l) + Br_2(g)$

3 **a** $m(CaCO_3) = 10.00 + 7.85 = 17.85\,g$

b $m(CO_2) = 100.1 - 56.1 = 44.0\,g$

c $m(CaO) = 75.07 - 33.01 = 42.06\,g$

4 **a** $n(Al(NO_3)_3) = \frac{m}{M} = \frac{30.5}{213.01} = 0.143\,mol$

b $n(N) = 3 \times 0.143 = 0.430$ mol

c Total number of mol of atoms = $13 \times 0.143 = 1.86$ mol

Total number of atoms = $1.86 \times 6.022 \times 10^{23} = 1.12 \times 10^{24}$ N atoms

d $\%O = 9 \times \frac{16.00}{213.01} \times 100 = 67.60\%$

5 The empirical formula provides the simplest whole number ratio of atoms in a compound. The amount of each atom is found by using $n = \frac{m}{M}$, where m is the mass in grams and M is the molar mass in g mol^{-1}. Use a periodic table to find the molar masses of W and S.

	W	S
Mass	1.84 g	0.64 g
Molar mass	183.9 g mol^{-1}	32.07 g mol^{-1}
Amount using $n = \frac{m}{M}$	$n = \frac{1.84}{183.9} = 0.01$ mol	$n = \frac{0.64}{32.07} = 0.02$ mol
Divide all by the smallest amount.	$\frac{0.01}{0.01} = 1$	$\frac{0.02}{0.01} = 2$
Round off to whole numbers.	1	2

So the empirical formula is WS_2.

6 **a** $m(O)$ in compound = $2.127 - 1.704 = 0.423$ g

$n(Cu) : n(O) = \frac{1.704}{63.55} : \frac{0.423}{16.00} = 0.0268 : 0.0264 = 1 : 1$

So empirical formula is CuO.

b If the sample was a mixture of CuO and Cu_2O, the ratio of Cu to O would be greater and so the mass of Cu extracted would be greater than 1.704 g.

7 **a** $m(O)$ in the compound = $0.907 - 0.753 = 0.154\,g$

$n(O)$ in the compound = $\frac{0.154}{16.00} = 0.00963\,mol$

Because the empirical formula is X_2O:

$n(X)$ in the compound = $2 \times 0.00963 = 0.0193\,mol$

Because $n = \frac{m}{M}$, $M(X) = \frac{m}{n} = \frac{0.752}{0.0193} = 39.1\,g\,mol^{-1}$

b The metal must be potassium.

8 %N in $KNO_3 = \frac{14.01}{101.11} \times 100 = 13.9\%$

So m(N) in 65.0 g KNO_3 = 9.01 g

%N in $(NH_4)_2SO_4 = \frac{28.02}{132.15} \times 100 = 21.2\%$

So $m((NH_4)_2SO_4)$ that contains 9.01 g N $= \frac{9.01}{21.2} \times 100 = 42.5$ g

9 **a** $m(\text{Hg}) = c \times V$

$= 20 \times 0.100$

$= 2.0$ mg

$= 0.020$ g

$\%(\text{w/w}) = \frac{m(\text{Hg})}{m(\text{fish})} \times 100$

$= \frac{0.020}{3.0} \times 100$

$= 0.67\%$

b concentration in ppm $= \frac{0.020}{3.0} \times 1 \times 10^6$

$= 6.7 \times 10^3$ ppm

10 **a** **i** $P = \frac{16500}{101.3} = 163$ atm

ii $P = 163 \times 760 = 1.24 \times 10^5$ mmHg

b $n(\text{He}) = \frac{PV}{RT} = \frac{16500 \times 25.0}{8.314 \times 288} = 172$ mol

$m(\text{He}) = nM = 172 \times 4.003 = 690$ g

c $V(\text{He}) = nV_m = 172 \times 24.79 = 4.27 \times 10^3$ L

d Total $V(\text{He}) = \frac{nRT}{P} = \frac{172 \times 8.314 \times 303}{108} = 4.01 \times 10^3$ L

$\Rightarrow$ number of balloons $= \frac{4.01 \times 10^3}{6.5} = 617$

Extended response

1 **a** $Ba(ClO_3)_2(s) \rightarrow BaCl_2(s) + 3O_2(g)$

b 7.25 – 4.96 = 2.29 g of oxygen

c 7.25 – 5.00 = 2.25 g more of barium chlorate would be required to produce 4.96 g of barium chloride.

d If you tripled the amount of barium chlorate, the total mass of the products would also triple, assuming the reaction goes to completion.

e 3.72 + 1.72 = 5.44; 7.25 – 5.44 = 1.81 g of barium chlorate did not decompose.

2 **a** $MgCO_3(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) + CO_2(g)$

b $n(CO_2) = \frac{V}{V_m}$

$= \frac{2.60}{24.79}$

$= 0.105$ mol

c $n(MgCO_3) = n(CO_2) = 0.105$ mol

$m(MgCO_3) = n \times M$

$= 0.105 \times 84.32$

$= 8.84$ g

$\%(\text{w/w}) = \frac{m(MgCO_3)}{m(\text{rock})} \times 100$

$= \frac{8.84}{12.8} \times 100$

$= 69.1\%$

d $n(\text{HCl}) = 2 \times n(\text{CO}_2)$
$= 2 \times 0.105$
$= 0.210\,\text{mol}$
$V(\text{HCl}) = \frac{n}{c}$
$= \frac{0.210}{0.50}$
$= 0.42\,\text{L}$

e **i** $\text{CO}_2\text{(g)} + \text{Ca(OH)}_2\text{(aq)} \rightarrow \text{CaCO}_3\text{(s)} + \text{H}_2\text{O(l)}$

ii $n(\text{CaCO}_3) = n(\text{CO}_2) = 0.105\,\text{mol}$
$m(\text{CaCO}_3) = n \times M$
$= 0.105 \times 100.09$
$= 10.5\,\text{g}$

3 **a** $n(\text{Pb(NO}_3)_2) = \frac{m}{M}$
$= \frac{9.80}{331.2}$
$= 0.0296\,\text{mol}$
$c = \frac{n}{V}$
$= \frac{0.0296}{0.0500}$
$= 0.592\,\text{mol L}^{-1}$

b The amount of $\text{Pb(NO}_3)_2$, in mol, has not changed, so still 0.0296 mol. Volume is now 80.0 mL.
$c = \frac{n}{V}$
$= \frac{0.0296}{0.0800}$
$= 0.370\,\text{mol L}^{-1}$

c **i** $\text{Pb(NO}_3)_2\text{(aq)} + 2\text{NaI(aq)} \rightarrow \text{PbI}_2\text{(s)} + 2\text{NaNO}_3\text{(aq)}$

ii $n(\text{Pb(NO}_3)) = 0.0296\,\text{mol}$
$n(\text{NaI}) = c \times V$
$= 0.650 \times 0.050$
$= 0.0325\,\text{mol}$

iii The amount of $\text{Pb(NO}_3)_2$, in mol, that would react with NaI is $\frac{0.0325}{2} = 0.0163\,\text{mol}$. Since 0.0296 mol of $\text{Pb(NO}_3)_2$ is present, $\text{Pb(NO}_3)_2$ is in excess $\therefore$ NaI is limiting.

iv $n(\text{PbI}_2) = \frac{n(\text{NaI})}{2}$
$= \frac{0.0325}{2}$
$= 0.0163\,\text{mol}$
$m(\text{PbI}_2) = n \times M$
$= 0.0163 \times 461.0$
$= 7.49\,\text{g}$

4 **a** The volume decreases non-linearly as the pressure increases.

b $PV = k$ or $P \propto \frac{1}{V}$

c Choose a convenient initial condition, say $P_1 = 70\,\text{kPa}$, $V_1 = 62\,\text{mL}$, $P_2 = 160\,\text{kPa}$, $V_2 = ?$
$P_1V_1 = P_2V_2$
$V_2 = \frac{P_1V_1}{P_2}$
$V_2 = \frac{70 \times 62}{160}$

$V_2 = 27\,\text{mL}$ (This can be verified by extrapolation of the graph.)

d The relationship $P \propto \frac{1}{V}$ holds for an ideal gas of negligibly sized particles that have no interactions. At high pressures, the gas particle size does become significant, as do the interactions between particles. Consequently, the inverse relationship will no longer hold.

e $PV = nRT$

where: $P = 70\,\text{kPa}$, $V = 62 \times 10^{-3}\,\text{L}$, $R = 8.314\,\text{J mol}^{-1}\,\text{K}^{-1}$, $T = 298\,\text{K}$, $n = ?$

$$\begin{aligned} PV &= nRT \\ n &= \frac{PV}{RT} \\ &= \frac{70 \times 62 \times 10^{-3}}{8.314 \times 298} \\ &= 1.8 \times 10^{-3}\,\text{mol} \end{aligned}$$

f If CO_2 was replaced by H_2 in this particular experiment, close to identical results should be obtained. At these conditions of low pressure and moderate temperatures, all gases approximate the behaviour of an ideal gas. That is, the particle size is negligible compared with the volume of the gas, and the interactions between particles are insignificant. At high pressures, H_2 will behave more like an ideal gas due to its weaker dispersion forces between molecules and smaller molecular size compared with CO_2.

5 **a** $P_4(s) + 3KOH(aq) + 3H_2O(l) \rightarrow PH_3(g) + 3KH_2PO_2(aq)$

b
$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{5.948}{123.88} \\ &= 0.04801\,\text{mol} \end{aligned}$$

c
$$\begin{aligned} n &= \frac{m}{M} \\ &= \frac{4.500}{56.11} \\ &= 0.08020\,\text{mol} \end{aligned}$$

The amount of P_4, in mol, that would react with KOH is $\frac{0.08020}{3} = 0.02673\,\text{mol}$. Since 0.04801 mol of P_4 was present, P_4 was in excess $\therefore$ KOH was limiting.

d
$$\begin{aligned} n(P_4)\text{excess} &= n(P_4)\text{total} - n(P_4)\text{limiting} \\ &= 0.04801 - 0.02673 \\ &= 0.02128\,\text{mol} \end{aligned}$$

e
$$\begin{aligned} n(PH_3) &= \frac{n(KOH)}{3} \\ &= \frac{0.04801}{3} \\ &= 0.01601\,\text{mol} \\ V(PH_3) &= n \times V_m \\ &= 0.01601 \times 24.79 \\ &= 0.3968\,\text{L} \end{aligned}$$

f
$$\begin{aligned} c &= \frac{n}{V} \\ &= \frac{0.08020}{0.2500} \\ &= 0.3208\,\text{mol L}^{-1} \end{aligned}$$

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Chapter 10 Chemical reactions

10.1 Chemical change

Worked example: Try yourself 10.1.1

BALANCING EQUATIONS

Aluminium metal reacts in hydrochloric acid (a hydrogen chloride solution) to form aluminium chloride solution and hydrogen gas. Write a balanced equation for this reaction.

Thinking	Working
Write an equation in words.	aluminium + hydrochloric acid → aluminium chloride + hydrogen.
Write the correct formulae for each substance and add states.	$Al(s) + HCl(aq) \rightarrow AlCl_3(aq) + H_2(g)$
Balance the equation by adjusting coefficients.	$2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$

10.1 KEY QUESTIONS

1 C. The liquid formed is still gold and can easily be solidified back to solid gold.

2 **a** chemical
  b physical
  c physical
  d chemical
  e physical
  f chemical

3 **a** reactants: carbon dioxide and water; products: glucose and oxygen
  b reactant: water; products: hydrogen gas and oxygen gas
  c reactant: calcium carbonate; products: calcium oxide and carbon dioxide
  d reactants: carbon and oxygen; product: carbon dioxide

4 **a** sulfur + oxygen gas → sulfur dioxide gas
  $S(s) + O_2(g) \rightarrow SO_2(g)$
  b silver nitrate + potassium iodide → silver iodide + potassium nitrate
  $AgNO_3(aq) + KI(aq) \rightarrow AgI(s) + KNO_3(aq)$

5 **a** Stage one is a physical change. The solid turns to a liquid, but the bonds in the sugar molecules have not changed.
  b Stage two is a chemical change. The bonds in the sugar molecules have been broken and new substances have been formed.

6 carbon + oxygen → carbon dioxide. The mass of oxygen required was $132\,g - 36\,g = 96\,g$

7 **a** The mass of the bag and contents will slowly decrease because a gas is being evolved from the bag. Gases have mass.
  b The mass of the bag and contents will not change because the gas has been trapped in the bag. No products are lost from the bag.

8 **a** unpleasant odour, change of colour of flesh, bloating of corpse that indicates a gas has been produced
  b light from candle or mass loss
  c surface of the trophy has a film on it, colour change on the surface of the trophy

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10.2 Synthesis reactions

10.2 KEY QUESTIONS

1 A. Glucose is a larger molecule formed from the atoms present in carbon dioxide and water.

2 Synthesis reactions might use elements as reactants, but reactions that produce elements are not synthesis reactions.

3 **a** hydrogen + chlorine → hydrogen chloride

b $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

c 40 molecules: 20 molecules of hydrogen gas and 20 molecules of chlorine gas are required.

d The mass of chlorine gas required will be much greater than the mass of hydrogen required. The number of molecules is the same, but the mass of chlorine molecules is greater than that of hydrogen molecules.

4 **a** sulfur dioxide + oxygen → sulfur trioxide

b $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

c The mass of SO_3 will equal the combined mass of SO_2 and O_2 that reacted.

d 200 g of SO_3 will not be not formed, because the reactants are not used in the mole ratio of 2:1.

$n(SO_2) = \frac{m}{M} = \frac{100}{64.07} = 1.56\,\text{mol}$

$n(O_2) = \frac{m}{M} = \frac{100}{32.00} = 3.13\,\text{mol}$

So, to react all of the SO_2 you will require:

$\frac{1}{2} \times n(SO_2)$ of O_2

$\frac{1}{2} \times 1.56 = 0.780$ mol

As there is 3.13 mol available, O_2 is in excess.

10.3 Decomposition reactions

10.3 KEY QUESTIONS

1 C. It shows one reactant forming more than one product. All other alternatives involve more than one reactant.

2 $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

3 No new products are formed when water boils. There is only a change of state, which is a physical change. The intermolecular bonds between water molecules can reform easily. The water has not chemically changed.

4 $MgCl_2(l) \rightarrow Mg(l) + Cl_2(g)$

5 **a** $H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g)$

b Once the container is opened, the carbonic acid starts to react. As its concentration drops and carbon dioxide gas escapes the container, the soft drink becomes less gaseous.

6 **a** $MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

b metal carbonate → metal oxide + carbon dioxide

10.4 Combustion reactions

Worked example: Try yourself 10.4.1

WRITING EQUATIONS FOR COMPLETE COMBUSTION OF HYDROCARBON FUELS

Write the equation, including state symbols, for the complete combustion of liquid hexane (C_6H_{14}).

Thinking	Working
Add oxygen as a reactant, and carbon dioxide and water as the products.	$C_6H_{14} + O_2 \rightarrow CO_2 + H_2O$
Balance carbon and hydrogen atoms, based on the formula of the hydrocarbon.	$C_6H_{14} + O_2 \rightarrow 6CO_2 + 7H_2O$
Find the total of the number of oxygen atoms on the product side.	Total O = (6 × 2) + 7 = 19
If this is an odd number, multiply all of the coefficients in the equation by two, except for the coefficient of oxygen.	$2C_6H_{14} + O_2 \rightarrow 12CO_2 + 14H_2O$

Balance oxygen by adding the appropriate coefficient to the reactant side of the equation.	$2C_6H_{14} + 19O_2 \rightarrow 12CO_2 + 14H_2O$
Add state symbols.	$2C_6H_{14}(l) + 19O_2(g) \rightarrow 12CO_2(g) + 14H_2O(l)$

Worked example: Try yourself 10.4.2

WRITING EQUATIONS FOR COMBUSTION REACTIONS OF ALCOHOLS

Write the equation, including state symbols, for the complete combustion of liquid methanol (CH_3OH).

Thinking	**Working**
Add oxygen as a reactant, and carbon dioxide and water as the products.	$CH_3OH + O_2 \rightarrow CO_2 + H_2O$
Balance carbon and hydrogen atoms, based on the formula of the alcohol.	$CH_3OH + O_2 \rightarrow CO_2 + 2H_2O$
Total the number of oxygen atoms on the product side. Then subtract the one oxygen atom in the alcohol molecule from the total number of oxygen atoms on the product side.	Total O on product side $= (1 \times 2) + (2 \times 1)$ $= 4$ Total O on product side – 1 in alcohol $= 4 - 1 = 3$
If this number is an odd number, multiply all of the coefficients in the equation by two, except for the coefficient of oxygen.	$2CH_3OH + O_2 \rightarrow 2CO_2 + 4H_2O$
Balance oxygen by adding the appropriate coefficient to the reactant side of the equation.	$2CH_3OH + 3O_2 \rightarrow 2CO_2 + 4H_2O$
Add state symbols.	$2CH_3OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 4H_2O(l)$

Worked example: Try yourself 10.4.3

WRITING EQUATIONS FOR INCOMPLETE COMBUSTION OF FUELS

Write an equation, including state symbols, for the incomplete combustion of liquid methanol (CH_3OH) to form carbon monoxide and water vapour.

Thinking	**Working**
Add oxygen as a reactant, and carbon monoxide and water as the products.	$CH_3OH + O_2 \rightarrow CO + H_2O$
Balance the carbon and hydrogen atoms, based on the formula of the hydrocarbon.	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
Balance oxygen by adding the appropriate coefficient to the reactant side of the equation.	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
If oxygen gas has a coefficient that is half of a whole number, multiply all of the coefficients in the equation by two.	$CH_3OH + O_2 \rightarrow CO + 2H_2O$
Add state symbols.	$CH_3OH(l) + O_2(g) \rightarrow CO(g) + 2H_2O(g)$

10.4 KEY QUESTIONS

1 D. The oxygen atoms should be left to last because oxygen appears in both products.

2 $2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$

3 $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$

4 **a** A bellows is used to increase the supply of oxygen to the coals, ensuring rapid and complete combustion is occurring.

b Air flow into a closed caravan is often low. Opening the door increases the supply of oxygen, often leading to increased combustion.

c A kerosene lamp often operates with a low oxygen supply. The incomplete combustion leads to carbon as one of the products.

d The temperature inside a haystack can be very high, and dry grass is a good fuel. However, the lack of oxygen penetrating to the heat source limits the degree of combustion.

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5 **a** If air is limited, incomplete combustion occurs and the toxic gas carbon monoxide forms.

b The reactants are diesel and oxygen; the products are carbon dioxide and water.

c Diesel is reacting with oxygen to form carbon dioxide and water. The mass of diesel and oxygen consumed will equal the mass of carbon dioxide and water produced.

10.5 Precipitation reactions

Worked example: Try yourself 10.5.1

PREDICTING THE PRODUCTS OF A PRECIPITATION REACTION

What precipitate, if any, will be produced when solutions of sodium sulfide (Na_2S) and copper(II) nitrate ($Cu(NO_3)_2$) are added together? You will need to refer to the solubility tables (Tables 10.5.1 and 10.5.2, pages 333 and 334 of the Student Book) to complete this question.

Thinking	**Working**
Identify which ions are produced by each of the ionic compounds in the mixture.	The ions produced are $Na^+(aq)$, $S^{2-}(aq)$, $Cu^{2+}(aq)$ and $NO_3^-(aq)$.
Identify which two new combinations of positive and negative ions are possible in the mixture of the solutions.	$Na^+(aq)$ and $NO_3^-(aq)$ $Cu^{2+}(aq)$ and $S^{2-}(aq)$
Use the solubility tables to check which, if any, of these combinations will produce an insoluble compound.	Compounds containing sodium ions or nitrate ions are soluble, so sodium nitrate will not form a precipitate. Copper sulphide is insoluble, so copper(II) sulfide will form as a precipitate.

Worked example: Try yourself 10.5.2

WRITING EQUATIONS FOR PRECIPITATION REACTIONS

Write a full equation for the reaction between copper(II) sulfate and sodium hydroxide, in which the precipitate is copper(II) hydroxide. Identify the spectator ions in this reaction.

Thinking	**Working**
Write an incomplete, unbalanced equation showing the reactants and the precipitate product. Include symbols of state.	$CuSO_4(aq) + NaOH(aq) \rightarrow Cu(OH)_2(s)$
Add to the equation the formula of the other compound formed in the reaction.	$CuSO_4(aq) + NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Balance the equation.	$CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)$
Write the formulae of the ions that do not form a precipitate in the reaction. These are the spectator ions.	$Na^+(aq)$ and $SO_4^{2-}(aq)$ are spectator ions.

Worked example: Try yourself 10.5.3

DETERMINING WHETHER IONIC COMPOUNDS ARE SOLUBLE OR INSOLUBLE IN WATER

Is ammonium phosphate ($(NH_4)_3PO_4$) soluble or insoluble in water? You will need to refer to the solubility tables to complete this question (Tables 10.5.1 and 10.5.2 on pages 333 and 334 of the Student Book).

Thinking	**Working**
Identify the ions that are present in the ionic compound.	Ammonium (NH_4^+) and phosphate (PO_4^{3-})
Check the solubility tables to see if compounds containing the cation are usually soluble or insoluble in water.	Compounds containing ammonium ions are soluble.
Check the solubility tables to see if compounds containing the anion are usually soluble or insoluble in water.	Ammonium phosphate is a listed water-soluble, phosphate ion–containing compound.

10.5 KEY QUESTIONS

1 B. Most carbonates are insoluble in water.

2 A, B, D, E, H

3 A, C, D, F, H

4 **a** a precipitate of silver carbonate
- **b** a precipitate of lead(II) hydroxide
- **c** a precipitate of magnesium sulfide
- **d** no precipitate

  Silver, lead and magnesium ions are not found in the solubility table. However, the anions that each of these are combined with in these questions all form compounds that are usually insoluble. It is also worth noting that compounds containing sodium and nitrate ions are usually soluble.

  The iron(II) ion is also not found in the solubility table. However, iron(II) does not appear in the 'exceptions' column of the table for either nitrates or sulfates, both of which generally form soluble compounds.

5 Chemicals containing the ions Na^+, K^+, NH_4^+ and NO_3^- almost never form a precipitate.
- **a** **i** magnesium sulfide
  - **ii** silver chloride
  - **iii** aluminium hydroxide
  - **iv** magnesium hydroxide
- **b** **i** $K_2S(aq) + MgCl_2(aq) \rightarrow MgS(s) + 2KCl(aq)$
  - **ii** $CuCl_2(aq) + 2AgNO_3(aq) \rightarrow 2AgCl(s) + Cu(NO_3)_2(aq)$
  - **iii** $AlCl_3(aq) + 3KOH(aq) \rightarrow Al(OH)_3(s) + 3KCl(aq)$
  - **iv** $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

6
- **a** **i** $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
  - **ii** $Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$
- **b** **i** $CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)$
  - **ii** $Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)$
- **c** **i** $(NH_4)_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2NH_4Cl(aq)$
  - **ii** $SO_4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO_4(s)$
- **d** **i** $K_2S(aq) + Pb(NO_3)_2(aq) \rightarrow PbS(s) + 2KNO_3(aq)$
  - **ii** $S^{2-}(aq) + Pb^{2+}(aq) \rightarrow PbS(s)$
- **e** **i** $3CaCl_2(aq) + 2Na_3PO_4(aq) \rightarrow 6NaCl(aq) + Ca_3(PO_4)_2(s)$
  - **ii** $3Ca^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Ca_3(PO_4)_2(s)$
- **f** **i** $2NaOH(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + Pb(OH)_2(s)$
  - **ii** $2OH^-(aq) + Pb^{2+}(aq) \rightarrow Pb(OH)_2(s)$

7
- **a** Na^+/NO_3^-
- **b** Na^+/SO_4^{2-}
- **c** NH_4^+/Cl^-
- **d** K^+/NO_3^-
- **e** Na^+/Cl^-
- **f** Na^+/NO_3^-

10.6 Reactions of acids and bases

Worked example: Try yourself 10.6.1

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND REACTIVE METALS

Write an ionic equation for the reaction that occurs when aluminium is added to a dilute solution of hydrochloric acid.

Thinking	Working
Write the general reaction. Identify the products formed.	Acid + reactive metal → salt + hydrogen Hydrogen gas and aluminium chloride solution are produced.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: aluminium is a solid, $Al(s)$. Hydrochloric acid is dissociated, forming $H^+(aq)$ and $Cl^-(aq)$ ions. Products: hydrogen gas, H_2. Aluminium chloride is dissociated into $Al^{3+}(aq)$ and $Cl^-(aq)$ ions.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Al(s) \rightarrow Al^{3+}(aq) + Cl^-(aq) + H_2(g)$
Identify the spectator ions.	$Cl^-(aq)$
Rewrite equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

Worked example: Try yourself 10.6.2

WRITING AN IONIC EQUATION FOR AN ACID BASE REACTION

Write an ionic equation for the reaction that occurs when sulfuric acid is added to potassium hydroxide solution.

Thinking	Working
Write the general reaction. Identify the products formed.	Acid + metal hydroxide → salt + water A solution of potassium sulfate and water is formed.
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: H_2SO_4 is dissociated in solution, forming $H^+(aq)$ and $SO_4^{2+}(aq)$. KOH is dissociated in solution, forming $K^+(aq)$ and $OH^-(aq)$. Products: potassium sulfate is dissociated and exists as $K^+(aq)$ and $SO_4^{2-}(aq)$. Water is a molecular compound and its molecular formula is $H_2O(l)$.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + SO_4^{2-}(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + SO_4^{2-}(aq) + H_2O(l)$
Identify the spectator ions: the ions that have an (aq) state both as a reactant and as a product.	$K^+(aq)$ and $SO_4^{2-}(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to number of atoms of each element and charge.	$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

Worked example: Try yourself 10.6.3

WRITING IONIC EQUATIONS FOR REACTIONS BETWEEN ACIDS AND METAL CARBONATES

What products are formed when a solution of hydrochloric acid is added to a solution of sodium carbonate? Write an ionic equation for this reaction.

Thinking	**Working**
Write the general reaction. Identify the products.	acid + metal carbonate → salt + water + carbon dioxide The products of this reaction are sodium chloride in solution, water and carbon dioxide gas
Identify the reactants and products. Indicate the state of each, i.e. (aq), (s), (l) or (g).	Reactants: hydrochloric acid is dissociated in solution, forming $H^+(aq)$ and $Cl^-(aq)$ ions. Sodium carbonate dissociates in solution and exists as $Na^+(aq)$ and $CO_3^{2-}(aq)$ ions. Products: sodium chloride is dissociated into $Na^+(aq)$ and $Cl^-(aq)$ ions. Water has the formula $H_2O(l)$. Carbon dioxide has the formula $CO_2(g)$.
Write the equation showing all reactants and products. (There is no need to balance the equation yet.)	$H^+(aq) + Cl^-(aq) + Na^+(aq) + CO_3^{2-}(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l) + CO_2(g)$
Identify the spectator ions.	$Na^+(aq) + Cl^-(aq)$
Rewrite the equation without the spectator ions. Balance the equation with respect to the number of atoms of each element and charge.	The balanced equation with the spectator ions omitted is: $2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$

10.6 KEY QUESTIONS

1 C. Acetic acid will neutralise the alkaline venom.

2 **a** $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$
   $Mg(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$
   b $Ca(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2(g)$
   $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   c $Zn(s) + 2CH_3COOH(aq) \rightarrow Zn(CH_3COO)_2 + H_2(g)$
   $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
   d $2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)$
   $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

3 **a** magnesium sulfate
   b calcium chloride
   c zinc acetate
   d aluminium chloride

4 **a** **i** $ZnO(s) + H_2SO_4(aq) \rightarrow ZnSO_4(aq) + H_2O(l)$
   ii $ZnO(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2O(l)$
   b **i** $Ca(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2(g)$
   ii $Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$
   c **i** $Cu(OH)_2(s) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l)$
   ii $Cu(OH)_2(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + 2H_2O(l)$
   d **i** $Mg(HCO_3)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l) + 2CO_2(g)$
   ii $Mg(HCO_3)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l) + 2CO_2(g)$
   e **i** $SnCO_3(s) + H_2SO_4(aq) \rightarrow SnSO_4(aq) + H_2O(l) + CO_2(g)$
   ii $SnCO_3(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2O(l) + CO_2(g)$

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5 The first step, as in the development of any equation, is to write the correct chemical formulae for each of the chemicals involved.

- **a** $2KOH(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + 2H_2O(l)$
  $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
- **b** $NaOH(aq) + HNO_3(aq) \rightarrow NaNO_3(aq) + H_2O(l)$
  $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
- **c** $Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$
  $Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
- **d** $CuCO_3(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l) + CO_2(g)$
  $CuCO_3(s) + 2H^+(aq) \rightarrow Cu^{2+}(aq) + H_2O(l) + CO_2(g)$
- **e** $KHCO_3(aq) + HF(aq) \rightarrow KF(aq) + H_2O(l) + CO_2(g)$
  $HCO_3^-(aq) + H^+(aq) \rightarrow H_2O(l) + CO_2(g)$
- **f** $Zn(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2(g)$
  $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
- **g** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$
  $CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$
- **h** $NaHCO_3(s) + CH_3COOH(aq) \rightarrow NaCH_3COO(aq) + H_2O(l) + CO_2(g)$
  $NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)$

6 **a** The product of zinc chloride would be formed from the reaction between zinc metal and hydrochloric acid.

- **b** $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

7 **a** The product lithium sulfate would indicate the reaction was between sulfuric acid and lithium hydroxide.

- **b** $H_2SO_4(aq) + 2LiOH(aq) \rightarrow Li_2SO_4(aq) + 2H_2O(l)$

10.7 Removing toxins from food

10.7 KEY QUESTIONS

1 C.

2 **a** Grinding the food will increase the surface area. It will be easier to remove each toxin particle if it is not protected by other parts of the food.

- **b** Some of the toxins will be relatively soluble in water. Sitting the ground-up food in a water supply will allow time for soluble toxins to wash from the food.
- **c** Cooking the food will lead to the decomposition reaction of some of the toxins into less harmful substances.

3 Fire leads to more effective seeding of cycads. The eventual number of cycad bushes in a particular area will be greater if the bushes are regularly exposed to fire.

4 Coffee beans can be placed in a solvent. The caffeine dissolves in the solvent. The solvent and caffeine are then washed from the coffee beans.

CHAPTER 10 REVIEW

1 **a** change of colour or odour

- **b** unpleasant odour, curdled appearance or discolouration
- **c** change in colour of egg white from colourless to white
- **d** exhaust fumes or soot or sound emitted
- **e** rapid volume change or temperature change

2 Gold – physical change. The gold melts to a liquid, then solidifies to form a gold bar. The substance stays as gold.

Mercury(II) oxide – chemical change. The mercury(II) oxide decomposes to mercury and oxygen gas.

Aluminium – chemical change. When the aluminium burns, it reacts with oxygen to form aluminium oxide.

3 **a** combustion of petrol, reaction of nitrogen in air with oxygen, or reactions that occur in the battery

- **b** action of baking powder to create a gas, cooking of an egg, or meat turning brown when cooked

4 Aluminium metal can react with oxygen gas to form solid aluminium oxide. The ratio of aluminium atoms to oxygen molecules required is 4:3.

5 **a** magnesium + sulfur → magnesium sulfide
$Mg(s) + S(s) \rightarrow MgS(s)$
b zinc + copper(II) sulfate → copper + zinc sulfate
$Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$
c magnesium carbonate → magnesium oxide + carbon dioxide
$MgCO_3(s) \rightarrow MgO(s) + CO_2(g)$

6 A. Reactants X and Y are combining to synthesise Z.

7 **a** nitrogen + oxygen → nitrogen monoxide
$N_2(g) + O_2(g) \rightarrow 2NO(g)$
b nitrogen monoxide + oxygen → nitrogen dioxide
$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

8 **a** thermal decomposition
b $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

9 **a** electrolysis
b $2Al_2O_3(l) \rightarrow 4Al(l) + 3O_2(g)$

10 **a** photolysis
b $2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$

11 C.

12 **a** methane + oxygen → carbon dioxide and water
b $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)$
c $2CH_4(g) + 3O_2(g) \rightarrow 2CO(g) + 4H_2O(l)$
d The Bunsen burner has an air valve. When the valve is open, complete combustion occurs, producing a hot blue flame. When the valve is closed, the flame is orange and produces less heat.

13 **a** Heaters such as this have an air intake valve or adjustment. When the valve is open, oxygen is readily available and combustion is relatively complete.
b If combustion is limited by air flow, less energy is released, but the fuel lasts longer. More soot is produced.
c Carbon dioxide, carbon monoxide and water are also produced in the reaction and escape through the chimney flue, as this is not a closed system. Once all the products are taken into account, the mass of the products is found to equal that of the reactants.

14 **a** true
b true
c false
d false
e true
f false

15

	NaOH	**KBr**	**NaI**	**$MgSO_4$**	**$BaCl_2$**
$Pb(NO_3)_2$	$Pb(OH)_2$	$PbBr_2$*	PbI_2	$PbSO_4$	$PbCl_2$*
KI	–	–	–	–	–
$CaCl_2$	$Ca(OH)_2$*	–	–	$CaSO_4$*	–
Na_2CO_3	–	–	–	$MgCO_3$	$BaCO_3$
Na_2S	–	–	–	MgS	BaS

*These compounds are only slightly soluble.

16 According to Tables 10.5.1 and 10.5.2 (pages 333 and 334 of the Student Book), Fe^{2+} would precipitate in the presence of PO_4^{3-} and S^{2-} to produce $Fe_3(PO_4)_2$ and FeS, respectively.

17 **a** barium sulfate
b none
c lead(II) sulfate
d none

18 **a** $\mathrm{NH_4Cl(aq) + AgNO_3(aq) \rightarrow NH_4NO_3(aq) + AgCl(s)}$

$\mathrm{Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)}$

b $\mathrm{FeCl_2(aq) + Na_2S(aq) \rightarrow FeS(s) + 2NaCl(aq)}$

$\mathrm{Fe^{2+}(aq) + S^{2-}(aq) \rightarrow FeS(s)}$

c $\mathrm{Fe(NO_3)_3(aq) + 3KOH(aq) \rightarrow 3KNO_3(aq) + Fe(OH)_3(s)}$

$\mathrm{Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)}$

d $\mathrm{CuSO_4(aq) + 2NaOH(aq) \rightarrow Cu(OH)_2(s) + Na_2SO_4(aq)}$

$\mathrm{Cu^{2+}(aq) + 2OH^-(aq) \rightarrow Cu(OH)_2(s)}$

e $\mathrm{Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaNO_3(aq)}$

$\mathrm{Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)}$

19 **a** $\mathrm{CuSO_4(aq) + Na_2CO_3(aq) \rightarrow CuCO_3(s) + Na_2SO_4(aq)}$

$\mathrm{Cu^{2+}(aq) + CO_3^{2-}(aq) \rightarrow CuCO_3(s)}$

spectator ions: $\mathrm{Na^+}$, $\mathrm{SO_4^{2-}}$

b $\mathrm{AgNO_3(aq) + KCl(aq) \rightarrow AgCl(s) + KNO_3(aq)}$

$\mathrm{Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)}$

spectator ions: $\mathrm{K^+}$, $\mathrm{NO_3^-}$

c $\mathrm{Na_2S(aq) + Pb(NO_3)_2(aq) \rightarrow PbS(s) + 2NaNO_3(aq)}$

$\mathrm{Pb^{2+}(aq) + S^{2-}(aq) \rightarrow PbS(s)}$

spectator ions: $\mathrm{Na^+}$, $\mathrm{NO_3^-}$

d $\mathrm{FeCl_3(aq) + 3NaOH(aq) \rightarrow Fe(OH)_3(s) + 3NaCl(aq)}$

$\mathrm{Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)}$

spectator ions: $\mathrm{Na^+}$, $\mathrm{Cl^-}$

e $\mathrm{Fe_2(SO_4)_3(aq) + 6KOH(aq) \rightarrow 2Fe(OH)_3(s) + 3K_2SO_4(aq)}$

$\mathrm{Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)}$

spectator ions: $\mathrm{K^+}$, $\mathrm{SO_4^{2-}}$

20 **a** $\mathrm{HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)}$

b $\mathrm{H_2SO_4(aq) + K_2CO_3(aq) \rightarrow K_2SO_4(aq) + H_2O(l) + CO_2(g)}$

c $\mathrm{2H_3PO_4(aq) + 3Ca(HCO_3)_2(s) \rightarrow Ca_3(PO_4)_2(s) + 6H_2O(l) + 6CO_2(g)}$

d $\mathrm{2HF(aq) + Zn(OH)_2(s) \rightarrow ZnF_2(aq) + 2H_2O(l)}$

21 E. An acid reacts with a base to produce a salt plus water. In this case, the salt produced is magnesium chloride, which is soluble in water and exists as magnesium ions and chloride ions.

22 $\mathrm{2H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)}$

23 a metal carbonate or metal hydrogen carbonate

24 **a** $\mathrm{2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)}$

b $\mathrm{2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)}$

25 **a** decomposition

b synthesis

c acid base

d precipitation

e combustion

26 **a** combustion

b precipitation

c synthesis

d decomposition

e acid base

27 **a** sodium + chlorine → sodium chloride

b $\mathrm{2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)}$

c synthesis

28 **a** $HCl(aq) + KOH(aq) \rightarrow KCl(aq) + H_2O(l)$

b $2HCl(aq) + Ca(s) \rightarrow CaCl_2(aq) + H_2(g)$

c $2HNO_3(aq) + CaCO_3(s) \rightarrow Ca(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

d $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$

e $2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

29 acid and metal carbonate reaction

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Chapter 11 Predicting reactions of metals

11.1 Reactions of metals

11.1 KEY QUESTIONS

1 **a** sodium oxide
  b magnesium oxide
  c aluminium oxide
  d tin(II) oxide
2 **a** potassium, calcium, zinc, iron and silver
  b $4Na(s) + O_2(g) \rightarrow 2Na_2O(s)$
  $2Ca(s) + O_2(g) \rightarrow 2CaO$
  $2Zn(s) + O_2(g) \rightarrow 2ZnO$
  $4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$
  $Ag(s) + O_2(g) \rightarrow$ no reaction
3 **a** **i** K, Na or Ca
  ii Mg, Al, Zn or Fe
  iii Sn, Pb, Cu or Ag
  b **i** hydrogen gas, metal ions and hydroxide ions (or metal hydroxides)
  ii hydrogen gas and metal oxides
4 **a** $Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$
  b $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$
  c $Sn(s) + 2H^+(aq) \rightarrow Sn^{2+}(aq) + H_2(g)$
5 The metal displaces hydrogen ions (H^+) from the acid to form hydrogen gas (H_2).
6 **a** **i** A reaction occurs producing silver metal and zinc ions.
  ii no reaction
  b $2Ag^+(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$

11.2 The activity series of metals

Worked example: Try yourself 11.2.1

PREDICTING METAL DISPLACEMENT REACTIONS

Predict whether copper will displace gold from a solution containing gold ions and, if appropriate, write the ionic equation for the reaction.

Thinking	Working
Locate the metal and the metal ions in the activity series.	Metals are found on the right-hand side of the activity series. $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $Ag(s) \rightarrow Ag^+(aq) + e^-$ $Au(s) \rightarrow Au^+(aq) + e^-$
Determine whether the metal ion is below and to the right of the metal in the table. If this is the case, there will be a reaction.	You can see from the activity series that Au^+ is on the right-hand side and below Cu. This means that Au^+ is the ion of a less reactive metal, so there will be a reaction.
The more reactive metal atom will lose electrons to form a metal ion. Write a half-equation for this reaction.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

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The ion of the metal ion lower down the activity series will gain electrons, forming the metal atom. Write a half-equation for this reaction.	$Au^+(aq) + e^- \rightarrow Au(s)$
Combine the two half-equations, balancing electrons, to give the ionic equation for the reaction.	$(Au^+(aq) + e^- \rightarrow Au(s)) \times 2$ $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$ $2Au^+(aq) + Cu(s) \rightarrow 2Au(s) + Cu^{2+}(aq)$

11.2 KEY QUESTIONS

1 **a** magnesium > manganese > copper
  b magnesium > manganese > copper

2 $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$
  $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
  $Ni(s) + Cu^{2+}(aq) \rightarrow Ni^{2+}(aq) + Cu(s)$

3 For reactions to occur, the metal ion in the solution must be lower down the activity series than the metal.
  a no
  b no
  c yes
  d yes
  e yes
  f yes
  g no

4 **a** $3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$
  b no reaction
  c $2Ag^+(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$

5 Use iron and lead to test the solutions.
  The copper(II) nitrate will react with both iron and lead.
  The tin(II) nitrate will react only with the iron.
  The zinc nitrate will not react with either of the metals.

6 The activity series ranks the reactivity of metals in decreasing order from most reactive to least reactive. The table of standard reduction potentials ranks the reactivity of metals and non-metals in increasing order of standard reduction potential, from the lowest potential at the top of the table to the highest at the bottom.

  In the activity series, the reactivity of metals is determined in conditions that are not specified or controlled. In the table of standard reduction potentials, the reactivity of metals is measured under standard, specified conditions.

  The activity series of metals is based on observations of the relative vigour of the reactions of different metals. The table of standard reduction potentials ranks both metals and non-metals in order of reactivity on the basis of a numerical value, electrode potential, derived from precise measurements made under standardised conditions.

11.3 Metal activity and the periodic table

11.3 KEY QUESTIONS

1 In metals, reactivity increases down a group and reactivity decreases across the period.

2 Calcium has one more electron shell than magnesium, and its atomic radius is consequently greater. Because calcium has a greater atomic radius, the electrostatic force of attraction between its valence electrons and the nucleus is weaker. As a result, the ionisation energy and electronegativity of calcium are less than that of magnesium. A calcium atom is more reactive, since it will lose the outer-shell or valence electrons more readily than a magnesium atom does.

3 The potassium atom has one more electron shell than the sodium atom. Because there is a greater distance between the valence electron and the nucleus, the electrostatic attraction is weaker, so it is easier for the potassium atom to lose its valence electron.

The potassium atom has a core charge of +1, whereas the calcium atom has a core charge of +2. The greater core charge of the calcium atom means that the attraction between its valence electrons and nucleus is stronger than the attraction between the valence electron and nucleus of the potassium atom.

4 Metal II is the most reactive. The most reactive metal will most readily lose its valence electron. Metal II has the greatest atomic radius. The force of attraction between the nucleus and valence electrons will be less, since the distance between the nucleus and the outer electron shell is the greatest. Metal II also has the lowest ionisation energy, indicating that it requires less energy to lose an electron. Electronegativity is a measure of the ability of an atom to attract electrons. Metal II has the lowest electronegativity, so it has the least ability to attract electrons.

CHAPTER 11 **REVIEW**

1 **a** calcium oxide
b hydrogen and hydroxide ions (or potassium hydroxide solution)
c hydrogen and $Zn^{2+}(aq)$ ions

2 B. In this reaction, the half-equation for sodium is $Na(s) \rightarrow Na^+(s) + e^-$. Each sodium atom loses one electron.

3 B

4 Cu

5 D

6 Add the mixture of metals to some dilute acid. The zinc metal will dissolve in the dilute acid, whereas copper metal will not dissolve.

$Zn(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + H_2(g)$

The copper can be recovered from the solution by filtration.

7 C. The equation is not balanced correctly as it is written. Because the silver ion has a +1 charge, Ag^+ ions and Cu atoms will react in a 2:1 ratio in the reaction. The equation is balanced when it is written as:

$2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

8 **a** $Ag^+(aq) + e^- \rightarrow Ag(s)$
b $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
c $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

9 Na, Li, Ca , Mg, Al, Fe, Sn, Cu, Ag

10 The activity series of metals is based on experimental observations of the relative rate of reaction of metals with other substances such as oxygen, water and dilute acids. The rates of reaction are described using terms such as vigorous fast, slow etc. Qualitative data is based on observations, whereas quantitative data is based on measurements.

11 **a** $Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$
b no reaction
c no reaction
d $Mg(s) + Pb^{2+}(aq) \rightarrow Mg^{2+}(aq) + Pb(s)$
e no reaction
f no reaction
g no reaction
h $Pb(s) + 2Ag^{2+}(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$

12 **a** no
b $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$
c $Zn(s) + SnCl_2(aq) \rightarrow ZnCl_2(aq) + Sn(s)$
d $Zn(s) + CuSO_4(aq) \rightarrow ZnSO_4(aq) + Cu(s)$

13 Place 10 mL of each solution into a clean test-tube. Add a small piece of magnesium ribbon to each solution. The magnesium should be coated by displaced metal in the silver nitrate and lead(II) nitrate solutions, but not in the sodium nitrate solution. The lead coating will be black. Over time, the silver coating will change from black to silver as more metal is deposited. To confirm the identity of the silver nitrate solution, take a fresh sample and add a small piece of copper. Copper will displace silver from the solution, giving a silver deposit and a blue solution. Copper will not displace lead from the solution.

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14 Zinc. Zinc is more reactive than iron, but less reactive than aluminium, so will only displace iron from solution. Cobalt is less reactive than both iron and aluminium so will displace neither from solution.

15 Coatings of metals other than iron would be expected on the nails placed in 1 mol L^{-1} solutions of $CuSO_4$ and $Pb(NO_3)_2$. Iron is more reactive than copper and lead, so iron will displace copper(II) and lead(II) ions from solution.

16 **a** Magnesium and sodium, with outer electrons in the third shell, are in the same period. Magnesium has a nuclear charge of +12, but with completed inner shells of $1s^22s^22p^6$, the outer electrons experience the attraction of a core charge of +2. The outer-shell electrons of sodium, which has a nuclear charge of +11 and the same number of inner shells as magnesium, are attracted by a core charge of +1. The stronger attraction of the magnesium electrons to the core means that more energy is required to remove an electron from a magnesium atom than from a sodium atom.

b Both lithium and sodium are in group 1, so the outer electron of each atom experiences the attraction of the same core charge. Because the outer-shell electrons of a lithium atom are closer to the nucleus than those of a sodium atom, they are attracted more strongly and more energy is needed to remove one.

17 **a** aluminium

b The aluminium atom has a core charge of +3 and a magnesium atom has a core charge of +2. Aluminium is the least reactive because the valence electrons in aluminium are more strongly attracted to the nucleus than are the valence electrons in magnesium. Sulfur, chlorine, fluorine and oxygen are non-metals.

18 **a** aluminium. The elements are all members of group 13 and have a core charge of +3. Aluminium has fewer electron shells than the other elements in the list. Since the valence electrons in aluminium are closer to the nucleus, the force of attraction between the valence electrons and the nucleus is stronger. Consequently, more energy is required to remove the first electron from the valence shell of an aluminium atom than from the valence shell of a gallium or indium atom.

b aluminium. The aluminium atom has a core charge of +3, magnesium atoms have a core charge of +2 and a sodium atom has a core charge of +1. The valence electrons in aluminium are more strongly attracted to the nucleus than the valence electrons in magnesium or sodium. Consequently, more energy is required to remove the first electron from the valence shell of an aluminium atom than from the valence shell of a magnesium or sodium atom.

19 **a** Mg(s)

b $Ag^+(aq)$

c A coating of silver will form on the lead when it is placed in silver nitrate solution.

d zinc and magnesium

20 Experiment 1 gives you evidence that T is a more reactive metal than R.

Experiment 2 suggests that T is a more reactive metal than S, because S does not displace T from a solution of its salt, $T(NO_3)_2$.

Experiment 3 indicates that R is a more reactive metal than S, because S does not displace R from a solution of its salt, RNO_3.

The information above tells you that S is less reactive than both T and R and that R is less reactive than T, so the correct order of decreasing reactivity is T > R > S.

21 List A contains metals, which tend to lose electrons to non-metals when they react. List B contains non-metals, which tend to gain electrons when they react.

22 Between Na and Ca: K, Na, X, Ca, Mg, Al, Zn, Cu, Ag.

The electronic structure shows that there are two electrons in the valence shell and that all the inner electron shells are filled, indicating that X is a group 2 metal. As the valence electrons are in the 5th shell, X is in the 5th period. This places X below calcium in group 2. The reactivity of group 2 metals in water increases down the group, so you would expect X to be more reactive than Ca. Metal X is placed after Na, since the reaction of group 1 metals in water is more vigorous than the reaction of group 2 metals in water.

23 More reactive metals have a larger atomic radius, lower ionisation energies and lower electronegativities than less reactive metals. The order of reactivity of the metals is II > IV > III > I.

24 The activity series of metals lists metals in decreasing order of reducing strength. Zinc is placed higher on the activity series of metals than iron, and therefore zinc metal is a stronger reducing agent than iron metal.

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Chapter 12 Redox reactions and galvanic cells

12.1 Introducing redox reactions

Worked example: Try yourself 12.1.1

RECOGNISING OXIDATION AND REDUCTION

Write the oxidation and reduction half-equations for the reaction with the overall equation:

$$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$$

Thinking	Working
Identify the ions in the product.	NaCl is made up of Na^+ and Cl^- ions.
Write the half-equation for the oxidation of the reactant that forms positive ions and balance the equation with electrons.	$Na(s) \rightarrow Na^+(s) + e^-$
Write the half-equation for the reduction of the reactant that forms negative ions and balance the equation with electrons.	$Cl_2(g) + 2e^- \rightarrow 2Cl^-(s)$

Worked example: Try yourself 12.1.2

WRITING SIMPLE HALF-EQUATIONS

When a piece of copper metal is placed into a silver nitrate solution, silver metal is formed and the solution gradually turns blue, indicating the presence of copper(II) ions in solution. The oxidation and reduction reactions can be represented by two half-equations.

Write these half-equations and identify the substances that are oxidised and reduced.

Thinking	Working
Identify one reactant and the product it forms and write them on each side of an equation. Balance the equation for the element.	$Ag^+(aq) \rightarrow Ag(s)$
Add electrons to balance the equation for charge.	$Ag^+(aq) + e^- \rightarrow Ag(s)$
To decide whether the reactant is oxidised or reduced, remember that oxidation is loss of electrons and reduction is gain of electrons.	Electrons are gained, so this is reduction. The $Ag^+(aq)$ is being reduced.
Identify the second reactant and the product it forms, and write them on each side of an equation. Balance the equation for the element.	$Cu(s) \rightarrow Cu^{2+}(aq)$
Add electrons to balance the equation for charge.	$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
To decide whether the reactant is oxidised or reduced, remember that oxidation is loss of electrons and reduction is gain of electrons.	Electrons are lost, so this is oxidation. The Cu(s) is being oxidised.

Worked example: Try yourself 12.1.3

WRITING OVERALL REDOX EQUATIONS FROM HALF-EQUATIONS

Potassium metal is oxidised by oxygen gas in air to form solid potassium oxide.

Write the half-equations for the reaction, and hence write the balanced overall equation.

Thinking	Working
Identify one reactant and the product it forms, and write them on either side of the equation. Balance the equation for the element.	$O_2(g) \rightarrow 2O^{2-}(s)$

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Add electrons to balance the equation for charge.	$O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$
Identify the second reactant and the product it forms, and write them on either side of the equation. Balance the equation for the element.	$K(s) \rightarrow K^+(s)$
Add electrons to balance the equation for charge.	$K(s) \rightarrow K^+(s) + e^-$
Multiply one equation by a suitable factor to ensure that the number of electrons is balanced.	$(K(s) \rightarrow K^+(s) + e^-) \times 4$ $4K(s) \rightarrow 4K^+(s) + 4e^-$
Add the oxidation and the reduction half-equations together, cancelling out electrons so that none are in the final equation. Combine ions to create the formula of the product.	$O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$ $4K(s) \rightarrow 4K^+(s) + 4e^-$ When the electrons have been cancelled, the overall equation is: $4K(s) + O_2(g) \rightarrow 2K_2O(s)$

Worked example: Try yourself 12.1.4

BALANCING COMPLEX HALF-EQUATIONS

Write the half-equation for the reduction of an acidified solution of MnO_4^- to solid MnO_2.

Thinking	**Working**
Balance all elements except hydrogen and oxygen in the half-equation.	There is 1 Mn in MnO_4^- and 1 Mn in MnO_2, so the Mn atoms are balanced. $MnO_4^- \rightarrow MnO_2$
Balance the oxygen atoms by adding water.	There are 4 O atoms in MnO_4^- and only 2 O atoms in MnO_2, so 2 H_2O must be added to the RHS. $MnO_4^- \rightarrow MnO_2 + 2H_2O$
Balance the hydrogen atoms by adding H^+ ions. Acids provide a source of H^+ ions.	There are now 4 H atoms on the RHS and none on the LHS, so $4H^+$ are added to the LHS. $MnO_4^- + 4H^+ \rightarrow MnO_2 + 2H_2O$
Balance the charge in the equation by adding electrons.	The charge on the LHS is $(-1) + (4 \times +1) = +3$ and on the RHS is 0, so 3 electrons are added to the LHS to make the charges equal. $MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$
Add states to complete the half-equation.	$MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$

Worked example: Try yourself 12.1.5

COMBINING COMPLEX HALF-EQUATIONS

Write balanced oxidation and reduction half-equations for the reaction in which $SO_3^{2-}(aq)$ and $ClO^-(aq)$ react to form $H_2S(g)$ and $ClO_3^-(aq)$. Then write the overall equation for the reaction.

Thinking	**Working**
Identify one reactant and the product it forms, and write the balanced half-equation.	$SO_3^{2-}(aq) + 8H^+(aq) + 6e^- \rightarrow H_2S(g) + 3H_2O(l)$
Identify the second reactant and the product it forms, and write the balanced half-equation.	$ClO^-(aq) + 2H_2O(l) \rightarrow ClO_3^-(aq) + 4H^+(aq) + 4e^-$
Multiply one or both equation(s) by a suitable factor to ensure that the number of electrons on both sides of the arrow is equal.	Lowest common multiple = 12 $2 \times [SO_3^{2-}(aq) + 8H^+(aq) + 6e^- \rightarrow H_2S(g) + 3H_2O(l)]$ $3 \times [ClO^-(aq) + 2H_2O(l) \rightarrow ClO_3^-(aq) + 4H^+(aq) + 4e^-]$ This gives: $2SO_3^{2-}(aq) + 16H^+(aq) + 12e^- \rightarrow 2H_2S(g) + 6H_2O(l)$ $3ClO^-(aq) + 6H_2O(l) \rightarrow 3ClO_3^-(aq) + 12H^+(aq) + 12e^-$

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Add the oxidation and the reduction half-equations together, cancelling electrons so that none are in the final equation. Also cancel H_2O and H^+ if these occur on both sides of the arrow.	$2SO_3^{2-}(aq) + \overset{4H^+}{\cancel{16H^+}}(aq) + 12e^- \rightarrow 2H_2S(g) + \cancel{6H_2O(l)}$ $3ClO^-(aq) + \cancel{6H_2O(l)} \rightarrow 3ClO_3^-(aq) + \cancel{12H^+(aq)} + 12e^-$ $2SO_3^{2-}(aq) + 4H^+(aq) + 3ClO^-(aq) \rightarrow 2H_2S(g) + 3ClO_3^-(aq)$

12.1 KEY QUESTIONS

1 **a** oxidation
b reduction
c reduction
d oxidation

2 **a** $Fe(s) \rightarrow Fe^{3+}(aq) + 3e^-$ (oxidation)
b $K(s) \rightarrow K^+(aq) + e^-$ (oxidation)
c $F_2(g) + 2e^- \rightarrow 2F^-(aq)$ (reduction)
d $O_2(g) + 4e^- \rightarrow 2O^{2-}(aq)$ (reduction)

3 **a** Fe(s) has been oxidised to $Fe^{2+}(aq)$.
b $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-$
c $H^+(aq)$
d $H^+(aq)$ has been reduced to $H_2(g)$.
e $2H^+(aq) + 2e^- \rightarrow H_2(g)$
f Fe(s)
g $Fe^{2+}(aq)/Fe(s)$ and $H^+(aq)/H_2(g)$

4 **a** Magnesium is oxidised, and copper ions are reduced.
b $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$
c $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$
d $Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$
e oxidizing agent Cu^{2+}; reducing agent Mg
f Copper(II) ions cause a solution to appear blue. The solution loses some of its blue colour due to the loss of $Cu^{2+}(aq)$, which is reduced to form Cu(s).

5 **a** $MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(l)$
b $MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$
c $SO_4^{2-}(aq) + 10H^+(aq) + 8e^- \rightarrow H_2S(g) + 4H_2O(l)$
d $SO_2(g) + 2H_2O(l) \rightarrow SO_4^{2-}(aq) + 4H^+(aq) + 2e^-$
e $H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-$
f $SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$

6 **a** CaO
b Ca(s)
c $Ca(s) \rightarrow Ca^{2+}(s) + 2e^-$
d $O_2(g)$
e $O_2(g) + 4e^- \rightarrow 2O^{2-}(s)$
f $2Ca(s) + O_2(g) \rightarrow 2CaO(s)$
g Calcium has been *oxidised* by *oxygen* to calcium ions. The *calcium* has lost electrons to the *oxygen*.
The oxygen has been *reduced* by *calcium* to oxide ions. The *oxygen* has gained electrons from the *calcium*.

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12.2 Oxidation numbers

Worked example: Try yourself 12.2.1

CALCULATING OXIDATION NUMBERS

Use the rules in Table 12.2.1 to determine the oxidation number of each element in $NaNO_3$.

Thinking	**Working**
Identify an element that has a set value.	Na is a main group metal in group 1. Applying rule 3a, the oxidation number of Na is +1.
Identify any other elements that have set values.	According to rule 3c, oxygen has an oxidation number of –2 unless attached to fluorine or in a peroxide.
Use algebra to work out the oxidation number of other elements.	Let the oxidation number of nitrogen in $NaNO_3$ be x. Solve the sum of the oxidation numbers for x: $+1 + x + (3 \times -2) = 0$ $+1 + x - 6 = 0$ $x - 5 = 0$ $x = +5$
Write oxidation numbers above the elements in the formula.	$\overset{+1}{Na}\overset{+5}{N}\overset{-2}{O}_3$ Note that the oxidation number of oxygen is written as –2 (not as –6), even though there are three oxygen atoms in the formula.

Worked example: Try yourself 12.2.2

USING OXIDATION NUMBERS TO IDENTIFY OXIDATION AND REDUCTION IN AN EQUATION

Use oxidation numbers to determine which element has been oxidised and which has been reduced in the following equation:

$$CuO(s) + H_2(g) \rightarrow Cu(s) + H_2O(l)$$

Thinking	**Working**
Determine the oxidation numbers of one of the elements on each side of the equation.	Choose Cu as the first element. As Cu can have a variable oxidation state, let the oxidation number = x $x + -2 = 0$ $x = +2$ $\overset{+2}{Cu}O(s) + H_2(g) \rightarrow \overset{0}{Cu}(s) + H_2O(l)$
Assess whether the oxidation number has changed. If so, identify whether it has increased (oxidation) or decreased (reduction).	The oxidation number of copper has decreased from +2 to 0, so the copper in CuO has been reduced.
Determine the oxidation numbers of a second element on the left-hand and the right-hand side of the equation.	Choose hydrogen as the second element. $CuO(s) + \overset{0}{H_2}(g) \rightarrow Cu(s) + \overset{+1}{H_2}O(l)$
Assess whether the oxidation number has changed. If so, identify whether it has increased (oxidation) or decreased (reduction).	The oxidation number of H has increased from 0 to +1, so H_2 has been oxidised.
Continue this process until the oxidation numbers of all elements have been determined.	Determine the oxidation numbers of oxygen. $Cu\overset{-2}{O}(s) + H_2(g) \rightarrow Cu(s) + H_2\overset{-2}{O}(l)$ The oxidation number of O has not changed.

12.2 KEY QUESTIONS

1 **a** +2
  b +4
  c –4
  d 0
  e +4

2 $KMnO_4$: the oxidation state of K is +1 O is –2
$(2 \times +1) + x + (4 \times -2) = 0$; hence, $x = +6$

3 **a** Ca +2 O –2
  b Ca +2 C –1
  c H +1 S +6 O –2
  d Mn +7 O –2
  e F 0
  f S +4 O –2
  g Na +1; N +5 O –2
  h K +1 Cr +6 O –2

4 **a** $\overset{0}{Mg}(s) + \overset{0}{Cl_2} \rightarrow \overset{+2}{Mg}\overset{-1}{Cl_2}(s)$ oxidant Cl_2 reductant Mg

  b $2\overset{+4-2}{SO_2}(g) + \overset{0}{O_2}(g) \rightarrow 2\overset{+6-2}{SO_3}(g)$ oxidant O_2 reductant SO_2

  c $\overset{+3\ -2}{Fe_2O_3}(s) + 3\overset{+2-2}{CO}(g) \rightarrow 2\overset{0}{Fe}(s) + 3\overset{+4-2}{CO_2}(g)$ oxidant Fe_2O_3 reductant CO

  d $2\overset{+2}{Fe}^{2+}(aq) + \overset{+1\ -1}{H_2O_2}(aq) + 2\overset{+1}{H}^{+}(aq) \rightarrow 2\overset{+3}{Fe}^{3+}(aq) + 2\overset{+1\ -2}{H_2O}(l)$ oxidant H_2O_2; reductant Fe^{2+}

5

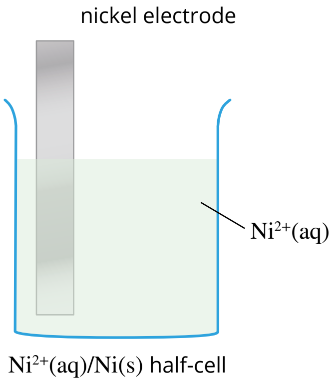
Redox reaction	**Conjugate redox pair (oxidation process)**	**Conjugate redox pair (reduction process)**
$Na(s) + Ag^+(aq) \rightarrow Na^+(aq) + Ag(s)$	$Na^+(aq)/Na(s)$	$Ag^+(aq)/Ag(s)$
$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$	$Zn^{2+}(aq)/Zn(s)$	$Cu^{2+}(aq)/Cu(s)$
$2K(s) + Cl_2(g) \rightarrow 2K^+(s) + 2Cl^-(s)$	$K^+(s)/K(s)$	$Cl_2(g)/Cl^-(s)$

12.3 Galvanic cells

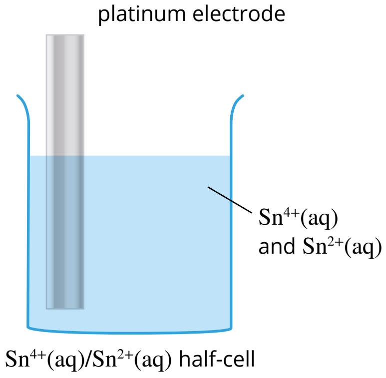
12.3 KEY QUESTIONS

1 D.
A is incorrect because in galvanic cells, the electrode in the half-cell in which oxidation occurs is called the anode.
B is incorrect because electrons flow from the anode to the cathode. The cathode is where reduction takes place.
C is incorrect because anions flow into the half-cell where oxidation occurs (the anode).
D is the only correct option—the electrode where oxidation occur i.e. the anode) has a negative polarity.

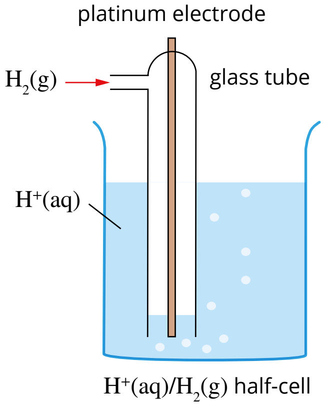
2 **a**



b

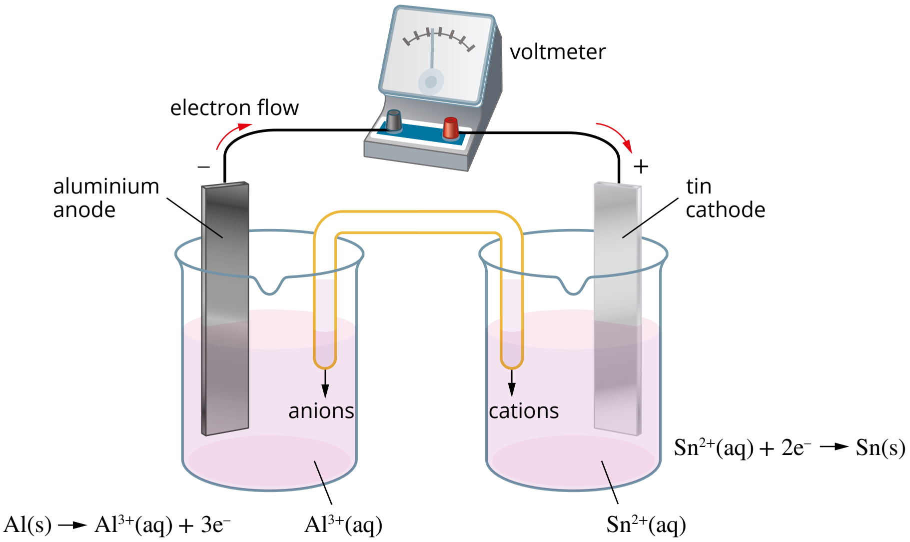


c



3 **a** Reduction occurs at the cathode: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$
  b Oxidation occurs at the anode: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

4



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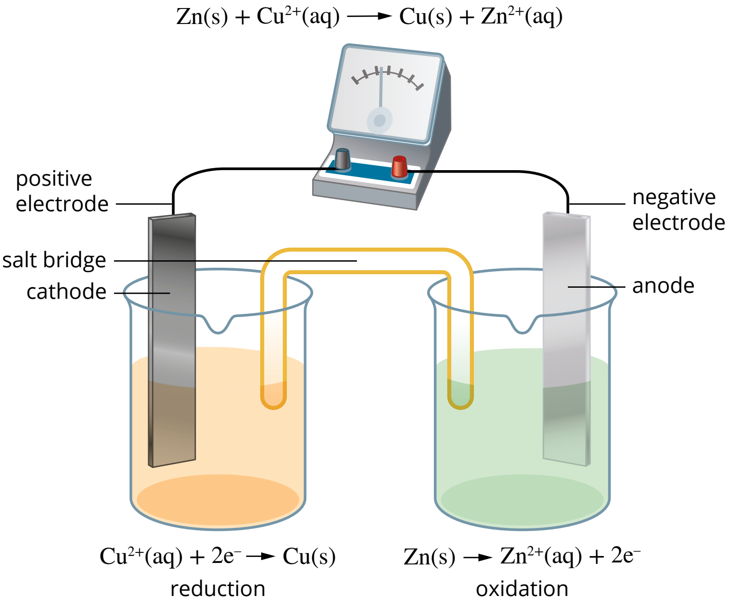

12.4 The table of standard reduction potentials

Worked example: Try yourself 12.4.1

PREDICTING THE OPERATION OF A GALVANIC CELL

A cell is made from $Sn^{2+}(aq)/Sn(s)$ and $Ni^{2+}(aq)/Ni(s)$ half-cells under standard conditions and at 25°C. Use the table of standard reduction potentials to predict the overall cell reaction, identify the anode and cathode, and determine the direction of electron flow.

Thinking	Working
Identify the two relevant half-equations in the table of standard reduction potentials.	$Ni^{2+}(aq) + 2e^- \rightleftharpoons Ni(s) \quad E^\circ = -0.24\,V$ $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s) \quad E^\circ = -0.14\,V$
Identify the strongest oxidising agent (the species on the left side of the table with the most positive E° value) and the strongest reducing agent (top right).	Because Sn^{2+} is lower on the left side of the table than Ni^{2+}, it is the stronger oxidising agent. Ni, being higher on the right side of the table than Sn, is the stronger reducing agent.
Write the two half-equations that will occur. The strongest oxidising agent will react with the strongest reducing agent. (Hint: The reduction equation has the most positive E° value, and the oxidation equation has the most negative E° value.)	Oxidation: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$ Reduction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

Write the overall cell equation.	The number of electrons in each half-equation is equal, so add the two equations together: $Sn^{2+}(aq) + 2e^- + Ni(s) \rightarrow Sn(s) + Ni^{2+}(aq) + 2e^-$ Cancel out the $2e^-$ to give: $Sn^{2+}(aq) + Ni(s) \rightarrow Sn(s) + Ni^{2+}(aq)$
Identify the anode and the cathode in this cell. The anode is the electrode at which oxidation occurs. The cathode is the electrode at which reduction occurs.	The nickel electrode will be the anode and the tin electrode will be the cathode.
Determine the polarities of the electrodes and the direction of electron flow in the cell. The anode is negative; the cathode is positive.	Electrons flow from the negative electrode (anode) to the positive electrode (cathode).

Worked example: Try yourself 12.4.2

PREDICTING DIRECT REDOX REACTIONS

Consider the following equations that appear in the order shown in the table of standard reduction potentials.

$Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s) \quad E^\circ = -1.18\,V$

$Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s) \quad E^\circ = -0.14\,V$

$Cu^{+}(aq) + e^- \rightleftharpoons Cu(s) \quad E^\circ = +0.52\,V$

Use the table of standard reduction potentials to predict the effect of mixing:

a $Mn(s)$ and $Cu^{+}(aq)$

b $Mn^{2+}(aq)$ and $Sn^{2+}(s)$

c $Cu(s)$ and $Sn^{2+}(s)$.

a $Mn(s)$ and $Cu^{+}(aq)$	
Thinking	**Working**
Identify the two relevant half-equations in the table of standard reduction potentials.	**a** $Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s) \quad E^\circ = -1.18\,V$ $Cu^{+}(aq) + e^- \rightleftharpoons Cu(s) \quad E^\circ = +0.52\,V$
Predict whether a reaction will occur. A chemical species on the left-hand side of the table of standard reduction potentials (an oxidising agent) reacts with a chemical species on the right-hand side that is higher in the series (a reducing agent).	A reaction occurs because the oxidising agent, Cu^+, is below the reducing agent, Mn, in the table of standard reduction potentials. The lower half-equation occurs in the forward direction: $Cu^{+}(aq) + e^- \rightarrow Cu(s)$ The higher half-equation occurs in the reverse direction: $Mn(s) \rightarrow Mn^{2+}(aq) + 2e^-$
Write the overall equation.	The overall equation is found by adding the half-equations: $2Cu^{+}(aq) + Mn(s) \rightarrow 2Cu(s) + Mn^{2+}(aq)$
b $Mn^{2+}(aq)$ and $Sn^{2+}(s)$	
Identify the two relevant half-equations in the table of standard reduction potentials. Predict whether or not a reaction occurs. A chemical species on the left-hand side of the table of standard reduction potentials (an oxidising agent) reacts with a chemical species on the right-hand side that is higher in the series (a reducing agent). Write the overall equation.	**b** $Mn^{2+}(aq) + 2e^- \rightleftharpoons Mn(s) \quad E^\circ = -1.18\,V$ $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s) \quad E^\circ = -0.14\,V$ No reaction occurs because the Mn^{2+} and Sn^{2+} are both oxidising agents.

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c Cu(s) and Sn^{2+}(s).	
Identify the two relevant half-equations in the table of standard reduction potentials. Predict whether or not a reaction occurs. A chemical species on the left-hand side of the table of standard reduction potentials (an oxidising agent) reacts with a chemical species on the right-hand side that is higher in the series (a reducing agent). Write the overall equation.	**c** $Sn^{2+}(aq) + 2e^- \rightleftharpoons Sn(s) \quad E° = -0.14\,V$ $Cu^{+}(aq) + e^- \rightleftharpoons Cu(s) \quad E° = +0.52\,V$ No reaction occurs because the oxidising agent, Sn^{2+}, is above the reducing agent, Cu, in the table of standard reduction potentials.

12.4 KEY QUESTIONS

1 **a** Because Pb^{2+} is lower on the left side of the table than Al^{3+}, it is the stronger oxidising agent and gets reduced. Al, being higher on the right side of the table than Pb, is a stronger reducing agent and gets oxidised.

reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

b Multiply the Pb^{2+}/Pb half-equation by 3 and the Al^{3+}/Al half-equation by 2 so that the number of electrons in each half-equation is equal, and then add the two equations together:

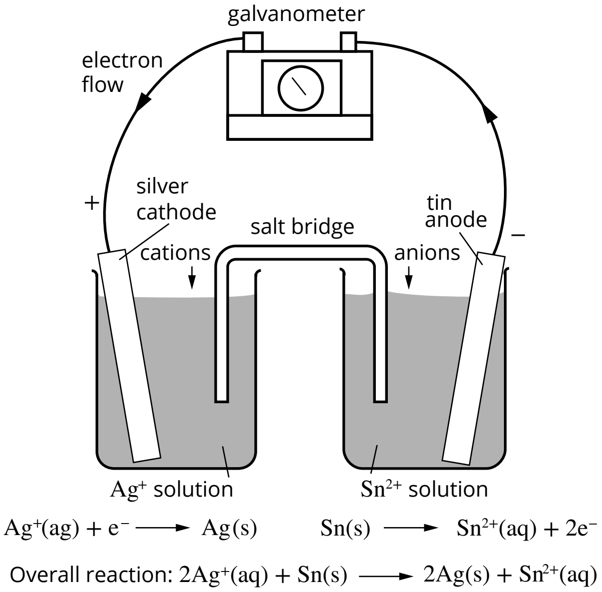
$[Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)] \times 3$

$[Al(s) \rightarrow Al^{3+}(aq) + 3e^-] \times 2$

$3Pb^{2+}(aq) + 2Al(s) \rightarrow 3Pb(s) + 2Al^{3+}(aq)$

c The lead electrode will be the cathode and the aluminium electrode will be the anode.

2



3 Cell potential difference = most positive half-cell $E°$ – most negative half-cell $E°$

$= E°(Ag^+(aq)/Ag(s)) - E°(Sn^{2+}(aq)/Sn(s))$

$= 0.80 - (-0.14)$

$= 0.94\,V$

4 **a** A reaction occurs because the oxidising agent, Cu^{2+}, is below the reducing agent, Zn, in the table of standard reduction potentials.

$$Cu^{2+}(aq) + Zn(s) \rightarrow Zn^{2+}(aq) + Cu(s)$$

b No reaction, because the oxidising agent, Ca^{2+}, is above the reducing agent, Mg, in the table of standard reduction potentials.

c No reaction, because Pb^{2+} and Fe^{2+} are both oxidising agents and cannot react with each other.

5 For a reaction to occur, the reducing agent, Fe, must be above the oxidising agent in the table of standard reduction potentials.

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$CuSO_4$: A reaction occurs because the reducing agent, Fe, is above the oxidising agent, Cu^{2+}. Cu^{2+} is reduced, forming a coating of copper on the nail.

$MgCl_2$: No reaction occurs because the reducing agent, Fe, is below the oxidising agent, Mg^{2+}. No reaction occurs between Fe and Cl^- because they are both reducing agents.

$Pb(NO_3)_2$: A reaction occurs because the reducing agent, Fe, is above the oxidising agent, Pb^{2+}. Pb^{2+} is reduced, forming a coating of lead on the nail.

$ZnCl_2$: No reaction occurs because the reducing agent, Fe, is below the oxidising agent, Zn^{2+}. No reaction occurs between Fe and Cl^- because they are both reducing agents.

CHAPTER 12 **REVIEW**

1 D. The first half-equation shows that sodium metal, Na(s), is losing electrons, so is being oxidised. It must therefore be the reducing agent. The second half-equation shows that $Ag^+(aq)$ is gaining electrons, so it is being reduced. The $Ag^+(aq)$ must therefore be the oxidising agent.

2 A.

3 B. In this reaction, the half-equation for magnesium is $Mg(s) \rightarrow Mg^{2+}(s) + 2e^-$. Each magnesium atom loses two electrons.

4 Oxidation and reduction occur together. Oxidation occurs when an atom *loses* electrons to form a *positive* ion, such as happens when a calcium atom, with an electronic configuration of 2,8,8,2 *loses* electrons to form a Ca^{2+} ion. Reduction occurs when an atom *gains* electrons to form a *negative* ion, or a cation *gains* electrons to become less positively charged. An example is when a bromine atom, with *7* electrons in its valence shell, *gains* an electron to form a *bromide* ion.

5 oxidation half-equation: $Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-$

reduction half-equation: $Ag^+(aq) + e^- \rightarrow Ag(s)$

overall: $Pb(s) + 2Ag^+(aq) \rightarrow Pb^{2+}(aq) + 2Ag(s)$

6
- **a** oxidation
- **b** oxidation
- **c** reduction
- **d** oxidation
- **e** reduction
- **f** reduction

7
- **a** Charges are not balanced. $Ag(s) \rightarrow Ag^+(aq) + e^-$
- **b** Electrons appear on both sides of the equation. $Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$
- **c** State symbols are incorrect. $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$
- **d** Atoms are not balanced. $I_2(aq) + 2e^- \rightarrow 2I^-(aq)$
- **e** Electrons are being subtracted. $Na^+(aq) + e^- \rightarrow Na(s)$

8

Step	**Task**	**How it's done**	**Half-equation**
1	Balance nitrogens.	already balanced	$NO_3^- \rightarrow NO_2$
2	Balance oxygens by adding *H_2O*.	Add *one* H_2O molecule(s) to the right-hand side of the equation.	$NO_3^- \rightarrow NO_2 + H_2O$
3	Balance hydrogens by adding *H^+*.	Add $2H^+$ ion(s) to the *left-hand* side of the equation.	$NO_3^- + 2H^+ \rightarrow NO_2 + H_2O$
4	Balance charge by adding *electrons*.	Charge on left-hand side $= -1 + 2$ $= +1$ Charge on right-hand side = *0* Add *one* e^- to the *left-hand side* of the equation.	$NO_3^- + 2H^+ + e^- \rightarrow NO_2 + H_2O$
5	Add state symbols to give the final half-equation.	Give the appropriate states for each reactant and product in the equation.	$NO_3^-(aq) + 2H^+(aq) + e^- \rightarrow NO_2(g) + H_2O(l)$

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9 $2IO_3^-(aq) + 12H^+(aq) + 10e^- \rightarrow I_2(aq) + 6H_2O(l)$

10 half-equations: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

$2MnO_2(s) + 2H^+(aq) + 2e^- \rightarrow Mn_2O_3(s) + H_2O(l)$

overall equation: $Zn(s) + 2MnO_2(s) + 2H^+(aq) \rightarrow Zn^{2+}(aq) + Mn_2O_3(s) + H_2O(l)$

11 **a** $2H_3AsO_4(aq) + 4H^+(aq) + 4e^- \rightarrow As_2O_3(s) + 5H_2O(l)$

$I^-(aq) + 3H_2O(l) \rightarrow IO_3^-(aq) + 6H^+(aq) + 6e^-$

b Common multiple for electrons = 12

Reduction equation needs to be multiplied by 3:

$6H_3AsO_4(aq) + 12H^+(aq) + 12e^- \rightarrow 3As_2O_3(s) + 15H_2O(l)$

Oxidation equation needs to be multiplied by 2:

$2I^-(aq) + 6H_2O(l) \rightarrow 2IO_3^-(aq) + 12H^+(aq) + 12e^-$

Cancel e^-, H_2O and H^+:

$6H_3AsO_4(aq) + 2I^-(aq) \rightarrow 3As_2O_3(s) + 2IO_3^-(aq) + 9H_2O(l)$

12 **a** +4
b –2
c +6
d +6
e +4
f +2

13

Compound	Element	Oxidation number
$CaCO_3$	Ca	+2
HNO_3	O	–2
H_2O_2	O	–1
HCO_3^-	C	+4
HNO_3	N	+5
$KMnO_4$	Mn	+7
H_2S	S	–2
Cr_2O_3	Cr	+3
N_2O_4	N	+4

14 K_3N, N_2, N_2O, NO, N_2O_3, N_2O_4, $Ca(NO_3)_2$

Oxidation numbers for nitrogen are as follows:

K_3N: –3
N_2: 0
N_2O: +1
NO: +2
N_2O_3: +3
N_2O_4: +4
$Ca(NO_3)_2$: +5

15 **b**, **c**, **e**, **f** and **h** are redox reactions because the elements in the reactions undergo changes in oxidation number during the course of the reaction.

The changes in oxidation number which occur are:

b Ag from 0 to +1; Cl from 0 to –1
c Fe from +3 to +2; Sn from +2 to +4
e P from +3 to +5; I from 0 to –1
f Cu from +1 to +2; Cu from +1 to 0
h P from 0 to –3; H from 0 to +1

16 D If no salt bridge were present, the solution in one half-cell would accumulate a negative charge and the solution in the other half-cell would accumulate a positive charge as the reaction proceeds. Such accumulations of charge would very quickly prevent further reaction, and hence prevent production of electricity.

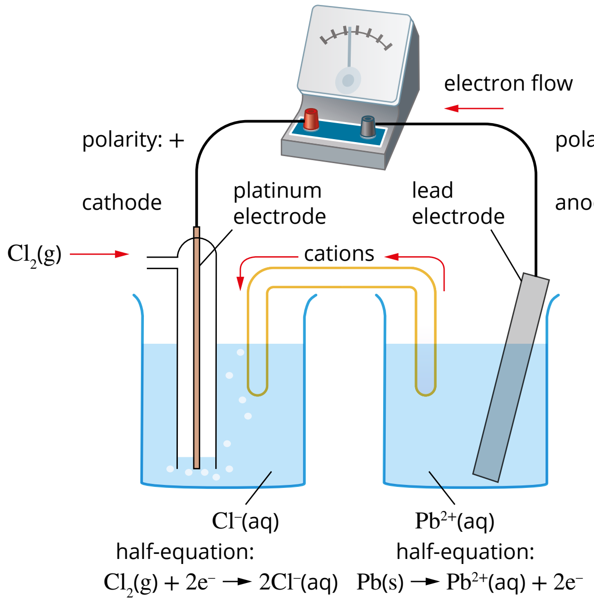
17 A If one of the conjugate redox pairs in a half-cell is not a metal, an inert (unreactive) electrode such as platinum or graphite is used. Compared with gold, iron is the least suitable for use as an electrode because it is more likely to react and become involved in a cell reaction.

18 **a** Oxidising agent a substance that causes another substance to be oxidised and is reduced in the process. Reducing agent a substance that causes another substance to be reduced and is oxidised in the process.

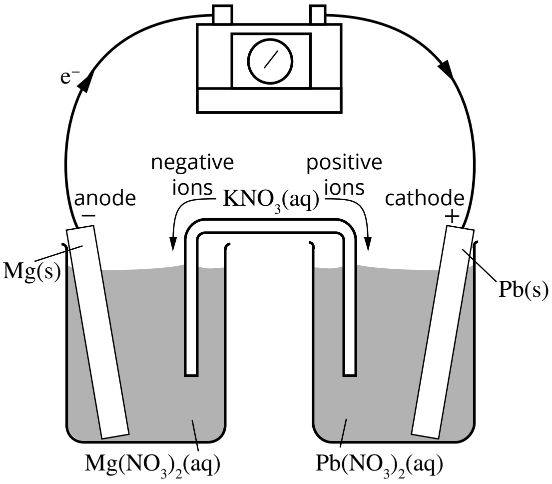
b Anode electrode at which oxidation occurs. Cathode: electrode at which reduction occurs.

c External circuit: section of a circuit where the electrons flow e.g. through wires. Internal circuit: part of cell where the current is due to the movement of ions, eg. in the salt bridge.

19



20 **a**



b

$$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$$

$$Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$$

Overall: $Mg(s) + Pb^{2+}(aq) \rightarrow Mg^{2+}(aq) + Pb(s)$

c The lead electrode is the cathode; the magnesium electrode is the anode.

d Anions will migrate to the $Mg^{2+}(aq)/Mg(s)$ half-cell; cations to the $Pb^{2+}(aq)/Pb(s)$ half-cell.

21 A reducing agent is oxidised. Oxidation occurs at the anode, which is the negative electrode.

In the first cell, the reducing agents are Sn^{2+} and Fe^{2+}. Because Sn^{2+} is contained in the half-cell with the anode, Sn^{2+} is the stronger reducing agent.

In the second cell, the reducing agents are Fe^{2+} and Br^-. Because Fe^{2+} is contained in the half-cell with the anode, Fe^{2+} is the stronger reducing agent.

It can be concluded that the order of reducing agent strength, from strongest to weakest is $Sn^{2+} > Fe^{2+} > Br^-$.

22 Reduction occurs at the cathode, which is the positive electrode. The half-cell containing the cathode contains the stronger oxidising agent, and therefore, the higher reduction half-cell potential. Only the cations in solution can be reduced, so they are the oxidising agents in each half-cell.

In the first cell, the oxidising agents are A^{2+} and B^{2+}. Because B^{2+} is contained in the half-cell with the cathode, B^{2+} is the stronger oxidising agent.

In the second cell, the oxidising agents are A^{2+} and C^{2+}. Because A^{2+} is contained in the half-cell with the cathode, A^{2+} is the stronger oxidising agent.

In the third cell, the oxidising agents are C^{2+} and D^{2+}. Because C^{2+} is contained in the half-cell with the cathode, C^{2+} is the stronger oxidising agent.

It can be concluded that the order of reduction half-cell potentials, from highest to lowest is $B^{2+} > A^{2+} > C^{2+} > D^{2+}$.

23 **a** $3Cu^{2+}(aq) + 2Al(s) \rightarrow 3Cu(s) + 2Al^{3+}(aq)$

b no reaction

c $2Ag^{+}(aq) + Zn(s) \rightarrow 2Ag(s) + Zn^{2+}(aq)$

24 **a** $Ag^{+}(aq)$; $Mg^{2+}(aq)$

b $Mg(s)$; $Ag(s)$

c A coating of silver will form on the lead when it is placed in silver nitrate solution, because Ag^{+} ions are stronger oxidants than Pb^{2+} ions.

d zinc and magnesium

25 **a** **i** +3

ii +2

iii +4

b oxidation: $S_2O_4^{2-}(aq) + 2H_2O(l) \rightarrow 2HSO_3^{-}(aq) + 2H^{+}(aq) + 2e^{-}$

reduction: $S_2O_4^{2-}(aq) + 2H^{+}(aq) + 2e^{-} \rightarrow S_2O_3^{2-}(aq) + H_2O(l)$

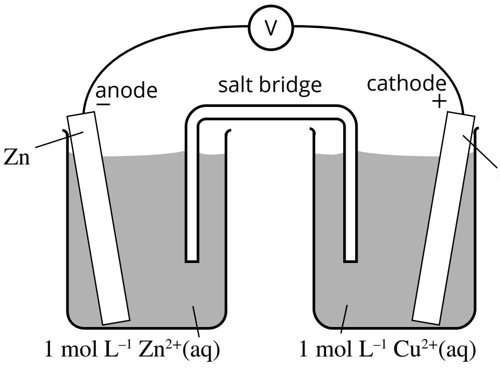
26 **a** $Zn(s) + 4OH^{-}(aq) + Cu^{2+}(aq) \rightarrow Zn(OH)_4^{2-}(aq) + Cu(s)$

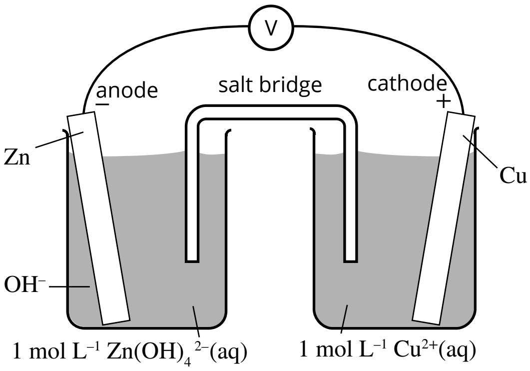
b The student should make two electrochemical cells consisting of the:

- Cu^{2+}/Cu half-cell and the ‘alkaline zinc half-cell’
- Cu^{2+}/Cu half-cell and the Zn^{2+}/Zn half-cell

as shown in the diagrams.

The cell voltages should be measured. And if they are identical, the two half-cells have the same *E*° values. (They would not be expected to have the same *E*° values.)





27 **a** The steel wool has undergone an oxidation reaction.

b Student answers will vary. Possible answers include: coating the steel with a stronger reducing agent and keeping the steel wool dry.

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Chapter 13 Rates of reactions

13.1 Collision theory

13.1 KEY QUESTIONS

1 A. A is incorrect because reactant particles only need to collide with energy equal to or greater than the activation energy for the reaction.

2 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

b $+3380\,kJ\,mol^{-1}$

3 $+167\,kJ\,mol^{-1}$

4 a very small activation energy

5 They must

- collide with the correct orientation
- have sufficient energy to allow bond breaking to occur (greater than or equal to the activation energy).

13.2 Measuring reaction rate

13.2 KEY QUESTIONS

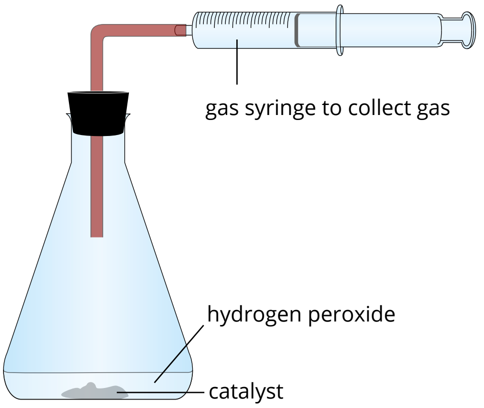
1 mass loss over time, gas volume produced over time, temperature change over time

2
- React Mg and HCl in a conical flask that is connected to a gas syringe and measure the gas volume at fixed time intervals; graph the data and determine the initial gradient.
- React Mg and HCl in a conical flask that has a calibrated pH probe in it; measure the pH at fixed time intervals; graph the data and determine the initial gradient.

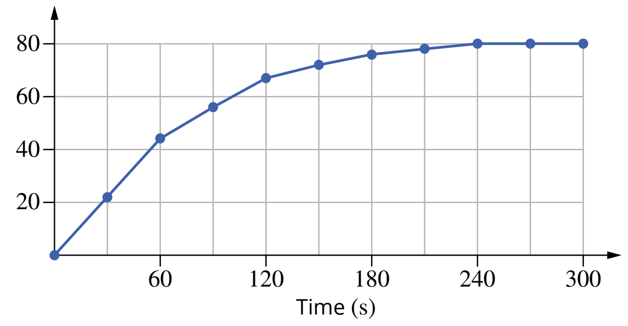
3 **a** $CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

b The reaction rate is decreasing, because the gradient of the graph is decreasing.

4 **a**



b



initial rate = initial gradient = $\frac{22}{30}$ = $7.3\,mL\,s^{-1}$

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5 $0.30\,\text{mol}\,L^{-1}$ has a higher concentration of copper(II) ions, so would absorb more light.

6 At regular time intervals: measure pH; use a colourimeter to measure the amount of light of an appropriate wavelength absorbed (the I_2 produced will cause the solution to turn orange/brown); or measure the electrical conductivity of the solution.

13.3 Effect of surface area, concentration and pressure on reaction rate

Worked example: Try yourself 13.3.1

USING COLLISION THEORY TO EXPLAIN CHANGES IN RATES OF REACTIONS

There have been many explosions in underground coal mines due to the presence of coal dust. Explain this observation in terms of collision theory.

Thinking	**Working**
Consider the states of the reactants.	Coal is a solid. In the mine, there would be lumps of coal and also powdered coal present.
Relate the states of the reactant to the factor that affects the reaction rate and explain in terms of collision theory.	The surface area of powdered coal is greater than that of solid coal. When the surface area increases, the frequency of collisions increases, so the rate of reaction increases.
Answer the question.	An explosion is a very fast reaction. The very large surface area of the coal dust allows for an increase in the frequency of collisions with reacting particles, which increases the reaction rate so much that explosions occur.

13.3 KEY QUESTIONS

1 C. Only decreasing the concentration of HCl(aq) will lower the collision frequency and decrease the reaction rate. Hydrochloric acid is not a gas; smaller zinc particles have a higher surface area and react faster; the volume of acid solution used does not alter the concentration of the acid, so the collision frequency is unaltered.

2 **a** *increasing* the surface area of solid reactants
  b *increasing* the concentration of a reactant in solution
  c *increasing* the pressure of gaseous reactants
  d *decreasing* the particle size of solid reactants

3 B, D, C, A

B is slowest because it has the lowest surface area and lowest concentration; all have the same temperature.

D is faster (than B) because it has a larger surface area.

C is faster (than D) because it has a larger surface area.

A is faster (than C) because it has a higher concentration.

4 Fine particles have a large surface area, resulting in a high frequency of collisions between aluminium particles and gas molecules (such as oxygen) in the air and hence a rapid reaction rate. The aluminium can burn vigorously and release a large quantity of heat.

5 **a** using smaller pieces of wood with a larger surface area
  b using a brick cleaner with a higher concentration

13.4 Effect of temperature on reaction rate

13.4 KEY QUESTIONS

1 C. The main reason why the reaction rate increases with temperature is that many more colliding particles have sufficient energy to overcome the activation energy barrier of the reaction.

2 A, B and C. A is true because a reaction with a large activation energy would have a low proportion of reactant particles colliding with energy equal to or greater than the activation energy for the reaction.

B is true because a low temperature means a low proportion of reactant particles collide with energy equal to or greater than the activation energy for the reaction.

C is true because for a reaction to occur, reactant particles need to collide with the correct orientation.

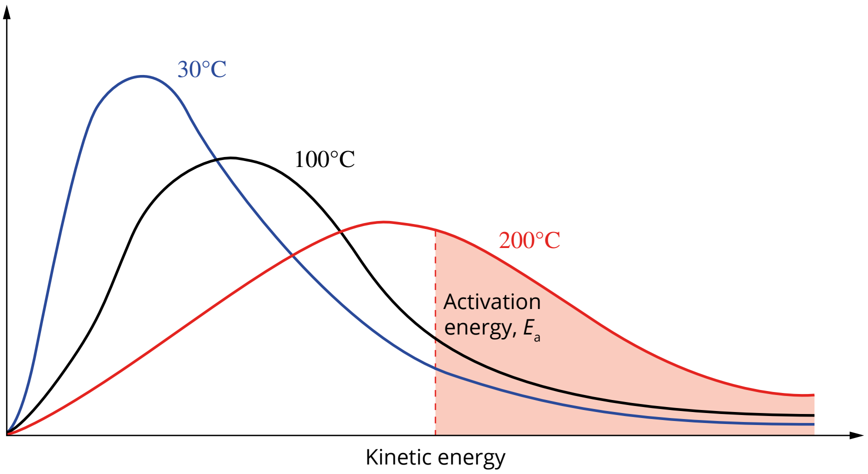
3 **a** At higher temperatures, the molecules that react to form fibreglass plastics have greater energy. They collide more frequently and are more likely to have a total energy exceeding the activation energy of the reaction involved, increasing the rate of reaction.

b At high altitude, such as in Nepal, air pressure is considerably lower than at any location in the Australian bush, so the water boils at a lower temperature in Nepal (up to 30°C lower). Thus, the average kinetic energy of the molecules in the potato will be lower, and they are less likely to have a total energy exceeding the activation energy of the reactions involved in cooking the potato. As a result, the potato will cook more slowly.

4 **a** a Maxwell–Boltzmann (distribution) curve or a kinetic energy distribution diagram

b

Maxwell–Boltzmann distribution curve of particles at varying temperatures



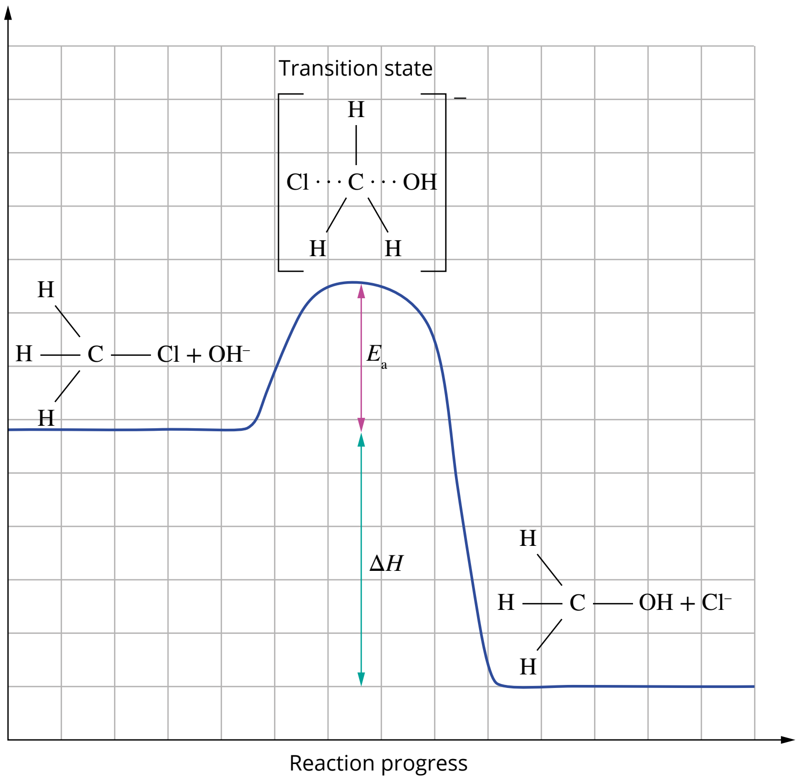
c The number of particles that have energy in excess of the activation energy is represented by the area under the curve to the right of the dashed line. At 200°C, a greater proportion of particles have energy greater than or equal to the activation energy of the reaction, E_a, than at 100°C.

CHAPTER 13 REVIEW

1 Reactant particles must: collide with each other, collide with sufficient energy to break the bonds within the reactants, and collide with the correct orientation to break the bonds within the reactants to allow the formation of new products.

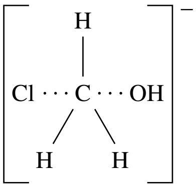
2 Collision 1 has the correct collision orientation, allowing bonds to break within the reactants and bonds to form within the products.

3 a Energy profile diagram for the reaction between CH_3Cl and OH^-



b Activation energy is the minimum amount of energy required by the reactants in order to form products in a reaction.

c



d The C–Cl bond is being broken in chloromethane. The C–O bond is beginning to form between chloromethane and the hydroxide ion.

4 a The single H–H bond in each hydrogen molecule and the double O=O bond in each oxygen molecule are broken in the course of this reaction.

b Two H–O bonds are formed in each new water molecule during the reaction.

c No reaction occurs until sufficient energy is supplied to give the reactant particles enough energy to overcome the activation energy barrier.

5 B. The rate of reaction is a measure of the change in concentration of a substance per unit time. The rate of bubbling, if a gas is produced, gives an indication of the rate.

6 B. At a lower temperature, the average kinetic energy of particles decreases, causing decreased collision frequency and a smaller proportion of collisions between reactant particles that have energy equal to or greater than the activation energy. Smaller solid particles have a larger surface area, leading to more frequent collisions.

7 The rate of reaction is a measure of the change in concentration of a substance over time. Units that correspond to concentration per unit time, such as $\text{mol L}^{-1}\text{s}^{-1}$ or $\text{mol L}^{-1}\text{min}^{-1}$, would be suitable.

8 **a** The student may have conducted a colourimetric investigation, as shown in Figures 13.2.6 and 13.2.7 (page 418 of the Student Book). As the reaction proceeds, the concentration of permanganate ion decreases, so the intensity of the purple solution should decrease. The student did not measure the volume of carbon dioxide produced, because if it was collected during the experiment it would have increased with time.

b graph B. At a higher temperature, the initial rate of reaction will be faster, so the gradient will be steeper.

9 D. An increase in concentration increases collision frequency between reactant particles.

10 **a** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$

b 2 mol L^{-1} HCl and powdered $CaCO_3$

11 **a** At lower temperatures, the molecules that react to cause the apple to brown have less energy. They collide less frequently and are less likely to have a total energy exceeding the activation energy of the reaction involved, decreasing the rate of reaction.

b Using pure oxygen gas instead of air increases the concentration of oxygen. This results in a higher frequency of collisions between hydrogen and oxygen molecules, and hence an increased reaction rate.

12 **a** $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

b $n(CaCO_3) = \frac{10.0}{100.09} = 0.0999\,\text{mol}$

$n(HCl) = 0.100\,\text{mol L}^{-1} \times 0.100\,\text{L} = 0.0100\,\text{mol}$

$\therefore$ $CaCO_3$ is in excess

c The rate of reaction could be measured by:

- the decrease in mass of the reaction mixture as $CO_2(g)$ escapes to the atmosphere
- the increase in pH measaured with a pH probe as acid is consumed.

d The rate of reaction with the smaller lumps will be faster. The smaller lumps have a larger surface area, so more collisions can occur per second.

e The rate of reaction could be increased by:

- increasing temperature: the average kinetic energy of the particles increases, causing increased collision frequency and also causing a greater proportion of the collisions between particles to have energy greater than the activation energy.
- increasing the concentration of the hydrochloric acid: there would be more HCl(aq) particles per unit volume, which would lead to more frequent collisions between reactants.

13 **a** A gas is produced, so mass is lost from the mixture.

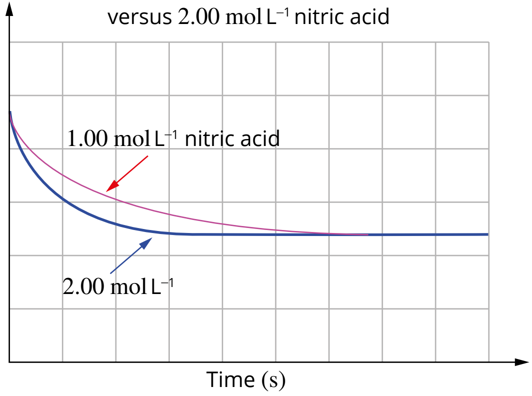
b
$$n(Cu) = \frac{m}{N} = \frac{5.00}{63.55} = 0.0787\,\text{mol}$$

$$n(HNO_3) = cV = 2.00 \times 0.500 = 1.00\,\text{mol}$$

$$n(HNO_3)\text{ reacting with Cu} = \frac{8}{3} \times n(Cu) = \frac{8}{3} \times 0.0787 = 0.210\,\text{mol}$$

0.210 < 1.00; therefore HNO_3 is in excess, so Cu is limiting.

c Decreasing mass of mixture indicating rate of reaction of copper with 1.00 mol L^{-1} versus 2.00 mol L^{-1} nitric acid

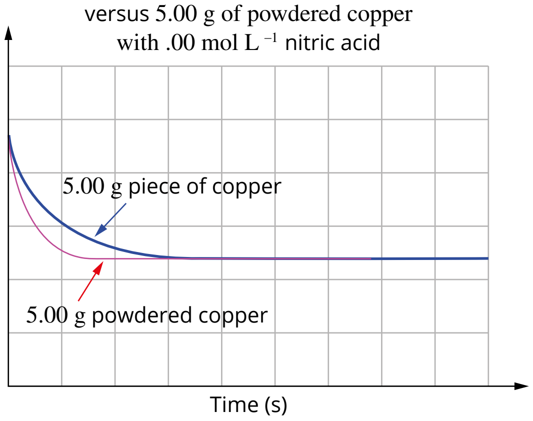


There would be a decreased rate of mass loss due to the lower nitric acid concentration.

($n(HNO_3) = cV = 1.00 \times 0.500 = 0.500\,\text{mol}$)

Copper would still be limiting, so the final mass would remain the same.

d Decreasing mass of mixture indicating rate of reaction of a 5.00 g piece of copper versus 5.00 g of powdered copper with .00 mol L^{-1} nitric acid



There would be an increased rate of mass loss due to the increased copper surface area.

Copper would still be limiting, so the final mass would remain the same.

14 D. According to collision theory, for a reaction to occur, the reactant particles must collide with sufficient energy equal to or greater than the activation energy of the reaction in order to break the bonds within the reactants.

15 **a** Higher body temperature increases the rate of reactions. Increased pulse and breathing rate increases the concentration of reactants, which also increases the rate of reactions.

b Lower body temperature decreases the rate of the metabolic reactions in the body that could cause brain damage.

16 IV, VI, II, V, III and I. IV is slowest because it has the lowest surface area, lowest concentration and lowest temperature.

VI is faster (than IV) because it has a larger surface area.

II is faster (than VI) because it has a higher concentration.

V is faster (than II) because it has a larger surface area.

III is faster (than V) because it has a higher temperature.

I is the fastest because it has the largest surface area, highest concentration and highest temperature.

17 **a** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants, temperature of the reaction

b **i** surface area of a solid reactant, concentration of reactants in a solution, pressure of any gaseous reactants

ii temperature of the reaction (also increases collision frequency)

18 To increase the rate of dissolution without ruining the toffee, you could:

- grind up the sugar crystals or use caster sugar
- use a cup of hot water to dissolve the sugar
- gently heat the sugar and water mixture while the sugar is dissolving
- stir the sugar and water mixture while the sugar is dissolving.

19 **a** The temperature of the new process (550°C) is much lower than the temperature of 1800°C in the blast furnace.

b $2Fe(s) + O_2(g) \rightarrow 2FeO(s)$

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

c high surface area of iron pellets; high temperature caused by trapped heat that was unable to escape rapidly

d Water would have caused the production of more hydrogen and increased the fire. The method used by the firefighters to extinguish the fire was to flood the hold with liquid nitrogen, which extinguished the surface fire but did not stop the deeper burning. A crane and clamshell bucket was then used to unload the iron into piles less than 1 m deep so the heat could escape.

20 **a** The energy of collisions is the same in both reactions because the temperature is the same. Because of an increase in surface area using powdered zinc compared with large zinc particles, the frequency of collisions between reactant particles increases.

b oxidation: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$; reduction: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

21 Crushed chalk has a greater surface area, which means the vinegar and chalk particles collide more frequently, increasing the rate of reaction.

Module 3 Review answers

Reactive chemistry

Multiple choice

1. D. When CO_2 sublimes it changes state but the change is not a permanent one. The other options are all chemical reactions in which new substances are formed.
2. C. The reaction is producing CaO and CO_2. The law of conservation of mass can be used to show that the mass of CO_2 formed will be equal to the difference in mass between $CaCO_3$ and CaO.
3. A. It is a reaction between an acid and a base that is forming a salt and water. This is a typical neutralisation reaction.
4. B. Applying the SNAAP rule, KNO_3 will be soluble because all nitrates are soluble, NH_4Cl is soluble because ammonium compounds are soluble and Na_2SO_4 is soluble because sodium compounds are soluble.
5. C. The equation is balanced correctly and it is an example of incomplete combustion as CO is formed rather than CO_2.
6. C. Aluminium reacts with oxygen to form aluminium oxide (Al_2O_3) which is an ionic compound consisting of Al^{3+} and O^{2-} ions.
7. B. Lithium metal reacts with water to produce hydrogen gas and a solution of lithium hydroxide. Zinc metal reacts with steam, producing hydrogen gas and zinc oxide. Copper does not react with dilute hydrochloric acid.
8. A. The reaction between zinc and oxygen involves a combination of zinc and oxygen atoms. Hydrogen is displaced from water in the reaction between sodium and water and in the reaction between iron and steam. Magnesium metal displaces Cu^{2+} ions from a solution of copper(II) sulfate.
9. C. A less reactive metal will lose its valence electrons less readily than a more reactive metal. Metal III has the smallest atomic radius indicating that it will have the strongest force of attraction between the valence electrons and the nucleus. Metal III has the highest ionisation energy, indicating it will lose its valence electrons less readily. Metal III also has the highest electronegativity, which means that it has the greatest attraction for electrons and consequently will less readily lose electrons.
10. D. There is no exchange of electrons between Ag^+ and Cl^-. This is a precipitation reaction.
11. A. Cu is less reactive than Mg so will not displace magnesium from a solution of its ions.
12. C. Salt and water have not been produced, so it is not a neutralisation reaction. There is a transfer of electrons from Mg to H^+ to form Mg^{2+} and H_2. The Mg loses electrons, so is oxidised by the H^+, which is therefore the oxidising agent.
13. B. The reaction tells us that metal X is more reactive than Pb as it displaces Pb from a solution of its salt. So A and D are not correct. Pb is sufficiently reactive to react with HCl to produce hydrogen gas, so X, being more reactive than Pb, will also react similarly with HCl, so B is correct. It is not known how much more reactive X is than Pb, or whether it is more or less reactive than Mg, so C could be true, but one cannot deduce from the data that it must be true.
14. B. A reducing agent causes reduction in another substance by donating electrons to it.
15. A. The strongest oxidising agent will be the species gaining electrons (i.e. on the left-hand side) with the highest E° value. The strongest reducing agent will be the species losing electrons (i.e. on the right-hand side) with the lowest E° value.
16. C. The anode is, by definition, the site of oxidation in any cell. In a galvanic cell, this is the electrode labelled negative.
17. D. The half-cell reactions will be as follows:

    At the positive cathode: $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$. Hence, $Cu^{2+}(aq)$ is consumed and the blue colour of the solution will fade.

    At the negative anode: $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$. Hence, Ni(s) is consumed and mass of the nickel electrode will decrease.
18. B. The activation energy represents the minimum energy required for a successful collision to occur between reactant particles. The area under the curve represents the total number of particles and the shaded area represents the number of particles with energy higher than the activation energy.
19. A. Increasing the temperature raises the average speed/energy of the particles, increasing the numbers of particles with higher energies. A catalyst reduces the activation energy, lowering the energy required for fruitful collisions.
20. C. A catalyst does not affect either the initial or the final states in a chemical reaction. It provides an alternative reaction pathway with a lower activation energy threshold.

Short answer

1 **a** precipitation. The $Cu(OH)_2(s)$ is a precipitate formed from two solutions.

b acid base. The reaction is between sulfuric acid and magnesium hydroxide. It produces a salt and water.

c synthesis. A single compound is formed from two smaller molecules.

d combustion. The products are an oxide of carbon and water.

e decomposition. The reactant is broken down into two smaller substances.

2 **a** $2HCl(aq) + Na_2CO_3(s) \rightarrow 2NaCl(aq) + H_2O(l) + CO_2(g)$

b $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$

c $Ba(s) + H_2SO_4(aq) \rightarrow BaSO_4(s) + H_2(g)$

d $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

e $HNO_3(aq) + KOH(aq) \rightarrow KNO_3(aq) + H_2O(l)$

3 **a** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

b synthesis

c $4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$

4 **a** $Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$

b $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

c $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_2(g) + H_2O(l)$

d $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

5 **a** $Fe(s) + 2H^+(aq) \rightarrow Fe^{2+}(aq) + H_2(g)$

b $2Rb(s) + 2H_2O(l) \rightarrow 2Rb^+(aq) + 2OH^-(aq) + H_2(g)$

c $Zn(s) + 2Ag^+ \rightarrow Zn^{2+}(aq) + 2Ag(s)$

6 **a** **i** Statement I uses qualitative data. Qualitative data can be observed but not measured.

ii Quantative data can be measured and expressed in numerical form. The quantitative data in statement II is underlined. 2.5 mL of hydrogen gas was collected in 60.0 s when a strip of magnesium weighing 0.5 g was placed into a test tube containing 10.0 mL of hydrochloric acid that had a concentration of $3.64\ \text{g L}^{-1}$.

b Similarities:

- Both series rank metals in relative order of reactivity.
- With some minor differences, the rankings are in agreement.

Differences:

- The activity series only ranks metals; the table of standard reduction potentials also includes non-metals and polyatomic ions in its ranking.
- The ranking in the table of standard reduction potentials is determined under controlled conditions. These conditions are not controlled in the determination of the activity series.
- While both series can be used to predict the reactions of metals with ions of other metals, the table of standard reduction potentials can also be used to predict reactions involving non-metals.

7 **a** $O_2(g) + 2H_2O(l) + 2Fe(s) \rightarrow 4OH^-(aq) + 2Fe^{2+}(aq)$

b The oxidising agent is the reactant that causes the other reactant to be oxidised. Fe loses electrons so it is oxidised. Oxygen is therefore the oxidising agent and Fe is the reducing agent.

c **i** Zn(s) is more reactive than iron, so will lose electrons in preference to Fe. When there is some Zn attached to the Fe, the Zn is oxidised and the Fe is protected.

ii Sn is less reactive than Fe, so if Fe were to be in contact with oxygen, it is the Fe that will preferentially corrode. If the Sn completely covers the Fe, the oxygen is prevented from reaching the Fe and the Fe is protected.

8 **a** **i** Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

Reduction: $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$

ii Overall: $2Al(s) + 3Br_2(aq) \rightarrow 2Al^{3+}(aq) + 6Br^-(aq)$

b **i** Oxidation: $S_2O_3^{2-}(aq) + 5H_2O(l) \rightarrow 2SO_4^{2-}(aq) + 10H^+(aq) + 8e^-$

Reduction: $ClO^-(aq) + 2H^+(aq) + 2e^- \rightarrow Cl^-(aq) + H_2O(l)$

ii Overall: $4ClO^-(aq) + S_2O_3^{2-}(aq) + H_2O(l) \rightarrow 2SO_4^{2-}(aq) + 4Cl^-(aq) + 2H^+(aq)$

c **i** Oxidation: $H_2O_2(aq) \rightarrow O_2(aq) + 2H^+(aq) + 2e^-$

Reduction: $MnO_4^-(aq) + 4H^+(aq) + 3e^- \rightarrow MnO_2(s) + 2H_2O(l)$

ii Overall: $3H_2O_2(aq) + 2MnO_4^-(aq) + 2H^+(aq) \rightarrow 3O_2(aq) + 2MnO_2(s) + 4H_2O(l)$

9 **a** $E°(Sn^{2+}/Sn) = –0.14$ V and $E°(Fe^{2+}/Fe) = –0.41$ V

  i cell voltage = (–0.14) – (–0.41) = 0.27 V

  ii Fe^{2+}/Fe

  iii $Sn^{2+}(aq) + Fe(s) \rightarrow Sn(s) + Fe^{2+}(aq)$

b $E°(Fe^{3+}/Fe^{2+}) = +0.77$ V and $E°(Al^{3+}/Al) = –1.71$ V

  i cell voltage = (+0.77) – (–1.71) = 2.48 V

  ii Al^{3+}/Al

  iii $3Fe^{3+}(aq) + Al(s) \rightarrow 3Fe^{2+}(s) + Al^{3+}(aq)$

c $E°(H^+/H_2) = 0.00$ V and $E°(I_2/I^-) = +0.54$ V

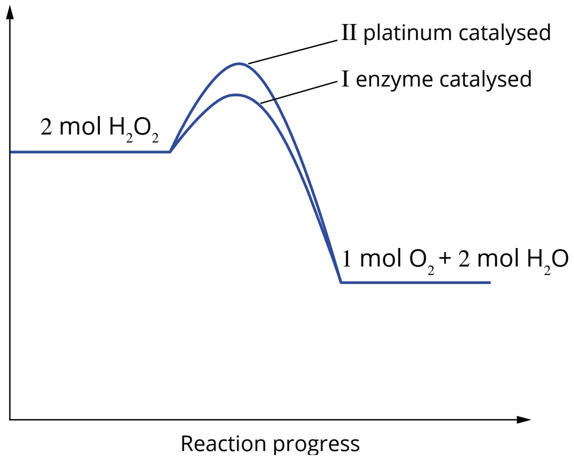
  i cell voltage = (+0.54) – (0.00) = 0.54 V

  ii H^+/H_2

  iii $I_2(aq) + H_2(g) \rightarrow 2I^-(aq) + 2H^+(aq)$

10 **a** The function is to act as a catalyst and lower the activation energy, giving a higher reaction rate.

b **Energy profiles of hydrogen peroxide decomposition**



c System I would be faster. The lower activation energy requirement means that there is a greater probability of any given collision having sufficient energy for a successful reaction, so fruitful collisions will occur more frequently.

Extended response

1 **a** $Pb(NO_3)_2(aq) + 2KCl(aq) \rightarrow PbCl_2(s) + 2KNO_3(aq)$

b $Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

c **i** lead(II) chloride

  ii potassium ions and nitrate ions

d **i** $n(Pb(NO_3)_2) = \frac{m}{M} = \frac{3.15}{331.2} = 0.00951$ mol

  ii $n(PbCl_2)$ formed = $n(Pb(NO_3)_2)$ reacted = 0.00951 mol

  So $m(PbCl_2)$ formed = $n \times M = 0.00951 \times 278.1 = 2.64$ g

2 **a** Place a sample of each metal into separate test-tubes containing a solution of Cu^{2+} ions. Observe and record which metals develop a copper coating. The copper coating indicates that the metal has displaced Cu^{2+} ions from solution. Repeat the experiment with each of the other metal ion solutions.

b The activity series of the metals listed is Mg > Zn > Fe > Sn > Pb > Cu.

c **i** Fe^{2+} or Pb^{2+} or Sn^{2+} or Cu^{2+}. A metal higher in the activity series will displace, from solution, the ion of a metal lower down the series. Zinc metal displaces Fe^{2+}, Pb^{2+}, Sn^{2+} and Cu^{2+} ions from solution.

  ii $Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)$ or

  $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$ or

  $Zn(s) + Sn^{2+}(aq) \rightarrow Zn^{2+}(aq) + Sn(s)$ or

  $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

d **i** Magnesium. Zinc ions are displaced from solution by a metal that is more reactive than zinc metal, i.e. magnesium.

  ii $Mg(s) + Zn^{2+}(aq) \rightarrow Mg^{2+}(aq) + Zn(s)$

3 Metal reactivity decreases from left to right across a period and increases down a group. The reactivity of a metal is determined by the ease with which a metal atom loses its valence electrons.

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Sodium and potassium are both group 1 metals. Sodium is the first element in the third period and potassium is the first element in the fourth period. Since potassium has one more electron shell than sodium, its atomic radius is greater. The shielding effect of the inner electron shells is also greater in potassium. Consequently, the force of attraction between the potassium atom's valence electron and its nucleus is less. As a result, the ionisation energy and electronegativity of potassium are less than that of sodium. The potassium atom is more reactive, since it will lose the outer shell or valence electrons more readily than a sodium atom.

Sodium and magnesium are sequential elements in the third period of the periodic table and are found in groups 1 and 2, respectively.

The electronic configuration of sodium is 2,8,1, and magnesium has an electronic configuration of 2,8,2. The number of occupied or partially occupied shells stays constant, but the core charge increases with increasing atomic number. Thus, the strength of attraction between valence electrons and the nucleus increases from sodium to magnesium. As a consequence, the atomic radius decreases across the period from left to right. The first ionisation energy and electronegativity increase due to the combined effects of increased core charge and decreased atomic radius.

The energy required to form an ion also depends upon the number of electrons that are lost in formation of the metal cation. The sodium atom only loses a single electron when it forms the Na^+ ion. The formation of a magnesium ion involves the loss of two electrons. A greater amount of energy is required to remove the second valence electron than the first valence electron. More energy is required to form the Mg^{2+} ion than the Na^+ ion. Even more energy is required to form the Al^{3+} ion, since this involves the loss of the three electrons.

As a result of the factors described above, the ability of metals to form ions by losing electrons decreases across a period and increases down a group. This is reflected in the decreasing order of their reactivities: K > Na > Ca > Mg.

4 **a** arrow A

b O_2 is the oxidising agent, and it enters the 'air space' from outside.

c $4Al(s) + 3O_2(g) + 6H_2O(l) \rightarrow 4Al^{3+}(aq) + 12OH^-(aq)$

d The cathode is the site of the reduction reaction involving oxygen/air, which must be in contact with the electrolyte. A porous electrode will allow a large surface area for this reaction and thereby maximise the rate of reaction and hence the current/power output of the cell.

e From the table of standard reduction potentials:

cathode reaction:	$O_2(g) + 2H_2O(l) + 4e^- \rightarrow 4OH^-(aq)$	$E^\circ = +0.40\,V$
anode reactions:	$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	$E^\circ = -0.76\,V$
	$Ag(s) \rightarrow Ag^+(aq) + e^-$	$E^\circ = +0.80\,V$

Since $E^\circ(Zn^{2+}/Zn) < E^\circ(O_2,H_2O/OH^-)$, a spontaneous cell reaction will occur.

Since $E^\circ(Ag^+/Ag) > E^\circ(O_2,H_2O/OH^-)$, a spontaneous cell reaction will not occur.

f smaller, since $E^\circ(Zn^{2+}/Zn) > E^\circ(Al^{3+}/Al)$

$E^\circ(O_2,H_2O/OH^-) - E^\circ(Zn^{2+}/Zn) = (+0.40) - (-0.76) = 1.16\ V$

$E^\circ(O_2,H_2O/OH^-) - E^\circ(Al^{3+}/Al) = (+0.40) - (-1.71) = 2.10\ V$

g With an acidic electrolyte, the cathode reaction becomes:

cathode reaction:	$O_2(g) + 4H^+(aq) + 4e^- \rightarrow 2H_2O(l)$	$E^\circ = +1.23\,V$
anode reactions:	$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	$E^\circ = -0.76\,V$
	$Ag(s) \rightarrow Ag^+(aq) + e^-$	$E^\circ = +0.80\,V$

Since $E^\circ(Zn^{2+}/Zn) < E^\circ(O_2,H^+/H_2O)$, a spontaneous cell reaction will occur.

Since $E^\circ(Ag^+/Ag) < E^\circ(O_2,H^+/H_2O)$, a spontaneous cell reaction will occur.

Essentially, O_2 becomes a stronger oxidising agent in acidic conditions, and is capable of spontaneous reaction with the relatively weak reducing agent, Ag.

5 **a** The concentration of sodium thiosulfate is higher, so there will be more frequent collisions between reactant particles. While the probability of any particular collision being successful is unchanged, the higher frequency of collisions overall will increase the frequency of successful collisions and hence the rate of reaction.

b The higher temperature means the particles in the mixture will be moving faster and hence colliding more frequently. Also, and more significantly, having faster-moving particles means there will be a higher proportion of collisions that are successful, since there is a greater chance of a collision having energy greater than the activation energy. These two factors combine to give a higher frequency of fruitful collisions and hence a higher rate of reaction.

c Apart from changing concentration and temperature, reaction rates can be affected by:

- changing reactant pressures for gases
- the presence of a catalyst
- changing the available surface area for reactions involving solids or for non-homogenous mixtures.

Chapter 14 Energy changes in chemical reactions

14.1 Exothermic and endothermic reactions

14.1 KEY QUESTIONS

1 D

2 **a** $0.180 \times (1 \times 10^3) = 180\,\text{kJ}$

  b $(1.5 \times 10^6) \times 10^{-3} = 1.5 \times 10^3\,\text{kJ}$

  c $10.0 \times 10^{-3} = 0.0100\,\text{kJ}$ or $1.00 \times 10^{-2}\,\text{kJ}$

  d $(2.0 \times 10^{-3}) \times 10^{-3} = 2.0 \times 10^{-6}\,\text{kJ}$

3 In chemistry, the system is usually the chemical reaction, whereas the surroundings refers to everything else (for example, the beaker or test-tube in which the reaction takes place).

4 In any reaction, the total amount of the chemical energy of the reactants is made up of the bonds between atoms within the reactants. If the total amount of chemical energy within the reactants is less than the total amount of chemical energy within the products, energy must be supplied to the system, and the reaction is said to be endothermic.

14.2 Thermochemical equations and energy profile diagrams

Worked example: Try yourself 14.2.1

CALCULATING ΔH FOR A RELATED EQUATION

Carbon reacts with hydrogen according to the equation:

$$6C(s) + 3H_2(g) \rightarrow C_6H_6(g) \quad \Delta H = +49\,\text{kJ mol}^{-1}$$

Calculate ΔH for the reaction represented by the equation:

$$3C_6H_6(g) \rightarrow 18C(s) + 9H_2(g)$$

Thinking	**Working**
The reaction has been reversed in the second equation, so the sign for ΔH is changed to the opposite sign.	ΔH for the second equation is negative.
Identify how the mole amounts in the equation have changed.	The mole amount of C_6H_6 has changed from 1 to 3, of H_2 has changed from 3 to 9 and of C has changed from 6 to 18. They have all tripled.
Identify how the magnitude of ΔH will have changed for the second equation.	The mole amounts of the chemicals have all tripled, so ΔH will also have tripled.
Calculate the new magnitude of ΔH. (You will write the sign of ΔH in the next step.)	$3 \times 49 = 147$
Write ΔH for the second equation, including the sign.	$\Delta H = -147\,\text{kJ mol}^{-1}$

14.2 KEY QUESTIONS

1 A negative ΔH value indicates that a reaction is exothermic. This is because the enthalpy of the reactants is greater than the enthalpy of the products. Energy is being released *to* the surroundings.

2 $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l) \quad \Delta H = -890\,\text{kJ mol}^{-1}$

3 It would be lower because the change of state of the H_2O from liquid to gas will require that energy be absorbed.

4 **a** endothermic

  b The total enthalpy of the product (HI) is greater than that of the reactants (hydrogen gas and iodine gas).

  c The activation energy is greater than the magnitude of the ΔH value.

5 $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(aq) + 6O_2(g) \quad \Delta H = +2803\,\text{kJ mol}^{-1}$

14.3 Heat of combustion

Worked example: Try yourself 14.3.1

CALCULATING ENERGY RELEASED BY A SPECIFIED MASS OF A PURE SUBSTANCE

Calculate the amount of energy released when 5.40 kg of propane (C_3H_8) is burnt in an unlimited supply of oxygen.

Thinking	**Working**
Calculate the number of mol of the compound using: $n = \frac{m\,(\text{in grams})}{M}$	$n(C_3H_8) = \frac{m}{M}$ $= \frac{5.40 \times 10^3}{44.09}$ $= 122.5\ \text{mol}$
Multiply the number of mol by the heat of combustion.	Energy $= n \times \Delta H_c$ $= 122.5 \times 2220$ $= 2.72 \times 10^5\ \text{kJ}$ or 272 MJ

14.3 KEY QUESTIONS

1 $2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$

2 propane:

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

$= \frac{1.000}{44.09}$

$= 2.268 \times 10^{-2}\ \text{mol}$

Energy $= n \times \Delta H_c$

$= (2.268 \times 10^{-2}) \times 2220$

$= 50.35\ \text{kJ}$

octane:

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

$= \frac{1.000}{114.22}$

$= 8.755 \times 10^{-3}\ \text{mol}$

energy $= n \times \Delta H_c$

$= (8.755 \times 10^{-3}) \times 5450$

$= 47.71\ \text{kJ}$

Propane releases more energy per gram than octane does.

3 $C_2H_5OH(l) + 2O_2(g) \rightarrow 2CO(g) + 3H_2O(l)$

4 **a** $n = \frac{m}{M}$

$= \frac{250}{16.04}$

$= 15.6\ \text{mol}$

energy $= n \times \Delta H_c$

$= 15.6 \times 890$

$= 1.39 \times 10^4\ \text{kJ}$ or 13.9 MJ

b $n = \frac{m}{M}$

$= \frac{9.64 \times 10^3}{44.09}$

$= 219\ \text{mol}$

energy $= n \times \Delta H_c$

$= 219 \times 2220$

$= 4.85 \times 10^5\ \text{kJ}$ or 485 MJ

c $n = \frac{m}{M}$

$= \frac{403 \times 10^3}{46.07}$

$= 8.75 \times 10^3$ mol

Energy $= n \times \Delta H_c$

$= (8.75 \times 10^3) \times 1367$

$= 1.20 \times 10^7$ kJ or 1.20×10^4 MJ or 12.0 GJ

14.4 Determining the heat of combustion

Worked example: Try yourself 14.4.1

CALCULATING THE AMOUNT OF ENERGY REQUIRED TO HEAT A SPECIFIED MASS OF WATER (USING SPECIFIC HEAT CAPACITY)

Calculate the heat energy, in kJ, needed to increase the temperature of 375 mL of water by 45°C.

Thinking	Working
Change the volume of water, in mL, to mass of water, in g. Remember that 1 mL of water has a mass of 1 g.	1 mL of water has a mass of 1 g, so 375 mL of water has a mass of 375 g.
Find the specific heat capacity of water from the data in Table 14.4.1 (page 455 of the Student Book).	The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.
To calculate the quantity of heat energy in joules, use the formula: $q = m \times c \times \Delta T$ Remember, a change of 1°C is the same as a change of 1 K.	$q = 375 \times 4.18 \times 45$ $= 7.05 \times 10^4$ J
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply by 10^{-3}.	$q = 7.05 \times 10^4 \times 10^{-3}$ $= 70.5$ kJ

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Worked example: Try yourself 14.4.2

CALCULATING THE HEAT OF COMBUSTION OF A FUEL FROM EXPERIMENTAL DATA

0.295 g of ethanol (C_2H_5OH) underwent complete combustion in a spirit burner. The heat energy released was used to heat 100 mL of water. The temperature of the water rose from 19.56°C to 38.85°C. Calculate the heat of combustion of ethanol in $kJ\,mol^{-1}$.

Thinking	Working
Calculate the temperature change of the water. Remember, a change of 1°C is the same as a change of 1 K.	$\Delta T = 38.85 - 19.56$ $= 19.29\,K$
Use the specific heat capacity of water to determine the energy required to heat the water. Use the formula: $q = m \times c \times \Delta T$ (m in this formula is the mass of water.)	$q = 100 \times 4.18 \times 19.29$ $= 8063\,J$
Express the quantity of energy in kJ. Remember that to convert from J to kJ, you multiply by 10^{-3}.	$q = 8063 \times 10^{-3}$ $= 8.063\,kJ$
Calculate the amount, in mol, of ethanol, using the formula: $n = \frac{m}{M}$	$n = \frac{0.295}{46.07}$ $= 0.00640\,mol$
Determine the heat of combustion of ethanol, in $kJ\,mol^{-1}$. $\text{Heat of combustion} = \frac{\text{heat energy released by sample}}{\text{amount of sample (in mol)}}$	$\text{heat of combustion} = \frac{-8.063}{0.00640}$ $= -1.26 \times 10^3\,kJ\,mol^{-1}$ (Note: The negative sign indicates that the reaction released energy, causing the temperature to rise.)

Worked example: Try yourself 14.4.3

WRITING A THERMOCHEMICAL EQUATION USING HEAT OF COMBUSTION DATA

Write a thermochemical equation for the complete combustion of ethane (C_2H_6).

Thinking	Working
Add oxygen as a reactant and carbon dioxide and water as the products in the equation.	$C_2H_6(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$
Balance the carbon, hydrogen, then oxygen atoms. Add states.	$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$
Obtain the heat of combustion, in $kJ\,mol^{-1}$, from Table 14.4.2 (page 459 of the Student Book).	The heat of combustion of ethane is $-1560\,kJ\,mol^{-1}$.
Determine ΔH for the thermochemical equation by multiplying the heat of combustion by the coefficient of the fuel in the balanced equation.	$\Delta H = 2 \times -1560 = -3120\,kJ\,mol^{-1}$
Write the thermochemical equation.	$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$ $\Delta H = -3120\,kJ\,mol^{-1}$

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Worked example: Try yourself 14.4.4

DETERMINING % EFFICIENCY FROM EXPERIMENTAL DATA

A student adds 26.0 kJ of heat energy to a calorimeter containing 100.0 g of water. The temperature of the water rises from 23.5°C to 58.1°C. What is the percentage efficiency of the calorimeter?

Thinking	**Working**
Identify the experimental value of the energy.	26.0 kJ is the experimental value of the energy.
Calculate the temperature change of the water. Remember, a change of 1°C is the same as a change of 1 K.	$\Delta T = 58.1 - 23.5$ $= 34.6\text{ K}$
Use the specific heat capacity of water to determine the expected energy required to heat the water. Use the formula: $q = m \times c \times \Delta T$ (m in this formula is the mass of water.)	$q = m \times c \times \Delta T$ $= 100.0 \times 4.18 \times 34.6$ $= 14\,463\text{ J}$ or 14.46 kJ
Use the % efficiency relationship to determine the efficiency of the calorimeter $\%\text{ efficiency} = \frac{\text{expected energy absorbed}}{\text{actual energy supplied}} \times 100\%$ Remember, expected and experimental values must be in the same units.	$\%\text{ efficiency} = \frac{\text{expected energy absorbed}}{\text{actual energy supplied}} \times 100\%$ $= \frac{14.46}{26.0} \times 100\%$ $= 55.6\%$
Answer the question.	55.6% of the energy supplied was absorbed by the calorimeter, and the rest was lost to the surrounding environment.

14.4 KEY QUESTIONS

1 $q = m \times c \times \Delta T$

$= 1.00 \times 10^3 \times 4.18 \times 25.0$

$= 1.05 \times 10^5\text{ J}$

$= 1.05 \times 10^2\text{ kJ}$

2 It is assumed that all of the energy released by the combustion of the fuel is used to heat the water. In reality, some of the energy will be used to heat the metal can or be lost to the surroundings.

3 $q = m \times c \times \Delta T$

$= 300 \times 4.18 \times 1.78$

$= 2.23 \times 10^3\text{ J}$

$= 2.23\text{ kJ}$

$\Delta H_c = \frac{q}{n}$

$= \frac{-2.23}{1.00 \times 10^{-3}}$

$= -2.23 \times 10^3\text{ kJ mol}^{-1}$ (The answer is negative because the reaction is exothermic.)

4 $q = m \times c \times \Delta T$

$= 500 \times 4.18 \times 11.5$

$= 24\,035\text{ J}$

$= 24.035\text{ kJ}$

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

$= \frac{0.500}{58.12}$

$= 0.00860\text{ mol}$

$\Delta H_c = \frac{q}{n}$

$= \frac{-24.035}{8.60 \times 10^{-3}}$

$= -2.79 \times 10^3\text{ kJ mol}^{-1}$ (The answer is negative because the reaction is exothermic.)

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5 **a** $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(l)\quad \Delta H = -890\,kJ\,mol^{-1}$

b $q = m \times c \times \Delta T$

$= 500 \times 4.18 \times 80.0$

$= 167\,200\,J$

$= 167.2\,kJ$

$\Delta H_c = \frac{q}{n}$

$890 = \frac{167.2}{n}$

$n = \frac{167.2}{890}$

$= 0.1879\ mol$

$m = n \times M$

$= 0.1879 \times 16.04$

$= 3.01\,g$

6 $q = m \times c \times \Delta T$

$= 150.0 \times 4.18 \times (97.0 - 16.0)$

$= 50\,787\,J$ or $50.79\,kJ$

$$\%\ \text{efficiency} = \frac{\text{expected energy absorbed}}{\text{actual energy supplied}} \times 100\%$$

$$= \frac{50.79}{73.2} \times 100$$

$$= 69.4\%$$

14.5 Enthalpy of dissolution

Worked example: Try yourself 14.5.1

DETERMINING THE ENTHALPY OF A DISSOCIATION PROCESS

When 5.00 g of ammonium nitrate (NH_4NO_3) ($M = 80.05\ g\,mol^{-1}$) is added to 100.0 g of water in a solution calorimeter, it dissolves according to the equation:

$$NH_4NO_3(s) \xrightarrow{H_2O(l)} NH_4^+(aq) + NO_3^-(aq)$$

The temperature of the water decreases from 23.0°C to a minimum temperature of 19.1°C. Determine the enthalpy of dissolution.

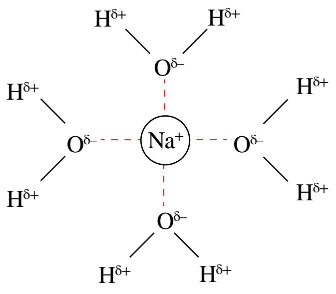
Thinking	Working
Determine the temperature change (ΔT). Remember, a change of 1°C is the same as a change of 1 K.	$\Delta T = 23.0 - 19.1$ $= 3.9\,K$
Determine the amount of energy absorbed/released by the water, resulting in the temperature change, using the expression: $q = m \times c \times \Delta T$ where m is the mass of the water.	$q = m \times c \times \Delta T$ $= 100 \times 4.18 \times 3.9$ $= 1630\,J$
Convert energy to kJ.	$1630 \times 10^{-3} = 1.63\,kJ$
Determine the amount of solute, in mol, using the formula: $n = \frac{m}{M}$	$n = \frac{m}{M}$ $= \frac{5.00}{80.05}$ $= 0.0625\,mol$
Determine ΔH_{soln} for the dissolution reaction by dividing the energy by the amount of substance according to: $\Delta H_{soln} = \frac{q}{n(\text{solute})}$ As the temperature decreased, the reaction is endothermic and hence ΔH_{soln} is positive.	$\Delta H_{soln} = \frac{q}{n(\text{solute})}$ $= \frac{1.63}{0.0625}$ $= +26\,kJ\ mol^{-1}$

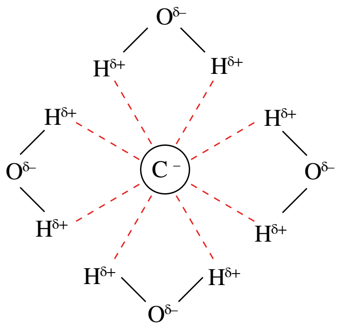
14.5 KEY QUESTIONS

1 $NaNO_3(s) \xrightarrow{H_2O(l)} Na^+(aq) + NO_3^-(aq)$

$Ca(OH)_2(s) \xrightarrow{H_2O(l)} Ca^{2+}(aq) + 2OH^-(aq)$

2 The positive sodium ion attracts the partial negative charges on the oxygen atoms in the water molecule. The negative chloride ion attracts the partial positive charges on the hydrogen atoms in the water molecule.





- - - - represents ion–dipole interaction

3 When solid potassium nitrate dissolves in water, the ionic bonds between ions must be broken, and new ion–dipole bonds form between ions and water molecules. The ions are said to be hydrated and the process is known as dissociation.

4 *Insulating* a calorimeter improves the accuracy of measurement of the energy changes occurring during a chemical reaction. Heat energy can be *lost* from a calorimeter, so a *lid* is a useful form of insulation. If the reaction occurring in a calorimeter is *exothermic*, the temperature of the water *increases*. If the reaction occurring is *endothermic*, the temperature of the water in a calorimeter decreases.

5 Determine the temperature change of the water.

$\Delta T = 38.90 - 17.90$

$= 21.00\,K$

Determine the amount of energy absorbed by the water.

$q = m \times c \times \Delta T$

$= 50.00 \times 4.18 \times 21.00$

$= 4389\,J$

$= 4.39\,kJ$

Determine the amount of calcium chloride, in mol.

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

$= \frac{6.00}{110.98}$

$= 0.0541\,mol$

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Determine the enthalpy of dissolution.

$$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$$
$$= \frac{-4.39}{0.0541}$$
$$= -81.2\,\text{kJ mol}^{-1}$$ (The answer is negative because the process is exothermic.)

6 Determine the amount of KNO_3, in mol.

$$n = \frac{m}{M}$$
$$= \frac{2.00}{101.11}$$
$$= 0.0198\,\text{mol}$$

Determine the amount of energy absorbed during dissolution.

$$q = \Delta H_{\text{soln}} \times n(\text{solute})$$
$$= 34.89 \times 0.0198$$
$$= 0.690\,\text{kJ}$$
$$= 690\,\text{J}$$

Determine the temperature change of the water.

$$\Delta T = \frac{q}{m \times c}$$
$$= \frac{690}{50.00 \times 4.18}$$
$$= 3.30\,\text{K}$$

Determine the final temperature of the water.

$$T_f = T_i - \Delta T$$
$$= 24.0 - 3.30$$
$$= 20.7\,^{\circ}\text{C}$$

(Note: The temperature of the water decreases because the dissociation process is endothermic.)

14.6 Catalysts

14.6 KEY QUESTIONS

1 D. The catalyst does not alter the energy of either the reactants (V + W) or the products (X + Z). Because it lowers the energy of the transition state, the activation energy, E_a, of both the forward reaction (A) and the backward reaction (C) are decreased.

2 **a** A catalyst is a substance that increases the rate of a chemical reaction without itself undergoing permanent change.

b Activation energy is the minimum amount of energy that colliding reactant particles need in order for successful collisions to occur that lead to a reaction.

3 When salt is mixed with sugar, the salt acts as a catalyst and lowers the activation energy of the combustion reaction between sugar and oxygen.

4 **a** Reactions involving a heterogeneous catalyst take place at the surface of the catalyst. Reactants form bonds with the catalyst, lowering the activation energy of reactions and allowing them to proceed more rapidly.

b A porous pellet has a much larger surface area than a solid lump. More reactants may be in contact with the surface of a porous pellet at any instant, producing a faster rate of reaction.

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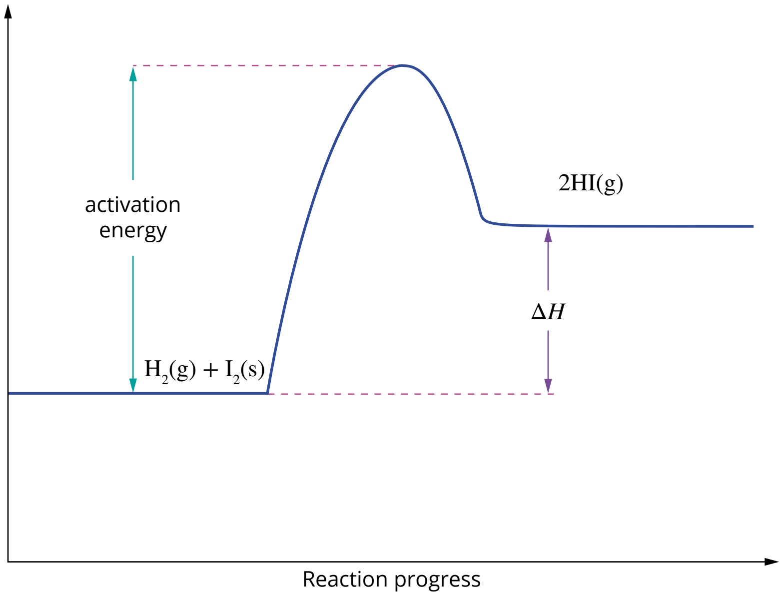
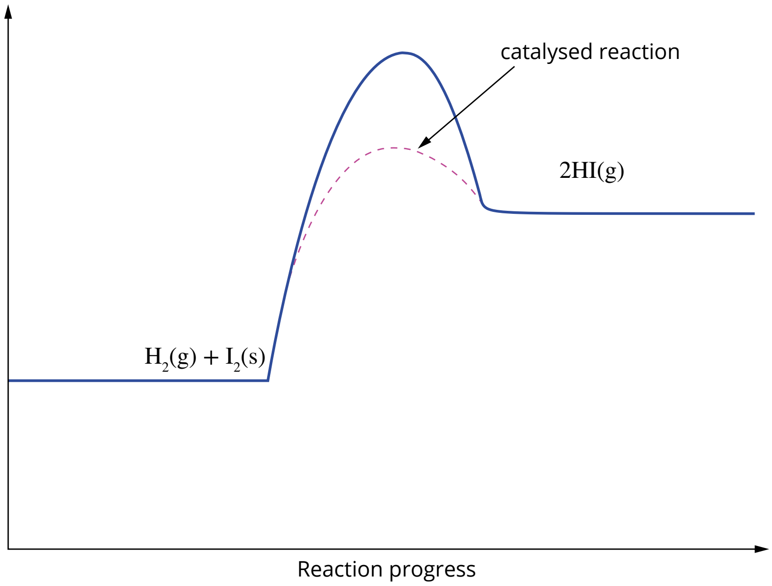
5 **a** Energy profile diagram for hydrogen iodide formation without a catalyst



b The reaction is endothermic.

c Energy profile diagram for hydrogen iodide formation with a catalyst



CHAPTER 14 REVIEW

1 **a** $2205 \times 10^{-3} = 2.21$ kJ

b $0.152 \times 10^{3} = 152$ J

c $1\,890\,000 \times 10^{-6} = 1.89$ MJ

d $0.0125 \times 10^{3} = 12.5$ kJ

2 **a** exothermic (Heat and light energy are released to the surrounding environment by the combustion of wood.)

b endothermic (Thermal energy is absorbed from the surrounding environment to melt ice.)

c endothermic (Electrical energy is consumed from a power supply as a battery is recharged.)

d exothermic (Heat energy is released to the surrounding environment as organisms in the compost heap decompose the plant material. The temperature of the compost heap rises as a consequence.)

3 C. Some energy is always absorbed to break bonds in the reactants.

4 **a** true

b false (All reactions have an activation energy. Reactions that start straight away are likely to have a very low activation energy and occur very readily.)

c false (All reactions require the absorption of some energy so that the reaction can proceed.)

d true

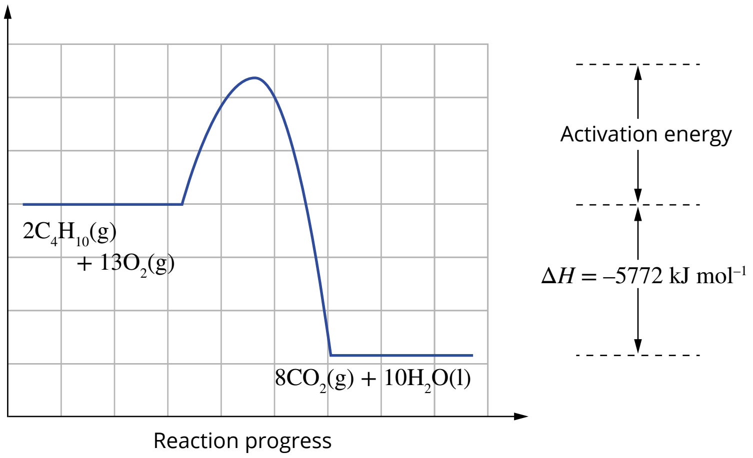
5 **a** exothermic (because ΔH is negative)

b The value of ΔH would now be doubled because the amount of all of the reactants and products has been doubled. $\Delta H = -5238$ kJ mol^{-1}

6 If a chemical equation is written for an endothermic reaction, ΔH is positive, telling you that energy is absorbed as the reaction proceeds. The enthalpy of the products must be higher than the enthalpy of the reactants. If this reaction is reversed, the enthalpy of the reactants is now higher than the enthalpy of the products. The reaction releases energy as it proceeds, so is exothermic. ΔH becomes negative.

7 **a** higher energy in bonds of reactants

b Energy profile diagram of butane combustion



8 $C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

9 $2C_4H_{10}(g) + 9O_2(g) \rightarrow 8CO(g) + 10H_2O(l)$

10 $n = \frac{m}{M}$

$= \frac{5.00\times10_3}{16.04}$

$= 311.7$ mol

Energy $= n \times \Delta H_c$

$= 311.7 \times 890$

$= 277\,431$ kJ

$= 277\,431 \times 10^{-3} = 277$ MJ

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11 These questions use the formula relating specific heat capacity and energy. It is important to use the correct units. As the density of water is 1 g mL^{-1}, 1 mL of water weighs 1 g. Remember to convert kg to g. As the temperature is only a difference, the unit can be either °C or K.

The formula is: energy = mass of water × specific heat capacity × temperature rise, which can be rewritten for water as:

$$q = m \times 4.18 \times \Delta T$$

a $q = 100 \times 4.18 \times 60.0$
$= 25\,080\ \text{J}$
$= 25.1\ \text{kJ}$

b $q = 250 \times 4.18 \times 75.0$
$= 78\,375\ \text{J}$
$= 78.4\ \text{kJ}$

c $q = 1.5 \times 10^3 \times 4.18 \times 10.0$
$= 62\,700\ \text{J}$
$= 63\ \text{kJ}$

d $q = 2300 \times 4.18 \times 82.0$
$= 788\,348\ \text{J}$
$= 788\ \text{kJ}$

e $q = 300 \times 2.2 \times 82.0$
$= 54\ 120\ \text{J}$
$= 54\ \text{kJ}$

12 $q = m \times c \times \Delta T$. Therefore,

$$\Delta T = \frac{q}{m \times c}$$
$$= \frac{10\,000}{4.18 \times 200}$$
$$= 12.0\text{°C}$$

The increase in temperature was 12.0°C, so the temperature reached by the beaker was 21.0 + 12.0 = 33.0°C

13 $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H = -286\ \text{kJ mol}^{-1}$

or

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$ $\Delta H = -572\ \text{kJ mol}^{-1}$

14 **a** Na^+/CO_3^{2-}
b Ca^{2+}/NO_3^-
c K^+/Br^-
d Fe^{3+}/SO_4^{2-}
e Cu^{2+}/Cl^-

15 dissociation

16 **a** $MgSO_4(s) \xrightarrow{H_2O(l)} Mg^{2+}(aq) + SO_4^{2-}(aq)$
b $Na_2S(s) \xrightarrow{H_2O(l)} 2Na^+(aq) + S^{2-}(aq)$
c $KOH(s) \xrightarrow{H_2O(l)} K^+(aq) + OH^-(aq)$
d $Cu(CH_3COO)(s) \xrightarrow{H_2O(l)} Cu^{2+}(aq) + 2CH_3COO^-(aq)$
e $Li_2SO_4(s) \xrightarrow{H_2O(l)} 2Li^+(aq) + SO_4^{2-}(aq)$

17 Potassium ions and bromide ions are held in an ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds break when it dissolves in water.

Water, the solvent, has hydrogen bonds between the water molecules. When potassium bromide is added to water, the partially positive hydrogen atoms of the water molecules are attracted to the negative bromide ions, and the partially negative oxygen atoms of the water molecules are attracted to the positive potassium ions. Ion–dipole bonds form between the ions and water molecules, and the surface ions are pulled into solution. Gradually, the ionic lattice dissociates and a solution is formed.

18 $\Delta T = 22.0 - 20.4$

$= 1.6\,\text{K}$

$q = m \times c \times \Delta T$

$= 50.0 \times 4.18 \times 1.6$

$= 334.4\,\text{J}$

$= 0.33\,\text{kJ}$

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

$= \frac{6.00}{58.44}$

$= 0.103\,\text{mol}$

$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$

$= \frac{0.33}{0.103}$

$= +3.3\,\text{kJ mol}^{-1}$

(positive enthalpy of dissolution, due to a decrease in the temperature of the water, so endothermic)

19 $\Delta T = 24.1 - 18.0$

$= 6.1\,\text{K}$

$q = m \times c \times \Delta T$

$= 300.0 \times 4.18 \times 6.1$

$= 7649\,\text{J}$

$= 7.6\,\text{kJ}$

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

$= \frac{7.50}{56.11}$

$= 0.134\,\text{mol}$

$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$

$= \frac{-7.6}{0.134}$

$= -57\,\text{kJ mol}^{-1}$

(negative enthalpy of dissolution, due to an increase in the temperature of the water, so exothermic)

20 $n = \frac{m}{M}$

$= \frac{1.50}{40.00}$

$= 0.0375\,\text{mol}$

$q = \Delta H_{\text{soln}} \times n(\text{solute})$

$= 44.51 \times 0.0375$

$= 1.67\,\text{kJ}$

$= 1.67 \times 10^3\,\text{J}$

$\Delta T = \frac{q}{m \times c}$

$= \frac{1.67 \times 10^3}{100.0 \times 4.18}$

$= 3.99^\circ\text{C}$

Determine the initial temperature of the water.

$T_i = T_f - \Delta T$

$= 28.0 - 3.99$

$= 24.0^\circ\text{C}$

Note: The temperature of the water increases because the dissociation process is exothermic.

21 A. The reactant is not gaseous, so increasing the pressure of oxygen gas would have no effect on the rate of reaction.

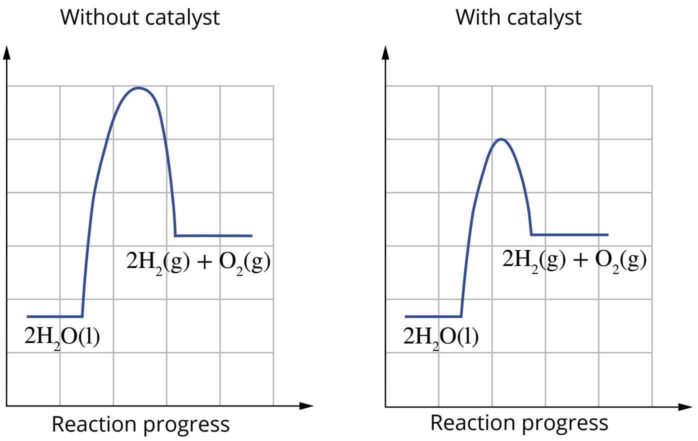
22 **i** Increasing the pressure of the gases would cause an increase in the number of collisions between the reactant molecules in a given time, so more frequent collisions between molecules with the correct orientation for reacting and with energy greater than or equal to the activation energy would occur. As a result, the rate of reaction would increase.

ii Adding a catalyst would allow the reaction to occur by a different pathway with a lower activation energy. The proportion of collisions with energy greater than the activation energy would thus be increased. As a result, the rate of reaction would increase.

23 **a** 1370 kJ mol^{-1}

b −572 kJ mol^{-1}

c



d +572 kJ mol^{-1}

24 **a** Collision frequency is not changed when a catalyst is added. However, the activation energy for the reaction is lowered so that a greater proportion of the collisions occurring involve particles with energy greater than or equal to the activation energy.

b With more particles per unit volume, the frequency of collisions between reactants is increased. The activation energy is not changed.

25 **a** 1 mol of CO(g) and 0.5 mol of $O_2(g)$. As the reaction is exothermic, the enthalpy of the reactants is higher.

b **i** As the coefficients are twice those in the given equation, ΔH will also be doubled.

$\Delta H = 2 \times (-283)$ kJ mol^{-1}

$= -566$ kJ mol^{-1}

ii As the coefficients are twice those in the given equation and the equation is reversed, ΔH will also be doubled, and have the opposite sign. $\Delta H = +566$ kJ mol^{-1}

26 $q = m \times c \times \Delta T$

$= 100.0 \times 4.18 \times 5.52$

$= 2307$ J

$= 2.307$ kJ

$n(C_2H_6) = \frac{m}{M}$

$= \frac{0.0450}{30.07}$

$= 0.00150$ mol

$\Delta H = \frac{q}{n(C_2H_6)}$

$= \frac{-2.307}{0.00150}$

$= -1.54 \times 10^3$ kJ mol^{-1}

Note: The negative sign indicates that the reaction released energy, causing the temperature to rise.

$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$ $\Delta H = -1.54 \times 10^3$ kJ mol^{-1}

or

$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$ $\Delta H = -3.08 \times 10^3$ kJ mol^{-1}

27 **a** $\Delta T = 31.0 - 25.0$

$= 6.0$ K

$q = m \times c \times \Delta T$

$= 100 \times 4.18 \times 6.0$

$= 2508$ J

$= 2.5$ kJ

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

$= \frac{3.00}{56.11}$

$= 0.0535$ mol

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$$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$$
$$= \frac{-2.5}{0.0535}$$
$$= -47\,\text{kJ mol}^{-1}$$

(negative enthalpy of dissolution, due to an increase in the temperature of the water, so exothermic)

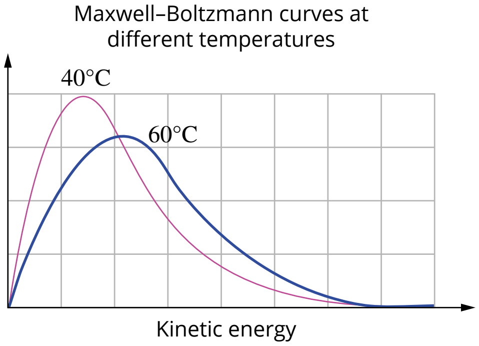
b $$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$$
$$q = \Delta H_{\text{soln}} \times n(\text{solute})$$
$$= 57.61 \times 0.0535$$
$$= 3.08\,\text{kJ}$$

$$\%\ \text{efficiency} = \frac{\text{actual energy released}}{\text{total energy expected to be released}} \times 100\%$$
$$= \frac{2.5}{3.08} \times 100\%$$
$$= 81\%$$

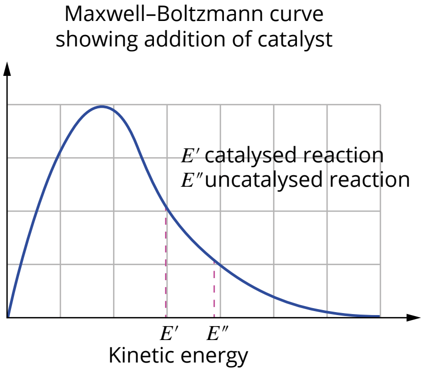
c Efficiency can be improved in any one of the following ways:

- putting a lid on the container (or if a lid is present, ensuring that the lid is tightly fastened, or the seal of the lid is improved)
- improving the insulation of the calorimeter
- ensuring that the solution is continuously stirred so that the heat is evenly distributed throughout the solution.

28 **a**



b



c When a catalyst is present, the reaction proceeds by an alternative reaction pathway with a lower activation energy than in the uncatalysed reaction. At a given temperature, presence of a catalyst means that there will be a greater proportion of the reacting particles that have a kinetic energy equal to or greater than the activation energy. As more reactants have sufficient energy to react, the rate of reaction increases.

29 The dissolution of sodium chloride in water is an endothermic physical reaction. As the salt dissolves into the water, heat energy is absorbed from the milk and sugar mixture.

Chapter 15 Enthalpy and Hess's law

15.1 Latent heat

Worked example: Try yourself 15.1.1

CALCULATING THE ENTHALPY CHANGE OF VAPORISATION

Calculate the amount of heat energy required to transform 5 mol of liquid water at its boiling point to steam.

Thinking	Working
State the amount of energy required to transform 1 mol of liquid water to steam.	$H_2O(l) \rightarrow H_2O(g)$ $\Delta H = +40.7\ \text{kJ mol}^{-1}$
5 mol of water will require 5 times the amount of energy to transform it from liquid to steam as 1 mol.	$5H_2O(l) \rightarrow 5H_2O(g)$ Energy = 5 × +40.7 = 203.5 kJ
Answer the question.	203.5 kJ of heat energy is required to transform 5 mol of liquid water at its boiling point to steam.

Worked example: Try yourself 15.1.2

CALCULATING THE ENTHALPY CHANGE OF MELTING

Calculate the amount of heat energy required to transform 23 g of solid ethanol at its melting point into a liquid.

Thinking	Working
Convert grams to moles.	$n = \frac{m}{M}$ $= \frac{23}{46.07}$ = 0.50 mol
Write the thermochemical equation for fusion for 1 mol of the substance. (Use the data in Student Book Table 15.1.1.)	$C_2H_5OH\ (s) \rightarrow C_2H_5OH\ (l)$ $\Delta H = +5.02\ \text{kJ mol}^{-1}$
Multiply this by the number of mol of ethanol.	$\frac{1}{2}C_2H_5OH\ (s) \rightarrow \frac{1}{2}C_2H_5OH\ (l)$ Energy = 0.50 × +5.02 = +2.5 kJ
Answer the question.	2.5 kJ of heat is required to transform 23 g of solid ethanol at its melting point to liquid.

15.1 KEY QUESTIONS

1. D, A, C, E, B. Solids have the least kinetic energy, while gases have the most kinetic energy.
2. The ΔH_{vap} is an indication of the intermolecular bond strength, because intermolecular bonds that hold the substance in the liquid state are broken when a substance vaporises.

   The ΔH_{vap} of water is $+40.7\ \text{kJ mol}^{-1}$, whereas that of ethanol is $+38.6\ \text{kJ mol}^{-1}$. This suggests that the intermolecular bonds in water are stronger than those in ethanol.
3. The process is endothermic, because energy is added to the liquid to vaporise it.

4 $C_2H_5OH(l) \rightarrow C_2H_5OH(g)$ $\Delta H = +38.6\,kJ\,mol^{-1}$
$12C_2H_5OH(l) \rightarrow 12C_2H_5OH(g)$ Energy $= 12 \times +38.6 = 463\,kJ$

5 $C_6H_6(g) \rightarrow C_6H_6(l)$ $\Delta H = -30.8\,kJ\,mol^{-1}$

6 $n = \frac{m}{M}$
$= \frac{100}{18.02}$
$= 5.55\,mol$
$H_2O(s) \rightarrow H_2O(l)$ $\Delta H = +6.01\,kJ\,mol^{-1}$
$5.55H_2O(s) \rightarrow 5.55H_2O(l)$ Energy $= 5.55 \times +6.01$
$= 33.4\,kJ$

7 $n = \frac{m}{M}$
$= \frac{1000}{18.02}$
$= 55.49\,mol$
$H_2O(g) \rightarrow H_2O(l)$ $\Delta H = -40.7\,kJ\,mol^{-1}$
$55.49H_2O(g) \rightarrow 55.49\ H_2O(l)$ Energy $= 55.49 \times 40.7$
$= 2.26 \times 10^{-3}\,kJ$

15.2 Bond energy

Worked example: Try yourself 15.2.1

USING AVERAGE BOND ENERGIES TO CALCULATE THE ENTHALPY CHANGE FOR A REACTION

Use data from Table 15.2.1 (page 483 of the Student Book) to estimate the enthalpy change for the following reaction: $H_2(g) + F_2(g) \rightarrow 2HF(g)$

<table>
<tr><th>Thinking</th><th>Working</th></tr>
<tr><td>Identify the number and type of bonds broken and made.</td><td>H–H + F–F → H–F + H–F
<table>
<tr><th>Bonds broken</th><th>Bonds formed</th></tr>
<tr><td>1 × F–F
1 × H–H</td><td>2 × H–F</td></tr>
</table></td></tr>
<tr><td>Sum the values for the bond energies of the bonds broken and bonds formed. Pay special attention to the sign.</td><td>
<table>
<tr><th></th><th></th><th>Average bond energy (kJ mol^{–1})</th></tr>
<tr><td>bonds broken</td><td>1 × F–F
1 × H–H</td><td>+155
+436</td></tr>
<tr><td>bonds formed</td><td>2 × H–F</td><td>2 × –567</td></tr>
<tr><td>enthalpy change</td><td></td><td>–543</td></tr>
</table></td></tr>
<tr><td>State the answer.</td><td>$H_2(g) + F_2(g) \rightarrow 2HF(g)\ \Delta H = -543\,kJ\,mol^{-1}$</td></tr>
</table>

15.2 KEY QUESTIONS

1. **a** (ii)
   b (i)
   c (i)

2

Name	Formula	Number and type of bonds
hydrogen	H_2	1 × H–H
water	H_2O	2 × O–H
methane	CH_4	4 × C–H
oxygen	O_2	1 × O=O
ammonia	NH_3	3 × N–H
carbon dioxide	CO_2	2 × C=O

3 Iodine is a larger atom than chlorine. The valence (bonding) electrons are a greater distance from the nucleus. Consequently, the electrons which form the covalent bond are less tightly held by the nuclei. Therefore, the bond energy of iodine is lower.

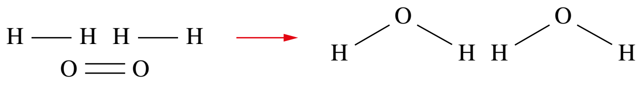
4 **a** $2HF(g) \rightarrow H_2(g) + F_2(g)$

H–F + H–F → H–H + F–F

		Average bond energy (kJ mol⁻¹)
bonds broken	2 × H–F	2 × +567
bonds formed	2 × H–H 1 × F–F	–436 1 × –155
enthalpy change		+543

$2HF(g) \rightarrow H_2(g) + F_2(g) \quad \Delta H = +543\ kJ\ mol^{-1}$

b $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

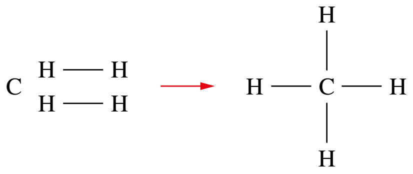


		Average bond energy (kJ mol⁻¹)
bonds broken	2 × H–H 1 × O=O	2 × +436 +498
bonds formed	4 × O–H	4 × –463
enthalpy change		–482

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g) \quad \Delta H = -482\ kJ\ mol^{-1}$

c $C(g) + 2H_2(g) \rightarrow CH_4(g)$

Note: C(g) is already atomised (it is not bonded).



		Average bond energy (kJ mol⁻¹)
bonds broken	2 × H–H	2 × +436
bonds formed	4 × C–H	4 × –414
enthalpy change		–784

$C(g) + 2H_2(g) \rightarrow CH_4(g) \quad \Delta H = -784\ kJ\ mol^{-1}$

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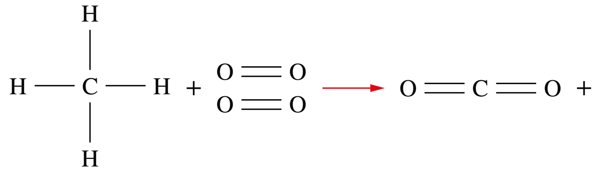
d $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$



		Average bond energy (kJ mol^{-1})
bonds broken	2 × O=O 4 × C–H	2 × +498 4 × +414
bonds formed	2 × C=O 4 × O–H	2 × –804 4 × –463
enthalpy change		–808 kJ

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g) \qquad \Delta H = -808\ \text{kJ mol}^{-1}$

15.3 Hess's law

Worked example: Try yourself 15.3.1

CALCULATING THE ENTHALPY CHANGE FOR A REACTION

Calculate the enthalpy change for the following reaction:

$CH_3COOH(l) + H_2(g) \rightarrow CH_3CH_2OH(l) + \frac{1}{2}O_2(g)$

given the following enthalpies of combustion:

$$CH_3CH_2OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \qquad \Delta H = -1360\ \text{kJ mol}^{-1} \qquad (1)$$
$$CH_3COOH(l) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \qquad \Delta H = -876\ \text{kJ mol}^{-1} \qquad (2)$$
$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta H = -285.8\ \text{kJ mol}^{-1} \qquad (3)$$

Thinking	**Working**
Rewrite equations to make the products for the reaction. If necessary, reverse the sign of the ΔH values.	Reverse (1): $2CO_2(g) + 3H_2O(l) \rightarrow CH_3CH_2OH(l) + 3O_2(g) \quad \Delta H = +1360\ \text{kJ mol}^{-1}$
Rewrite equations to make reactants for the reaction. If necessary, reverse the sign of the ΔH values.	Equation (2) $CH_3COOH(l) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \quad \Delta H = -876\ \text{kJ mol}^{-1}$ Equation (3) $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H = -285.8\ \text{kJ mol}^{-1}$
Make sure the molar quantities are correct. Apply a multiplier to each equation if necessary.	No multiplier is required for any of the equations.
Add the equations and their ΔH values. Cancel reactants and products that occur on both sides of the equation.	$\cancel{2CO_2(g)} + \cancel{3H_2O(l)} \rightarrow CH_3CH_2OH(l) + \cancel{3O_2(g)} \quad \Delta H = +1360\ \text{kJ mol}^{-1}$ $CH_3COOH(l) + \cancel{2O_2(g)} \rightarrow \cancel{2CO_2(g)} + \cancel{2H_2O(l)} \quad \Delta H = -876\ \text{kJ mol}^{-1}$ $H_2(g) + \cancel{\frac{1}{2}O_2(g)} \rightarrow \cancel{H_2O(l)} \quad \Delta H = -285.8\ \text{kJ mol}^{-1}$ $CH_3COOH(l) + H_2(g) \rightarrow CH_3CH_2OH(l) + \frac{1}{2}O_2(g) \quad \Delta H = +198\ \text{kJ mol}^{-1}$
State the answer.	The enthalpy change for $CH_3COOH(l) + H_2(g) \rightarrow CH_3CH_2OH(l) + \frac{1}{2}O_2(g)$ is +198 kJ mol^{-1}.

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Worked example: Try yourself 15.3.2

USING STANDARD ENTHALPY OF FORMATION DATA TO CALCULATE THE ENTHALPY CHANGE FOR A REACTION

Calculate the enthalpy change for the following reaction:

$Cl_2(g) + C_3H_8(g) \rightarrow C_3H_7Cl(g) + HCl(g)$

given the following standard enthalpies of formation:

$$3C(s) + 4H_2(g) \rightarrow C_3H_8(g) \qquad \Delta H^\circ_f = -104.6\,\text{kJ mol}^{-1}$$

$$\tfrac{1}{2}H_2(g) + \tfrac{1}{2}Cl_2(g) \rightarrow HCl(g) \qquad \Delta H^\circ_f = -92.3\,\text{kJ mol}^{-1}$$

$$3C(s) + \tfrac{7}{2}H_2(g) + \tfrac{1}{2}Cl_2(g) \rightarrow C_3H_7Cl(g) \qquad \Delta H^\circ_f = -132.5\,\text{kJ mol}^{-1}$$

Thinking	**Working**
Calculate the sum of the enthalpies of formation of the products, $\Sigma\Delta H^\circ_f$(products). As ΔH°_f is given per mol, multiply the ΔH°_f of each product by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{products}) = 1 \times \Delta H^\circ_f(C_3H_7Cl) + 1 \times \Delta H^\circ_f(HCl)$ $= (1 \times -132.5) + (1 \times -92.3)$ $= -224.8\,\text{kJ mol}^{-1}$
Calculate the sum of the enthalpies of formation of the reactants, $\Sigma\Delta H^\circ_f$(reactants). As ΔH°_f is given per mol, multiply the ΔH°_f of each reactant by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{reactants}) = 1 \times \Delta H^\circ_f(C_3H_8) + 1 \times \Delta H^\circ_f(Cl_2)$ $= (1 \times -104.6) + (1 \times 0)$ $= -104.6\,\text{kJ mol}^{-1}$
Calculate $\Delta H^\circ_{\text{reaction}}$ using the expression: $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$	$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$ $= -224.8 - (-104.6)$ $= -120.2\,\text{kJ mol}^{-1}$

Worked example: Try yourself 15.3.3

USING STANDARD ENTHALPY OF FORMATION DATA TO CALCULATE ENTHALPY OF COMBUSTION

Calculate the enthalpy change for the complete combustion of butan-1-ol:

$C_4H_9OH(l) + 6O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

given the following standard enthalpies of formation:

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^\circ_f = -393.5\,\text{kJ mol}^{-1}$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta H^\circ_f = -285.8\,\text{kJ mol}^{-1}$$

$$4C(s) + \tfrac{1}{2}O_2(g) + 5H_2(g) \rightarrow C_4H_9OH(l) \qquad \Delta H^\circ_f = -325.8\,\text{kJ mol}^{-1}$$

Thinking	**Working**
Calculate the sum of the enthalpies of formation of the products, $\Sigma\Delta H^\circ_f$(products). As ΔH°_f is given per mol, multiply the ΔH°_f of each product by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{products}) = 4 \times \Delta H^\circ_f(CO_2) + 5 \times \Delta H^\circ_f(H_2O)$ $= (4 \times -393.5) + (5 \times -285.8)$ $= -3003.3\,\text{kJ mol}^{-1}$
Calculate the sum of the enthalpies of formation of the reactants, $\Sigma\Delta H^\circ_f$(reactants). As ΔH°_f is given per mol, multiply the ΔH°_f of each reactant by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{reactants}) = 1 \times \Delta H^\circ_f(C_4H_9OH) + 6 \times \Delta H^\circ_f(O_2)$ $= (1 \times -325.8) + (6 \times 0)$ $= -325.8\,\text{kJ mol}^{-1}$
Calculate $\Delta H^\circ_{\text{reaction}}$ using the expression: $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$	$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$ $= -3003.3 - (-325.8)$ $= -2677.2\,\text{kJ mol}^{-1}$

Worked example: Try yourself 15.3.4

CALCULATING THE ENTHALPY CHANGE FOR ANAEROBIC RESPIRATION

Fermentation is a form of anaerobic respiration performed by yeasts. Fermentation of glucose to alcohol and carbon dioxide by yeasts is crucially important in the bread making and alcohol industries.

Calculate the enthalpy change associated with the fermentation of glucose:

$C_6H_{12}O_6(s) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$

given the following standard enthalpies of formation:

$6C(s) + 6H_2(g) + 3O_2(g) \rightarrow C_6H_{12}O_6(s)$ $\Delta H^\circ_f = -1271\ kJ\ mol^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ_f = -393.5\ kJ\ mol^{-1}$

$2C(s) + \frac{1}{2}O_2(g) + 3H_2(g) \rightarrow C_2H_5OH(l)$ $\Delta H^\circ_f = -277.7\ kJ\ mol^{-1}$

Thinking	**Working**
Calculate the sum of the enthalpies of formation of the products, $\Sigma\Delta H^\circ_f$(products). As ΔH°_f is given per mol, multiply the ΔH°_f of each product by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{products}) = 2 \times \Delta H^\circ_f(C_2H_5OH) + 2 \times \Delta H^\circ_f(CO_2)$ $= (2 \times -277.7) + (2 \times -393.5)$ $= -1342.4\ kJ\ mol^{-1}$
Calculate the sum of the enthalpies of formation of the reactants, $\Sigma\Delta H^\circ_f$(reactants). As ΔH°_f is given per mol, multiply the ΔH°_f of each reactant by its coefficient in the overall equation.	$\Sigma\Delta H^\circ_f(\text{reactants}) = 1 \times \Delta H^\circ_f(C_6H_{12}O_6)$ $= 1 \times -1271$ $= -1271\ kJ\ mol^{-1}$
Calculate $\Delta H^\circ_{reaction}$ using the expression: $\Delta H^\circ_{reaction} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$	$\Delta H^\circ_{reaction} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$ $= -1342.4 - (-1271)$ $= -71\ kJ\ mol^{-1}$

15.3 KEY QUESTIONS

1 Step 1. Reverse equation (1) to make NO(g) a product.

$N_2(g) + O_2(g) \rightarrow 2NO(g)$ $\Delta H^\circ = +180.6\ kJ\ mol^{-1}$

Step 2. Rewrite equation (2) to make NOCl(g) a reactant and $Cl_2(g)$ a product.

$2NOCl(g) \rightarrow N_2(g) + O_2(g) + Cl_2(g)$ $\Delta H^\circ = -103.4\ kJ\ mol^{-1}$

No multipliers are necessary.

Step 3. Add equations and cancel species that occur on both sides of the equation:

$\cancel{N_2(g)} + \cancel{O_2(g)} \rightarrow 2NO(g)$ $\Delta H^\circ = +180.6\ kJ\ mol^{-1}$

$2NOCl(g) \rightarrow \cancel{N_2(g)} + \cancel{O_2(g)} + Cl_2(g)$ $\Delta H^\circ = -103.4\ kJ\ mol^{-1}$

$2NOCl(g) \rightarrow 2NO(g) + Cl_2(g)$ $\Delta H^\circ = +77.2\ kJ\ mol^{-1}$

2 Step 1. Reverse equation (1) to make $H_2SO_4(l)$ a reactant.

$H_2SO_4(l) \rightarrow H_2S(g) + 2O_2(g)$ $\Delta H^\circ = +235.5\ kJ\ mol^{-1}$

Step 2. Use equation (2) as is.

$H_2S(g) + 2O_2(g) \rightarrow SO_3(g) + H_2O(l)$ $\Delta H^\circ = -207\ kJ\ mol^{-1}$

Step 3. Use equation (3) as is.

$H_2O(l) \rightarrow H_2O(g)$ $\Delta H^\circ = +44\ kJ\ mol^{-1}$

Step 4. Note: No multipliers necessary. Add equations and cancel species that occur on both sides of the equation:

$H_2SO_4(l) \rightarrow \cancel{H_2S(g)} + \cancel{2O_2(g)}$ $\Delta H^\circ = +235.5\ kJ\ mol^{-1}$

$\cancel{H_2S(g)} + \cancel{2O_2(g)} \rightarrow SO_3(g) + \cancel{H_2O(l)}$ $\Delta H^\circ = -207\ kJ\ mol^{-1}$

$\cancel{H_2O(l)} \rightarrow H_2O(g)$ $\Delta H^\circ = +44\ kJ\ mol^{-1}$

$H_2SO_4(l) \rightarrow SO_3(g) + H_2O(g)$ $\Delta H^\circ = +73\ kJ\ mol^{-1}$

3 Step 1. Use equation (1) as is:

$$\mathrm{CaCO_3(s) \rightarrow CaO(s) + CO_2(g)} \qquad \Delta H^\circ = +175\ \mathrm{kJ\ mol^{-1}}$$

Step 2. Rewrite equation (2) to make $Ca(OH)_2(s)$, an intermediate product.

$$\mathrm{H_2O(l) + CaO(s) \rightarrow Ca(OH)_2(s)} \qquad \Delta H^\circ = -67\ \mathrm{kJ\ mol^{-1}}$$

Step 3. Use equation (3) as is.

$$\mathrm{Ca(OH)_2(s) + 2HCl(g) \rightarrow CaCl_2(s) + 2H_2O(l)} \qquad \Delta H^\circ = -198\ \mathrm{kJ\ mol^{-1}}$$

Step 4. Note: No multipliers necessary. Add equations and cancel species that occur on both sides of the equation:

$$\begin{array}{ll}
\mathrm{CaCO_3(s) \rightarrow \cancel{CaO(s)} + CO_2(g)} & \Delta H^\circ = +175\ \mathrm{kJ\ mol^{-1}} \\
\mathrm{\cancel{H_2O(l)} + \cancel{CaO(s)} \rightarrow \cancel{Ca(OH)_2(s)}} & \Delta H^\circ = -67\ \mathrm{kJ\ mol^{-1}} \\
\mathrm{\cancel{Ca(OH)_2(s)} + 2HCl(g) \rightarrow CaCl_2(s) + \cancel{2}H_2O(l)} & \Delta H^\circ = -198\ \mathrm{kJ\ mol^{-1}} \\
\hline
\mathrm{CaCO_3(s) + 2HCl(g) \rightarrow CaCl_2(s) + H_2O(l) + CO_2(g)} & \Delta H^\circ = -90\ \mathrm{kJ\ mol^{-1}}
\end{array}$$

4

$$\begin{array}{ll}
\mathrm{C(s) + 2H_2(g) \rightarrow CH_4(g)} & \Delta H^\circ_f = -74.8\ \mathrm{kJ\ mol^{-1}} \\
\mathrm{C(s) + O_2(g) \rightarrow CO_2(g)} & \Delta H^\circ_f = -393.5\ \mathrm{kJ\ mol^{-1}} \\
\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)} & \Delta H^\circ_f = -285.8\ \mathrm{kJ\ mol^{-1}}
\end{array}$$

$$\begin{aligned}
\Sigma\Delta H^\circ_f(\text{products}) &= 1 \times \Delta H^\circ_f(\mathrm{CO_2}) + 2 \times \Delta H^\circ_f(\mathrm{H_2O}) \\
&= (1 \times -393.5) + (2 \times -285.8) \\
&= -965.1\ \mathrm{kJ\ mol^{-1}} \\
\Sigma\Delta H^\circ_f(\text{reactants}) &= 1 \times \Delta H^\circ_f(\mathrm{CH_4}) + 2 \times \Delta H^\circ_f(\mathrm{O_2}) \\
&= (1 \times -74.8) + (2 \times 0) \\
&= -74.8\ \mathrm{kJ\ mol^{-1}} \\
\Delta H^\circ_{\text{reaction}} &= \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants}) \\
&= -965.1 - (-74.8) \\
&= -890.3\ \mathrm{kJ\ mol^{-1}}
\end{aligned}$$

5

$$\begin{array}{ll}
\mathrm{N_2(g) + 2H_2(g) \rightarrow N_2H_4(l)} & \Delta H^\circ_f = +50.6\ \mathrm{kJ\ mol^{-1}} \\
\mathrm{H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(g)} & \Delta H^\circ_f = -241.8\ \mathrm{kJ\ mol^{-1}}
\end{array}$$

$$\begin{aligned}
\Sigma\Delta H^\circ_f(\text{products}) &= 1 \times \Delta H^\circ_f(\mathrm{N_2}) + 2 \times \Delta H^\circ_f(\mathrm{H_2O}) \\
&= (1 \times 0) + (2 \times -241.8) \\
&= -483.6\ \mathrm{kJ\ mol^{-1}} \\
\Sigma\Delta H^\circ_f(\text{reactants}) &= 1 \times \Delta H^\circ_f(\mathrm{N_2H_4}) + 1 \times \Delta H^\circ_f(\mathrm{O_2}) \\
&= (1 \times +50.6) + (1 \times 0) \\
&= +50.6\ \mathrm{kJ\ mol^{-1}} \\
\Delta H^\circ_{\text{reaction}} &= \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants}) \\
&= -483.6 - (+50.6) \\
&= -534.2\ \mathrm{kJ\ mol^{-1}}
\end{aligned}$$

6

$$\begin{aligned}
\Sigma\Delta H^\circ_f(\text{products}) &= 1 \times \Delta H^\circ_f(\mathrm{C_2F_4}) + 4 \times \Delta H^\circ_f(\mathrm{HF}) \\
&= (1 \times -675.0) + (4 \times -273.3) \\
&= -1768.2\ \mathrm{kJ\ mol^{-1}} \\
\Sigma\Delta H^\circ_f(\text{reactants}) &= 1 \times \Delta H^\circ_f(\mathrm{C_2H_4}) + 4 \times \Delta H^\circ_f(\mathrm{F_2}) \\
&= (1 \times +52.0) + (4 \times 0) \\
&= +52.0\ \mathrm{kJ\ mol^{-1}} \\
\Delta H^\circ_{\text{reaction}} &= \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants}) \\
&= -1768.2 - (+52.0) \\
&= -1820.2\ \mathrm{kJ\ mol^{-1}}
\end{aligned}$$

7

$$\begin{aligned}
\Sigma\Delta H^\circ_f(\text{products}) &= \left(\tfrac{-270.3}{2}\right) + 0 = +135.15\ \mathrm{kJ\ mol^{-1}} \\
\Sigma\Delta H^\circ_f(\text{reactants}) &= -74.8 + \left(\tfrac{-91.8}{2}\right) = -120.7\ \mathrm{kJ\ mol^{-1}} \\
\Delta H^\circ_{\text{reaction}} &= \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants}) \\
&= +135.15 - (-120.7) \\
&= +255.9\ \mathrm{kJ\ mol^{-1}}
\end{aligned}$$

8 **a** $C_2H_2(g) + \frac{5}{2}O_2(g) \rightarrow 2CO_2(g) + H_2O(l)$

b $\Sigma\Delta H^\circ_f(\text{products}) = 2 \times \Delta H^\circ_f(CO_2) + 1 \times \Delta H^\circ_f(H_2O)$

$= (2 \times -393.5) + (1 \times -285.8)$

$= -1072.8\,\text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = 1 \times \Delta H^\circ_f(C_2H_2) + \frac{5}{2} \times \Delta H^\circ_f(O_2)$

$= (1 \times +228) + (\frac{5}{2} \times 0)$

$= +228\,\text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -1072.8 - (+228)$

$= -1300.8\,\text{kJ mol}^{-1}$

c Calculate the number of mol of ethyne.

$n = \frac{m}{M} = \frac{100.0}{26.04} = 3.841\,\text{mol}$

100 g of ethyne produces 3.841 × 1300.8 = 4995 kJ of energy.

CHAPTER 15 **REVIEW**

1 **a** iv $C(s) + O_2(g) \rightarrow CO_2(g)$ $\Delta H^\circ = -393.5\,\text{kJ mol}^{-1}$

b ii $H_2O(l) \rightarrow H_2O(g)$ $\Delta H^\circ = +40.7\,\text{kJ mol}^{-1}$

2 **a** $10H_2O(l) \rightarrow 10H_2O(g)$ Energy = $10 \times 40.7 = 4.1 \times 10^2$ kJ required

b Calculate mol: $n = \frac{m}{M}$

$= \frac{1000}{18.02}$

$= 55.49$

$55.49H_2O(l) \rightarrow 55.49H_2O(g)$ Energy = $55.49 \times 40.7 = 2.26 \times 10^3$ kJ required

3 **a** A positive value of ΔH_{vap} indicates an endothermic process, i.e. vaporisation requires energy.

b Intermolecular bonds are broken during vaporisation of water. Metallic bonds are broken during vaporisation of mercury.

c The metallic bonding in mercury is stronger than the hydrogen bonds between water molecules in water, so more heat energy is required to vaporise mercury.

4 **a** endothermic, as indicated by the positive value of ΔH_{vap}

b $30.8 \times 3 = 92.4$ kJ

5 The greater value for the ΔH_{fus} for NaCl indicates stronger bonding between the particles in solid NaCl than in solid water. In fact, the ions of NaCl(s) are held together by strong ionic bonds, whereas water molecules are held together by hydrogen bonding.

6 The particles of steam have more kinetic energy than those of water.

7 **a** $C(g) + O_2(g) \rightarrow CO_2(g)$

		Average bond energy (kJ mol^{-1})
bonds broken	O=O	+498
bonds formed	2 × C=O	2 × –804
enthalpy change		–1110

b $H_2(g) + F_2(g) \rightarrow 2HF(g)$

		Average bond energy (kJ mol^{-1})
bonds broken	H–H F–F	+436 +159
bonds formed	2 × H–F	2 × –567
enthalpy change		–539

c $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

		Average bond energy (kJ mol^{-1})
bonds broken	N≡N 3 × H–H	+945 3 × +436
bonds formed	6 × N–H	6 × –391
enthalpy change		–93

8 **a** $C_{15}H_{32}(l) + 23O_2(g) \rightarrow 15CO_2(g) + 16H_2O(g)$

b

		Average bond energy (kJ mol^{-1})
bonds broken	14 × C–C 32 × C–H 23 × O=O	14 × +346 32 × +414 23 × +498
bonds formed	30 × C=O 32 × O–H	30 × –804 32 × –463
enthalpy change		–9390

9 **a** $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$

b

		Average bond energy (kJ mol^{-1})
bonds broken	C–C 6 × C–H 3.5 × O=O	+346 6 × +414 3.5 × +498
bonds formed	4 × C=O 6 × O–H	4 × –804 6 × –463
enthalpy change		–1421

10

		Average bond energy (kJ mol^{-1})
bonds broken	4 × C–H 1 × Cl–Cl	4 × +414 +242
bonds formed	3 × C–H 1 × C–Cl 1 × H–Cl	3 × –414 –324 –431
enthalpy change		–499

11

		Average bond energy (kJ mol^{-1})
bonds broken	6 × N–H 2 × C=O	6 × +391 2 × +804
bonds formed	2 × O–H 4 × N–H 2 × N–C 1 × C=O	2 × –463 4 × –391 2 × –286 –804
enthalpy change		+88

12 One reason is that H–H is a much smaller molecule than F–F. Hydrogen's atomic radius is smaller, so the shared valence electrons in the covalent bond are more strongly attracted to the hydrogen nuclei. Second, the fluorine molecule has many more electron pairs, so there is a significant amount of coulombic (electrostatic) repulsion that reduces its bond strength.

13 A straight-line relationship is apparent in the graph because with each successive alcohol, an extra C–C bond, 2 C–H bonds and $\frac{3}{2}$ O=O bonds are broken, while an extra 2 C=O bonds and 2 O–H bonds are formed.

14 Reverse equation (2) and change the sign.

$$Fe_2O_3(s) \rightarrow \quad 2Fe(s) + \tfrac{3}{2}O_2(g) \qquad \Delta H^\circ = +824\,kJ\,mol^{-1}$$

Use equation (1) as is.

$$2Al(s) + \tfrac{3}{2}O_2(g) \rightarrow \quad Al_2O_3(s) \qquad \Delta H^\circ = -1676\,kJ\,mol^{-1}$$

Add the equations and their ΔH° values.

Cancel reactants and products that occur on both sides of the equation.

$$\begin{array}{rll}
Fe_2O_3(s) \rightarrow & 2Fe(s) + \cancel{\tfrac{3}{2}O_2(g)} & \Delta H^\circ = +824\,kJ\,mol^{-1} \\
2Al(s) + \cancel{\tfrac{3}{2}O_2(g)} \rightarrow & Al_2O_3(s) & \Delta H^\circ = -1676\,kJ\,mol^{-1} \\
\hline
Fe_2O_3(s) + 2Al(s) \rightarrow & Al_2O_3(s) + 2Fe(s) & \Delta H^\circ = -852\,kJ\,mol^{-1}
\end{array}$$

15 Reverse equation (2).

$$CO(g) \rightarrow \quad C\,(s) + \tfrac{1}{2}O_2(g) \qquad \Delta H^\circ = +111\,kJ\,mol^{-1}$$

Use equation (1) as is.

$$C(s) + O_2(g) \rightarrow \quad CO_2(g) \qquad \Delta H^\circ = -394\,kJ\,mol^{-1}$$

Add the equations and their ΔH° values.

Cancel reactants and products that occur on both sides of the equation.

$$\begin{array}{rll}
CO(g) \rightarrow & \cancel{C\,(s)} + \cancel{\tfrac{1}{2}O_2(g)} & \Delta H^\circ = +111\,kJ\,mol^{-1} \\
\cancel{C(s)} + O_2(g) \rightarrow & CO_2(g) & \Delta H^\circ = -394\,kJ\,mol^{-1} \\
\hline
CO(g) + \tfrac{1}{2}O_2(g) \rightarrow & CO_2(g) & \Delta H^\circ = -283\,kJ\,mol^{-1}
\end{array}$$

16 Use equation (1) as is.

$$C(s) + H_2O(g) \rightarrow \quad CO(g) + H_2(g) \qquad \Delta H^\circ = +131\,kJ\,mol^{-1}$$

Rewrite equation (2) to make $CO_2(g)$ and $H_2(g)$ products

$$CO(g) + H_2O(g) \rightarrow \quad CO_2(g) + H_2(g) \qquad \Delta H^\circ = -40\,kJ\,mol^{-1}$$

Add the equations and their ΔH° values.

Cancel reactants and products that occur on both sides of the equation.

$$\begin{array}{rll}
C(s) + H_2O(g) \rightarrow & \cancel{CO(g)} + H_2(g) & \Delta H^\circ = +131\,kJ\,mol^{-1} \\
\cancel{CO(g)} + H_2O(g) \rightarrow & CO_2(g) + H_2(g) & \Delta H^\circ = -40\,kJ\,mol^{-1} \\
\hline
C(s) + 2H_2O(g) \rightarrow & CO_2(g) + 2H_2(g) & \Delta H^\circ = +91\,kJ\,mol^{-1}
\end{array}$$

17 $\Delta H_3 = \Delta H_1 + (-\Delta H_2) = +180 + (-112) = +68\,kJ$

18 $\Delta H^\circ_{reaction} = \Sigma\Delta H^\circ_f(products) - \Sigma\Delta H^\circ_f(reactants)$

$\Sigma\Delta H^\circ_f(products) = (-393.5) + (2 \times -296.8) = -987.1\,kJ\,mol^{-1}$

$\Sigma\Delta H^\circ_f(reactants) = (+87.9) + (0) = +87.9\,kJ\,mol^{-1}$

$$\begin{aligned}
\Delta H^\circ_{reaction} &= \Sigma\Delta H^\circ_f(products) - \Sigma\Delta H^\circ_f(reactants) \\
&= -987.1 - (+87.9) \\
&= -1075.0\,kJ\,mol^{-1}
\end{aligned}$$

19 Helium and neon are elements, so their enthalpies of formation are zero.

20

$$\begin{array}{rll}
S(s) + \tfrac{3}{2}O_2(g) \rightarrow & \cancel{SO_3(g)} & \Delta H^\circ_f = -295\,kJ\,mol^{-1} \\
\cancel{SO_3(g)} \rightarrow & SO_2(g) + \cancel{\tfrac{1}{2}O_2(g)} & \Delta H^\circ_f = -98\,kJ\,mol^{-1} \\
\hline
S(s) + O_2(g) \rightarrow & SO_2(g) & \Delta H^\circ_f = -393\,kJ\,mol^{-1}
\end{array}$$

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21 **a** $C_2H_4(g) + Cl_2(g) \rightarrow C_2H_3Cl(g) + HCl(g)$

b $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = (+22.0) + (-92.3) = -70.3\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (+52.3) + (0) = +52.3\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -70.3 - (+52.3)$

$= -122.6\ \text{kJ mol}^{-1}$

22 $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = -277.7\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (+52.3) + (-241.8) = -189.5\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -277.7 - (-189.5)$

$= -88.2\ \text{kJ mol}^{-1}$

23 $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = (-110.5) + (0) = -110.5\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (-74.8) + (-241.8) = -316.6\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -110.5 - (-316.6)$

$= +206.1\ \text{kJ mol}^{-1}$

24 **a** $CH_4(g) + O_2(g) \rightarrow C(s) + 2H_2O(l)$

b $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = (0) + (2 \times -258.5) = -517.0\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (-74.8) + (0) = -74.8\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -517 - (-74.8)$

$= -442.2\ \text{kJ mol}^{-1}$

25 $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = (0) + (3 \times -241.8) = -725.4\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (-824.2) + (0) = -824.2\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -725.4 - (-824.2)$

$= +98.8\ \text{kJ mol}^{-1}$

26 **a** $\Delta H_2 = \Delta H^\circ_f(CH_4) = -74.8\ \text{kJ mol}^{-1}$

b $\Delta H_3 = \Delta H$ for sublimation of carbon + 2 × H–H bond energy = $(+717) + (2 \times 436) = +1589\ \text{kJ mol}^{-1}$

c $\Delta H_1 = -\Delta H_2 + \Delta H_3 = -(-74.8) + (+1589) = +1664\ \text{kJ mol}^{-1}$

27 $\Delta H_1 = \Delta H_2 + \Delta H_3$

$\Delta H_3 = \Delta H_1 - \Delta H_2$

$= -890 - (-802)$

$= -88\ \text{kJ mol}^{-1}$

28 **a** $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

b $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{products}) = (4 \times -393.5) + (5 \times -285.8) = -3003.0\ \text{kJ mol}^{-1}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (-126) + (0) = -125.6\ \text{kJ mol}^{-1}$

$\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$= -3003.0 - (-126)$

$= -2877\ \text{kJ mol}^{-1}$

29 Melting requires energy—the enthalpy of fusion. For melting ice to liquid water, this is 6.01 kJ mol^{-1}. Heat energy is absorbed to cause a change of state instead of a change in temperature.

Chapter 16 Entropy and Gibbs free energy

16.1 Energy and entropy changes in chemical reactions

Worked example: Try yourself 16.1.1

DETERMINING THE ENTROPY CHANGE FOR A CHEMICAL REACTION

Determine the change in entropy for the following chemical reaction:

$$(NH_4)_2CO_3(s) \rightarrow 2NH_3(g) + CO_2(g) + H_2O(l)$$

Thinking	Working
Determine whether there is a change in the number of particles in the system and, if so, the effect on entropy.	There is 1 mol of reactant particles and there are 4 mol of product particles. The system becomes more disordered, so ΔS is positive.
Determine whether there is a change of state in the system and, if so, the effect on entropy.	There is an increase in the number of moles of liquid and gas during the reaction. The system becomes more disordered, so ΔS is positive.
Determine the overall change in entropy for the system.	Overall, the system becomes more disordered, so ΔS is positive.

16.1 KEY QUESTIONS

1 The student is adding two liquids together. While the system can be chosen freely, a sensible choice would be to choose the volume of the two liquids as the system. The boundary would be the flask. As the two liquids are mixed on the bench, energy may pass to and from the surroundings. As the flask has a stopper, matter cannot escape from the system; therefore, the system is closed.

2 There is no change in entropy. There is a different arrangement of buttons, but not a difference in the number of possible arrangements.

3 It has decreased. After compression, the gas occupies a smaller volume, so there are fewer possible arrangements of the gas throughout the volume.

4 **a** decreased, as 4 mol of gas become 2 mol of gas
  b decreased, as 1 mol of gas is lost and the number of mol of solid is unchanged
  c increased, as 1 mol of gas is formed and the number of mol of solid is unchanged
  d increased, as going from a liquid to a gas

5 III → VI → IV → I → II → V

16.2 Entropy and spontaneous processes

Worked example: Try yourself 16.2.1

DETERMINING THE STANDARD ENTROPY CHANGE FOR A CHEMICAL REACTION FROM STANDARD ENTROPIES

Determine the entropy change for the following reaction:

$H_2(g) + I_2(s) \rightarrow 2HI(g)$

given the following standard entropies:

$S°(H_2) = +130.6\ J\,mol^{-1}\,K^{-1}$

$S°(I_2) = +116.7\ J\,mol^{-1}\,K^{-1}$

$S°(HI) = +206.3\ J\,mol^{-1}\,K^{-1}$

Thinking	**Working**
Calculate the total entropy of the products, $\Sigma S°$(products). As $S°$ is given per mol, multiply the $S°$ of each product by its coefficient in the overall equation.	$\Sigma S°(\text{products}) = 2 \times S°(HI)$ $= 2 \times +206.3$ $= +412.6\ J\,mol^{-1}\,K^{-1}$
Calculate the total entropy of the reactants, $\Sigma S°$(reactants). As $S°$ is given per mole, multiply the $S°$ of each reactant by its coefficient in the overall equation.	$\Sigma S°(\text{reactants}) = 1 \times S°(H_2) + 1 \times S°(I_2)$ $= (1 \times +130.6) + (1 \times +116.7)$ $= +247.3\ J\,mol^{-1}\,K^{-1}$
Calculate the entropy change of the reaction, $\Delta S°$, using the expression: $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$	$\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$ $= +412.6 - (+247.3)$ $= +165.3\ J\,mol^{-1}\,K^{-1}$

16.2 KEY QUESTIONS

1 $\Delta S° = \Sigma S°\ (\text{products}) - \Sigma S°(\text{reactants})$

$= +257 - (+32 + (+205))$

$= +20\ J\,mol^{-1}\,K^{-1}$

2 **a** There is little change in ΔS_{system} because neither a gas nor a liquid is produced. The thermite reaction is exothermic ($\Delta H < 0$), so $\Delta S_{surrounding} > 0$. $\Delta S_{system} + \Delta S_{surrounding} > 0$, so the reaction is spontaneous.

b The exothermic reaction releases heat to the surroundings ($\Delta H < 0$), so $\Delta S_{surroundings} > 0$. The reaction has a loss of entropy ($\Delta S_{system} < 0$), because the reaction reduces 3 mol of gas to 2 mol of gas. The magnitude of $\Delta S_{surroundings}$ is larger than the magnitude of ΔS_{system}, so the sum of the two values is positive and the reaction is spontaneous.

c The process is endothermic ($\Delta H > 0$), so $\Delta S_{surroundings} < 0$. The increase in entropy when 1 mol of solid is dissolved to produce 2 mol of ions in solution is high enough ($\Delta S_{system} > \Delta S_{surroundings}$), so $\Delta S_{system} + \Delta S_{surroundings} > 0$; thus, the process is spontaneous.

3 **a** The change in entropy for the reaction is:

$\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$

$= (+89.6 + (+130.6)) - (+32.7 + (2 \times +186.9))$

$= 220.2 - 406.5$

$= -186.3\ J\,mol^{-1}\,K^{-1}$

b For this reaction, there is a net decrease in entropy. The reaction requires 1 mol of solid and 2 mol of gas to become 1 mol of solid and 1 mol of gas. The net loss of 1 mol of gas is expected to reduce the entropy of the system. This matches the negative entropy change calculated in part **a**.

4 $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$

$+25 = 2 \times S°(NO) - (+191.6 + 205)$

$2 \times S°(NO) = 25 + 396.6$

$S°(NO) = \frac{+422}{2} = +211\ J\,mol^{-1}\,K^{-1}$

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16.3 Gibbs free energy

Worked example: Try yourself 16.3.1

CALCULATING THE SPONTANEITY OF A REACTION

Ice melting is a change of state in which solid water melts to liquid water.
The reaction is:

$$H_2O(s) \rightarrow H_2O(l)$$

For this reaction, $\Delta H° = +6.01\ \text{kJ mol}^{-1}$ and $\Delta S° = +22.1\ \text{J mol}^{-1}\ \text{K}^{-1}$. Is this reaction spontaneous at –40°C?

Thinking	**Working**
State the known values.	$\Delta H° = +6.01\ \text{kJ mol}^{-1}$ $\Delta S° = +22.1\ \text{J mol}^{-1}\ \text{K}^{-1}$ $T = -40°\text{C}$
Convert $\Delta S°$ to $\text{kJ mol}^{-1}\ \text{K}^{-1}$ by dividing by 1000.	$\Delta S° = +22.1\ \text{J mol}^{-1}\ \text{K}^{-1}$ $= +0.0221\ \text{kJ mol}^{-1}\ \text{K}^{-1}$
Convert T to kelvin by adding 273.	$T = -40°\text{C}$ $= -40 + 273$ $= -233\ \text{K}$
Substitute the values into the Gibbs free energy change formula.	$\Delta G° = \Delta H° - T\Delta S°$ $= +6.01 - (233 \times +0.0221)$ $= 6.01 - 5.15$ $= +0.86\ \text{kJ mol}^{-1}$
Answer the question.	As $\Delta G° = +0.86\ \text{kJ mol}^{-1}$, ice will not spontaneously melt at –40°C.
Check your answer using the values of $\Delta H°$ and $\Delta S°$ and Table 16.3.1 (page 515 of the Student Book).	As both $\Delta H°$ and $\Delta S°$ are positive, Table 16.3.1 states that this reaction is spontaneous at relatively high T. This agrees with the calculated answer.

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Worked example: Try yourself 16.3.2

CALCULATING THE TEMPERATURE AT WHICH A REACTION BECOMES SPONTANEOUS

Ammonia reacts with hydrogen chloride according to the equation:

$$NH_3(g) + HCl(g) \rightarrow NH_4Cl(s)$$

For this reaction, $\Delta H^\circ = -176.7\,kJ\,mol^{-1}$ and $\Delta S^\circ = -284.8\,J\,mol^{-1}\,K^{-1}$. Calculate the highest temperature at which the reaction is spontaneous.

Thinking	**Working**
State the known variables.	$\Delta H^\circ = -176.7\,kJ\,mol^{-1}$ $\Delta S^\circ = -248.8\,J\,mol^{-1}\,K^{-1}$ $\Delta G^\circ = 0$ $T = ?$
Convert ΔS° to $kJ\,mol^{-1}\,K^{-1}$ by dividing by 1000.	$\Delta S = -248.8\,J\,mol^{-1}\,K^{-1}$ $= -0.2488\,kJ\,mol^{-1}\,K^{-1}$
Substitute the values into the Gibbs free energy change formula.	$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ $0 = -176.7 - T \times (-0.2488)$ Solving for T: $T = \frac{-176.7}{-0.2488}$ $= 710.2\,K$
Answer the question.	The highest temperature at which the reaction is spontaneous is 710.2 K or 437.2°C.
Check your answer using the values of ΔH° and ΔS° and Table 16.3.1 (page 515 of the Student Book).	As both the ΔH° and ΔS° are negative, Table 16.3.1 states that this reaction is temperature dependent and favoured at a relatively low T.

16.3 KEY QUESTIONS

1 B (a reaction in which ΔG° is negative). Option A is not necessarily true. A reaction may be spontaneous and slow. Options C and D are incorrect, because changes in entropy of the system alone do not determine whether a reaction is spontaneous.

2 A process that is always spontaneous is exothermic (has $\Delta H^\circ < 0$) and has an increase in entropy ($\Delta S^\circ > 0$).

3 Standard state conditions are defined at a temperature of 25°C, but the human cell operates at around 37°C. The difference in temperature will give a different Gibbs free energy value.

4 The change in entropy for the system is $\Delta S^\circ = \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants})$

$$\begin{aligned} &= (+26.9) - (+32.7 + (\tfrac{1}{2} \times +205.5)) \\ &= 26.9 - 135.5 \\ &= -108.6\,J\,mol^{-1}\,K^{-1} \end{aligned}$$

The entropy decreases.

5 $\Delta H° = \Sigma\Delta H°_f(\text{products}) - \Sigma\Delta H°_f(\text{reactants})$
$= (-641.6 + 0) - (0 + 2 \times (-92.3))$
$= -641.6 + 184.6$
$= -457.0\ \text{kJ mol}^{-1}$

$\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$
$= (+89.6\ + (+130.7)) - (+32.7 + 2 \times (+186.9))$
$= -186.2\ \text{J mol}^{-1}\ \text{K}^{-1}$
$= -0.1862\ \text{kJ mol}^{-1}\ \text{K}^{-1}$

$T = 298\ \text{K}$ at standard state conditions

$\Delta G° = \Delta H° - T\Delta S°$
$= -457.0 - (298 \times -0.1862)$
$= -401.5\ \text{kJ mol}^{-1}$

$\Delta G° < 0$, so the reaction is spontaneous at standard state conditions.

CHAPTER 16 **REVIEW**

1 A closed system allows energy but not mass to pass through the boundary; an open system allows both energy and matter to be transferred between the system and the surroundings.

2 Scenarios A, B and D lead to an increase in entropy. An increase in the number of mol of material or an increase in the number of mol of higher–entropy materials will increase the entropy of the system. In scenarios C and E there is an increase in disorder during the changes of state, so entropy will increase.

3 **a** decrease (gas becomes a liquid)
b decrease (4 mol of gas becomes 2 mol of gas)
c increase (gas produced)

4 **a** $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$
$= (+210.0) - (+43.0 + (\frac{3}{2} \times +130.7)$
$= 210.0 - 239.1$
$= -29.1\ \text{J mol}^{-1}\ \text{K}^{-1}$

There is a decrease in entropy, which is unsurprising because a $\frac{1}{2}$ mol of gas and 1 mol of liquid is lost in the process.

b $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$
$= (+114.4 + (3 \times +188.8)) - ((2 \times +210.0) + (4 \times +205.0))$
$= 680.8 - 1240.0$
$= -559.2\ \text{J mol}^{-1}\ \text{K}^{-1}$

5 The reaction is exothermic, so $\Delta S_{\text{surroundings}} > 0$. The reaction generates 1 mol of gas plus 2 mol of liquid per 2 mol of liquid reactant, so there is a net increase in the entropy of the system, $\Delta S_{\text{system}} > 0$. As $\Delta S_{\text{system}} + \Delta S_{\text{surroundings}} > 0$, the reaction is spontaneous.

6 No. The forward reaction is exothermic. The reverse reaction must be endothermic ($\Delta H > 0$), and therefore $\Delta S_{\text{surroundings}} < 0$. The forward reaction transforms 1 mol of liquid to 17 mol of gas, with an accompanying large increase in entropy. Therefore, the reverse reaction has a large decrease in entropy, $\Delta S_{\text{system}} < 0$. For the reverse reaction, $\Delta S_{\text{system}} + \Delta S_{\text{surroundings}} < 0$, so it is not spontaneous.

7 **a** The total entropy of the universe will increase.
b The total entropy of the universe will increase.

The total entropy of the universe will always increase if the reaction is spontaneous.

8 **a** ΔS_{system} increases, so the products have greater degree of disorder.

b An exothermic reaction will result in $\Delta S_{\text{surroundings}} > 0$. This coupled with $\Delta S_{\text{system}} > 0$ results in $\Delta S_{\text{universe}} > 0$, so the reaction is spontaneous.

9 **a** Melting is endothermic, as heat has to be absorbed to cause the phase change from a solid to a liquid.

b A liquid has more entropy than a solid, so $\Delta S_{\text{system}} > 0$.

c An endothermic process will result in $\Delta S_{\text{surroundings}} < 0$. This fact, coupled with $\Delta S_{\text{system}} > 0$, results in $\Delta S_{\text{universe}} > 0$ only at relatively high temperatures. Therefore, the process is only spontaneous above the melting point.

10 **a** $C_5H_{12}(l) + 8O_2(g) \rightarrow 5CO_2(g) + 6H_2O(l)$

b $\Delta S^\circ = \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants})$

$-414.5 = ((5 \times +213.8) + (6 \times +70.0)) - (S^\circ(C_5H_{12}) + (8 \times +205))$

$S^\circ(C_5H_{12}) = 414.5 + (5 \times +213.8) + (6 \times +70.0) - (8 \times +205)$

$= +264\ \text{J mol}^{-1}\ \text{K}^{-1}$

11 B. Since both change in entropy and enthalpy are positive, the reaction is spontaneous at high temperatures. To quantify the temperature at which the reaction becomes spontaneous, use:

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

Solve for T when $\Delta G^\circ = 0$. Also, convert ΔS° to $\text{kJ mol}^{-1}\ \text{K}^{-1}$ by dividing by 1000.

$0 = +30 - T \times (+0.050)$

Solving for T:

$T = \frac{30}{0.050}$

$= 600\ \text{K}$

The best answer is B; the reaction is spontaneous at temperatures above 600 K.

12 To quantify the temperature at which the reaction becomes spontaneous, use:

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

Solve for T when $\Delta G^\circ = 0$. Also, convert ΔS° to $\text{kJ mol}^{-1}\ \text{K}^{-1}$ by dividing by 1000.

$0 = -25 - T \times (-0.058)$

Solving for T:

$T = \frac{-25}{-0.058}$

$= 4.3 \times 10^2\ \text{K}$

The reaction becomes spontaneous when it reaches temperatures of 4.3×10^2 K and above.

13 To quantify the temperature at which the process becomes spontaneous, use:

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

Solve for T when $\Delta G^\circ = 0$. Also, convert ΔS° to $\text{kJ mol}^{-1}\ \text{K}^{-1}$ by dividing by 1000.

$0 = +30.9 - T \times (+0.0930)$

Solving for T:

$T = \frac{30.9}{0.0930}$

$= 332\ \text{K}$

$= 59^\circ\text{C}$

14 $T = 298\ \text{K}$ at standard state conditions.

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

$= +115 - (298 \times +0.125)$

$= 115 - 37.3$

$= +78\ \text{kJ mol}^{-1}$

15 $T = 298\ \text{K}$ at standard state conditions.

a $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$= -128 - (298 \times +0.205)$

$= -189\ \text{kJ mol}^{-1}$

b $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$= -100 - (298 \times +0.309)$

$= -192\ \text{kJ mol}^{-1}$

c The formation of both are spontaneous ($\Delta G^\circ < 0$). As $\Delta G^\circ_{\text{gas}} < \Delta G^\circ_{\text{liquid}}$, the formation of the gas is preferred (more spontaneous).

16 At $T = 10°C = 283\,K$

$\Delta G° = \Delta H° - T\Delta S°$

$= -238 - (283 \times +0.127)$

$= -238 - (-36)$

$= -202\,kJ\,mol^{-1}$

As $\Delta G° < 0$, the reaction is spontaneous. Similar calculations can be made at the other temperatures. However, the simple observations that $\Delta H° < 0$ and $\Delta S° > 0$ indicate that the reaction will be spontaneous at all temperatures.

17 $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$

$= (+109.6) - (+205.2 + 130.7)$

$= 109.6 - 335.9$

$= -226.3\,J\,mol^{-1}\,K^{-1}$

$= -0.2263\,kJ\,mol^{-1}\,K^{-1}$

$\Delta G° = \Delta H° - T\Delta S°$

$= -187.8 - (298 \times (-0.2263))$

$= -187.8 + 67.4$

$= -120.4\,kJ\,mol^{-1}$

The reaction is spontaneous.

18 $\Delta S° = \Sigma S°(\text{products}) - \Sigma S°(\text{reactants})$

$= (121.2) - (191.6 + 2 \times 130.7)$

$= 121.2 - 453.0$

$= -331.8\,J\,mol^{-1}\,K^{-1}$

$= -0.3318\,kJ\,mol^{-1}\,K^{-1}$

$\Delta G° = \Delta H° - T\Delta S°$

$= 50.6 - (298 \times (-0.3318))$

$= +149.5\,kJ\,mol^{-1}$

The reaction is not spontaneous.

19 $\Delta H° = \Sigma\Delta H°_f(\text{products}) - \Sigma\Delta H°_f(\text{reactants})$

$= (-126) - (+110)$

$= -236\,kJ\,mol^{-1}$

$\Delta G° = \Delta H° - T\Delta S°$

$\Delta S° = \Delta H° - \frac{\Delta G°}{T}$

$T = 298\,K$ at standard state conditions

$\Delta S° = -236 - \frac{(-168)}{298}$

$= -0.228\,kJ\,mol^{-1}\,K^{-1}$

$= -228\,J\,mol^{-1}\,K^{-1}$

20 In stretching a rubber band, the released warmth indicates that the processes are exothermic, so ΔH is negative. The stretching requires effort and is not spontaneous, so $\Delta G > 0$. From the Gibbs free energy change formula, the sign of ΔS must also be negative. Stretching a rubber band reduces the entropy in the rubber band system. The stretching straightens out the kinks in rubber, which reduces the disorder and is reflected in the decrease in entropy.

21 **a** The self-assembly of drinking straws occurs naturally, without continuous input of energy, so it is a spontaneous process. The sign of ΔG for a spontaneous process is negative.

b Since the straws become more ordered, the entropy decreases. The sign of ΔS for the process is negative.

c Since $\Delta G = \Delta H - T\Delta S$ (the Gibbs free energy change formula), then $\Delta H = \Delta G + T\Delta S$. If ΔG and ΔS are both negative, then ΔH must be negative. The process is exothermic.

d Since $\Delta G = \Delta H - T\Delta S$ and ΔS is negative, the term $-T\Delta S$ is positive. Increasing the temperature therefore makes the process less spontaneous (ΔG become more positive) and the straws become less likely to self-assemble.

e The same process occurs during crystallisation. When crystals are forming in a solution, under those conditions the entropy of the system is decreasing and the process is spontaneous and exothermic.

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Module 4 **Review answers**

Drivers of reactions

Multiple choice

1 D is correct.

$= 3.7 \times 10^{-8}$ MJ

2 B is correct.

A has a positive ΔH and heat is released to the surrounding environment. Positive ΔH correlates with energy absorption from the surroundings.

B. It has a positive ΔH and heat is absorbed from the surrounding environment.

C has a negative ΔH and heat is released to the surrounding environment. This is true of an exothermic reaction.

D has a negative ΔH and heat is absorbed from the surrounding environment. Negative ΔH correlates with energy being released to the surroundings.

3 A is correct.

$E = n \times \Delta H_c = \frac{16.0}{60} \times 2021 = 539\,\text{kJ}$

4 C is correct.

$\Delta T = 80.0 - 23.0 = 57.0\,\text{K}$

$q = m \times c \times \Delta T = 250.0 \times 4.184 \times 57.0 = 5.96 \times 10^4\,\text{J} = 59.6\,\text{kJ}$

5 C. The negative ΔH value indicates a decrease in enthalpy overall, but bond breaking is necessary for initiation so there will be an initial increase (i.e. an activation energy barrier).

6 B. $n(C_3H_8) = \frac{100}{44.09} = 2.27\,\text{mol}$

energy released (exothermic) $= \frac{2220}{1} \times 2.27 = 5.04 \times 10^3\,\text{kJ}$

7 B. 1 MJ = 1000 kJ

$n(CO_2) = \frac{1000}{2200} \times 3 = 1.35\,\text{mol}$

$V(CO_2) = n \times V_m = 1.35 \times 24.79 = 33.8\,\text{L}$

8 B. The enthalpy of formation is the enthalpy change associated with the formation of 1 mol of a compound from its elements at standard state conditions.

9 B. This reaction shows the methane molecule (CH_4) being broken into carbon and four hydrogen atoms. This requires breaking the four C–H bonds.

10 C.

		Bond energy (kJ mol^{-1})
bonds broken	2 × H–H 1 × O=O	2 × +436 +498
bonds formed	4 × O–H	4 × –463
enthalpy change		–482

11 A. $\Delta H^\circ_f = \Delta H_1 + (-\Delta H_2) = -966 + (-(-890)) = -76\,\text{kJ mol}^{-1}$

12 B. $\Delta H = -\Delta H_2 + (-\Delta H_1) = -1560 + (-(-1626)) = +86\,\text{kJ mol}^{-1}$

13 D.

Reverse equation 1: $\frac{1}{2}N_2(g) + \frac{3}{2}H_2(g) \rightarrow NH_3(g)$ $\Delta H^\circ = -46\,\text{kJ mol}^{-1}$

Reverse equation 2: $NO_2(g) + 2H_2(g) \rightarrow \frac{1}{2}N_2(g) + 2H_2O(l)$ $\Delta H^\circ = -170\,\text{kJ mol}^{-1}$

Overall: $NO_2(g) + \frac{7}{2}H_2(g) \rightarrow 2H_2O(l) + NH_3(g)$ $\Delta H^\circ = -46 + (-170) = -216\,\text{kJ mol}^{-1}$

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14 A.

$\Delta H^\circ = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$+205.3 = \Delta H^\circ_f(CO(g)) - (\Delta H^\circ_f(CH_4(g)) + \Delta H^\circ_f(H_2O(g)))$

$+205.3 = -110.5 - (\Delta H^\circ_f(CH_4(g)) + (-241.8))$

$+205.3 = -110.5 - \Delta H^\circ_f(CH_4(g)) + 241.8$

$\Delta H^\circ_f(CH_4(g)) = -205.3 - 110.5 + 241.8$

15 D. A solution has less entropy than a gas. Gases have greater entropy than solids and liquids, and a solution has greater entropy than a solid.

16 C. A gas turning into a solid is associated with a decrease in entropy.

17 D. The sublimation of CO_2 involves the conversion of a solid into a gas, so ΔS_{system} increases (+). This is an endothermic process, so heat is absorbed from the surroundings; thus, $\Delta S_{\text{surroundings}}$ decreases (+).

18 D. $\Delta G = \Delta H - T\Delta S$

For positive ΔH and negative ΔS, ΔG will be positive at all temperatures and therefore not spontaneous at any temperature.

19 A. $\Delta S^\circ = \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants}) = (2 \times 70.0 + 192.3) - (240.0 + \frac{7}{2} \times 130.6) = 332.3 - 697.1 = -364.8\,\text{J mol}^{-1}\,\text{K}^{-1}$

20 A. $\Delta G = \Delta H - T\Delta S = -104.0 - (303 \times -0.0608) = -104.0 - (-18.4) = -85.6\,\text{kJ mol}^{-1}$. $\Delta G < 0$ so the reaction is spontaneous.

Short answer

1 **a** A positive ΔH value correlates with an endothermic reaction.

b Coefficients are all tripled; therefore, triple the ΔH value: $3 \times (+129.2) = +387.6\,\text{kJ mol}^{-1}$

c $n(Al_2(SO_4)_3) = \frac{m}{M} = \frac{30.0}{342.17} = 0.0877\,\text{mol}$

$E = n \times \Delta H = 0.0877 \times 129.2 = 11.3\,\text{kJ}$ absorbed.

d Changes of state have an associated latent heat, which needs to be factored in when determining ΔH for the reaction.

2 **a** $q = m \times c \times \Delta T$

$= 50 \times 4.184 \times (70.0 - 15.0)$

$= 50 \times 4.184 \times 55.0$

$= 1.15 \times 10^4\,\text{J}$

$= 11.5\,\text{kJ}$

b $q = m \times c \times \Delta T$

$= 700 \times 4.184 \times (98.0 - 32.0)$

$= 700 \times 4.184 \times 66.0$

$= 1.93 \times 10^5\,\text{J}$

$= 193\,\text{kJ}$

c $q = m \times c \times \Delta T$

$= 2600 \times 4.184 \times (10.0 - 8.0)$

$= 2600 \times 4.184 \times 2.0$

$= 2.18 \times 10^4\,\text{J}$

$= 21.8\,\text{kJ}$

d $q = m \times c \times \Delta T$

$= 1000 \times 1.039 \times (200 - 175)$

$= 1000 \times 1.039 \times 25$

$= 25\,975\,\text{J}$

$= 26.0\,\text{kJ}$

e $q = m \times c \times \Delta T$

$= 270 \times 0.385 \times (54.0 - 26.0)$

$= 270 \times 0.385 \times 28.0$

$= 291\,\text{J}$

$= 2.9\,\text{kJ}$

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3 **a** Equation 1 is doubled. Hence, $\Delta H = 2 \times (+181) = +336\,\text{kJ mol}^{-1}$.

b Equation 1 is reversed. Hence, $\Delta H = -181\,\text{kJ mol}^{-1}$.

c Equation 2 is halved and reversed. Hence, $\Delta H = \frac{1}{2} \times (+114) = +57\,\text{kJ mol}^{-1}$.

d Equation 1 + equation 2. Hence, $\Delta H = (+181) + (-114) = +67\,\text{kJ mol}^{-1}$.

4 $\Delta H^\circ_{\text{reaction}} = \Sigma\Delta H^\circ_f(\text{products}) - \Sigma\Delta H^\circ_f(\text{reactants})$

$\Sigma\Delta H^\circ_f(\text{reactants}) = \Sigma\Delta H^\circ_f(\text{products}) - \Delta H^\circ_{\text{reaction}}$

$\Sigma\Delta H^\circ_f(\text{reactants}) = (3 \times -393.5 + 3 \times -285.8) - (-1790)$

$\Sigma\Delta H^\circ_f(\text{reactants}) = -247.9\,\text{kJ mol}^{-1}$

5 2 × reverse of equation 1: $2N_2O_3(g) \rightarrow 2NO(g) + 2NO_2(g) \quad \Delta H = +80\,\text{kJ mol}^{-1}$

Reverse equation 2: $2NO_2(g) \rightarrow N_2O_4(g) \quad \Delta H = -58\,\text{kJ mol}^{-1}$

Overall: $2N_2O_3(g) \rightarrow 2NO(g) + N_2O_4(g) \quad \Delta H = +80 + (-58) = +22\,\text{kJ mol}^{-1}$

6 Reverse equation 1: $KCl(s) + 3MgO(s) \rightarrow KClO_3(s) + 3Mg(s) \quad \Delta H = +1852\,\text{kJ mol}^{-1}$

Equation 2: $K(s) + \frac{1}{2}Cl_2(g) \rightarrow KCl(s) \quad \Delta H = -437\,\text{kJ mol}^{-1}$

3 × equation 3: $3Mg(s) + \frac{3}{2}O_2(g) \rightarrow 3MgO(s) \quad \Delta H = -1806\,\text{kJ mol}^{-1}$

$K(s) + \frac{1}{2}Cl_2(g) + \frac{3}{2}O_2(g) \rightarrow KClO_3(s) \quad \Delta H = +1852 + (-437) + (-1806) = -391\,\text{kJ mol}^{-1}$

7 **a**

	Species	Energy (kJ mol^{-1})
bonds broken	1 × C–C 6 × C–H 3.5 × O=O	1 × +346 6 × +414 3.5 × +498
bonds formed	4 × C=O 6 × O–H	4 × −804 6 × −463
enthalpy change		−1421

$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \quad \Delta H^\circ = -1421\,\text{kJ mol}^{-1}$

b Combustion: $C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H_c = ?$

$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g) \quad \Delta H^\circ = -1421\,\text{kJ mol}^{-1}$

$3H_2O(g) \rightarrow 3H_2O(l) \quad \Delta H^\circ = 3 \times -44\,\text{kJ mol}^{-1} = -132\,\text{kJ mol}^{-1}$

$C_2H_6(g) + \frac{7}{2}O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \quad \Delta H_c = -1553\,\text{kJ mol}^{-1}$

8 **a** A reaction changes from spontaneous to non-spontaneous when $\Delta G = \Delta H - T\Delta S = 0$. Rearranging, the reaction is spontaneous when $T = \frac{\Delta H}{\Delta S} = \frac{-44.1}{-0.0908} = 485\,\text{K} = 213°\text{C}$.

b Since the reaction has $\Delta H < 0$ and $\Delta S < 0$, the reaction is spontaneous at low values of T; so it is non-spontaneous at high values of T, when the temperature is greater than 213°C.

9 **a**

$$\begin{aligned} \Delta S^\circ &= \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants}) \\ &= S^\circ(C_2H_6(g)) - (S^\circ(C_2H_2(g)) + 2 \times S^\circ(H_2(g))) \\ -232.7 &= S^\circ(C_2H_6(g)) - (200.9 + 261.4) \end{aligned}$$

$S^\circ(C_2H_6(g)) = -232.7 + 462.3 = +229.6\,\text{J mol}^{-1}\,\text{K}^{-1}$

b $\Delta H^\circ = \Sigma\Delta H^\circ(\text{products}) - \Sigma\Delta H^\circ(\text{reactants})$

$= (-84.7) - (226.7 + 2 \times 0) = -311.4\,\text{kJ mol}^{-1}$

Standard state, so $T = 298\,\text{K}$

From Gibbs free energy change formula:

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

$= -311.4 - (298 \times -0.2327) = -311.4 + 69.3 = -242.1\,\text{kJ mol}^{-1}$

c The change in Gibbs free energy is $-242.1\,\text{kJ mol}^{-1}$ and therefore the reaction is spontaneous.

10 As all components are gaseous, and the reaction leads to an increase in the number of moles of gas (1 mol reactants becomes 2 mol products), ΔS° is positive. From the Gibbs free energy change formula, $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$, since ΔS° is positive, $T\Delta S^\circ$ is positive. As T increases, a larger value of $T\Delta S^\circ$ is subtracted from ΔH°, and so ΔG° becomes more negative.

Extended response

1 **a** $2C_5H_{11}OH(l) + 15O_2(g) \rightarrow 10CO_2(g) + 12H_2O(g)$

b $q = m \times c \times \Delta T$

$= 100 \times 4.184 \times 30.0$

$= 1.26 \times 10^4\,\text{J}$

$= 12.6\,\text{kJ}$

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

$= \frac{0.400}{88.15}$

$= 0.004\,54\,\text{mol}$

$\Delta H_c = \frac{q}{n}$

$= \frac{-12.6}{0.004\,54}$

$= -2.77 \times 10^3\,\text{kJ mol}^{-1}$

c total energy supplied $= \Delta H_c \times n(\text{pentan-1-ol})$

$= 3330 \times 0.004\,55$

$= 15.1\,\text{kJ}$

$\%\ \text{efficiency} = \frac{\text{actual energy absorbed}}{\text{total energy supplied}} \times 100\%$

$= \frac{12.6}{15.1} \times 100\%$

$= 82.8\%$

d $m(C_5H_{11}OH) = 42.5 \times 1000 = 42\,500\,\text{g}$

$n(C_5H_{11}OH) = \frac{m}{M} = \frac{42500}{88.15} = 482.13\,\text{mol}$

$E_{\text{chemical}} = n \times \Delta H_c = 3330 \times 482.13 = 1\,605\,502\,\text{kJ} = 1.61 \times 10^6\,\text{kJ}$ or $1.61 \times 10^3\,\text{MJ}$

$E_{\text{mechanical}} = \frac{45}{100} \times 1\ 61 \times 10^3\,\text{MJ} = 7.2 \times 10^2\,\text{MJ}$

2 **a** $NaCH_3COO \xrightarrow{H_2O(l)} CH_3COO^-(aq) + Na^+(aq)$

b Sodium ions and ethanoate ions are held in an ionic lattice by ionic bonds that are based on electrostatic attraction. These ionic bonds in the solute break when it dissolves in water. Water, the solvent, has hydrogen bonds between the water molecules. When sodium acetate is added to water, the hydrogen atoms of the water molecules are attracted to the negative acetate ions, and the oxygen atoms of the water molecules are attracted to the positive sodium ions. Ion–dipole bonds form between the ions and water molecules and the surface ions are pulled into solution. Gradually, the ionic network dissociates and a solution is formed.

c $\Delta T = 26.8 - 25.0$

$= 1.8\,\text{K}$

$q = m \times c \times \Delta T$

$= 75.0 \times 4.184 \times 1.8$

$= 565\,\text{J}$

$= 0.565\,\text{kJ}$

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

$= \frac{3.50}{82.03}$

$= 0.0427\,\text{mol}$

$\Delta H_{\text{soln}} = \frac{q}{n(\text{solute})}$

$= \frac{-0.565}{0.0427}$

$= -13.2\,\text{kJ mol}^{-1}$

There is a negative enthalpy of dissolution due to an increase in temperature of the water; therefore, the reaction is exothermic.

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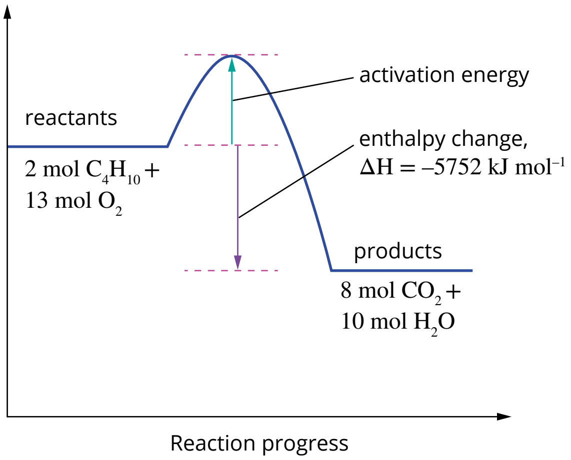
d $\Delta H_{soln} = \frac{q}{n(\text{solute})}$

$$q = \Delta H_{soln} \times n(\text{solute})$$
$$= 17\ 3 \times 0.0427$$
$$= 0.738\,\text{kJ}$$

$$\%\ \text{efficiency} = \frac{\text{actual energy absorbed}}{\text{total energy supplied}} \times 100\%$$
$$= \frac{0.565}{0.738} \times 100\%$$
$$= 76.5\%$$

3 **a** The equation has combustion of 2 mol butane, hence $\Delta H = 2 \times (-2876) = -5752\,\text{kJ mol}^{-1}$

b **Energy level diagram of butane combustion**



c Some initial input of energy is required to allow some reactants to overcome the activation energy barrier to initiate the rapid reaction.

d heat of combustion = $\frac{-2876}{58.12} = -49.48\,\text{kJ g}^{-1}$

e energy released = $3.00 \times 49.48 = 148\,\text{kJ}$

Therefore energy absorbed by water = $\frac{30}{100} \times 148 = 44.5\,\text{kJ} = 4.45 \times 10^4\,\text{J}$

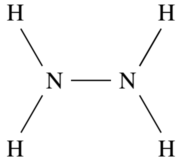
Using $q = m \times c \times \Delta T$:

$$\Delta T = \frac{q}{mc} = \frac{4.45 \times 10^4}{150 \times 4.184} = 71.0\,\text{K}$$

Therefore, final temperature = 20.0 + 71.0 = 91.0°C

4 The entropy of a perfect ordered substance at 0 K is defined as having a zero entropy. By measuring the entropy at any temperature, the change in entropy can be provided as an absolute value. There is no analogous zero point for enthalpy. The enthalpy of formation of an element is given as zero, but this is an arbitrary mark. There is no way of knowing the total enthalpy of a substance; only changes in enthalpy can be measured.

5 **a**



b

	Species	Energy (kJ mol⁻¹)
bonds broken	4 × N–H N–N O=O	4 × +391 +158 +498
bonds formed	N≡N 4 × O–H	–945 4 × –463
enthalpy change		–577

c $\Delta H^\circ = \Sigma\Delta H^\circ(\text{products}) - \Sigma\Delta H^\circ(\text{reactants})$

$$= (2 \times -241.8) - (+95.8)$$
$$= -579.4\,\text{kJ mol}^{-1}$$

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d There is a slight difference between the standard enthalpy change calculated using bond energies (–577) and using standard enthalpies of formation (–579.4). This difference is due to the use of average bond energies for the N–H, N–N and O–H bonds, which is an approximation of the actual bond energies of these bonds in the reactants and products of this reaction.

e $\Delta S^\circ = \Sigma S^\circ(\text{products}) - \Sigma S^\circ(\text{reactants})$

$= (191.6 + 2 \times 188.8) - (238.7 + 243.6)$

$= +86.9\,\text{J mol}^{-1}\,\text{K}^{-1}$

$= 0.0869\,\text{kJ mol}^{-1}\,\text{K}^{-1}$

f Standard state conditions, so $T = 298\,\text{K}$.

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

$= -579.4 - (298 \times -0.0869)$

$= -553\,\text{kJ mol}^{-1}$

Because $\Delta G^\circ < 0$, the reaction is spontaneous.

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