STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 1: Review of basic concepts

Check your understandng 1.1/1.2

1 Homogenous – substance with a uniform composition throughout. Heterogeneous – substance with non-uniform composition, where pieces of the material are different from other pieces.

2 Pure – substance that is not contaminated by any other substances. Impure – substance contaminated with small amounts of one or more other substances.

3 Element – substance that cannot be decomposed into simpler substances. Compound – substance that can be decomposed into simpler substances, thus composed of two or more simpler substances.

4 (Examples) Homogenous – pure water, sugar, petrol, apple juice. Heterogeneous – jam, wood, beef, concrete.

5 (Examples) Mixtures – milk, orange juice, dishwashing liquid, jam. Pure substances – pure water, sugar, salt, copper.

6 (Examples) Elements – aluminium, copper, oxygen (air), gold, nitrogen air), carbon (diamond). Compounds – sodium chloride (table salt), sucrose (sugar), sodium carbonate (washing soda), ethanol (alcohol).

7 **a** homogeneous

b heterogeneous

c homogeneous

d homogeneous

e heterogeneous

f heterogeneous

g heterogeneous

h homogeneous

heterogeneous or homogenous (depending on composition) – if striped or containing particles, will be heterogeneous; if smooth such as regular white, will be homogeneous

8 **a** Homogeneous – a diamond (pure substance), white wine (mixture), mercury (pure substance), dry ice (pure substance – carbon dioxide solid), toothpaste (mixture).

b All heterogeneous substances are mixtures.

9 Compound. Decomposed into two simpler substances, silver liquid and colourless gas, upon heating. Unable to be easily reformed into original substance.

10 Element. Did not decompose upon heating. Retained original properties after heating and cooling back to original temperature. Has a single constant melting point.

Check your understanding 1.3/1.4

1 **a** Melting
  b Evaporation or boiling or vaporisation
  c Freezing or solidification
  d Condensation or liquefaction
  e Sublimation

2 Solution – homogenous mixture in which the particles cannot be seen, even by a microscope, and never settle out. Example is salt or sugar in water. Suspension – mixture in which the particles are able to be seen in the solvent and are large enough to eventually settle out on standing. Example is sand in water, or milk.

3 Solvent – liquid in which a substance is dissolved. Solute – the substance that is being dissolved. Example in sugar water, sugar is the solute and water is the solvent.

4 Solutions – salt water (solute – salt, solvent – water), sugar water (solute – sugar, solvent – water), brandy (solute – ethanol, solvent – water). Suspensions – milk, paint, orange juice with pulp. All three have particles that can be seen, and upon standing will settle out of the solution.

5 Note: many of these depend on the exact composition – will vary between brands and types of the substance.
  a suspension
  b suspension
  c solution (if clear), suspension (if visible particles)
  d solution (if made from concentrate), suspension (if included pulp)
  e solution
  f solution
  g suspension
  h solution
  suspension
  j solution

Check your understanding 1.5/1.6

1 Gas is evolved; a solid (precipitate) forms; a change in colour is seen; a significant change in temperature occurs; a solid disappears (not just dissolving) or an odour is produced.

2 Refer to Figure 1.7 page 9.

3 **a** Particles are packed closely in an orderly array. Particles vibrate slightly about their rest positions.
  b Particles possess more kinetic energy than particles in a solid. Move about in a less orderly manner than solids. Motion involves random translation from one place to another within the liquid.
  c Particles have rapid, random motion. Particles are spread out far apart and, due to rapid motion, fill the container they are in.

4 Physical changes – changing the state of a substance, dissolving a solid in a liquid, separating mixtures, changing physical appearance by grinding. Chemical changes – gas is evolved, a precipitate forms when two solutions are mixed, a change in colour occurs, an odour is produced.

5 **a** chemical

  b physical

  c chemical

  d physical

  e physical

  f chemical

6 Chemical. Colour change from blue t colourless occurs. A brown solid forms on the zinc.

7 **a** Particles gain kinetic energy and ibrate faster. If enough kinetic energy is added to the particles they will begin to move and separate a the solid melts.

  b Particles gain kinetic energy and increase their translational motion, moving faster. If enough kinetic energy is added they will begin to separate ore, forming a gas.

8 Gases have large spaces between the partices, more than liquids. When the gas particles are brought together they will mix faster than the liquid particles which have less space between them for the new partiles.

Check your understandng 1.7/1.8

1 oxygen – O, nitrogen – N, carbon – C, hydrogen – H, fluorine – F, sulfur – S, phosphorus – P, boron – B

2 **a** beryllium – Be, bromine – Br, barium – Ba

  b chlorine – Cl, chromium – Cr, calcium – Ca

  c nickel – Ni, sodium – Na, neon – Ne

3 sodium – Na, potassium – K, iron – Fe, lead – Pb, tin – Sn, mercury – Hg

4 The elements present in the compound and the ratio in which they are present. In molecules, it shows the number of atoms present ineach molecule. In ionic substances, it shows the ratio of ions present in the ionic comound.

5 Oxygen does not exist as single atoms, but as a molecule with two atoms of oxygen joined together to form a molecule of O_2

6 Diatomic gases – nitrogen, N_2; oxygen, O_2; hydrogen, H_2; chlorine, Cl_2 Monatomic gases – helium, He; neon, Ne; argon, Ar.

7 **a** nitrogen – 2, oxygen – 1

  b sulfur – 1, oxygen – 2

  c hydrogen – 3, phosphorus – 1, oxygen – 4

  d carbon – 2, hydrogen – 6, oxygen – 1

8 **a** nitric oxide – NO, water – H_2O, carbon disulfide – CS_2, hydrogen peroxide – H_2O_2, carbon monoxide – CO, carbon dioxide – CO_2

  b PH_3 **ii** N_2O_5 **iii** CON_2H_4 **v** $C_{12}H_{22}O_{11}$

9 **a** 5 **b** 14 **c** 9 **d** 11

10 **a** K O = 2 : 1

b iron : sulfur = 2 : 3

c nitrogen : hydrogen :sulfur : oxygen = 2 : 8 : 1 : 4

d calcium : phosphorus : oxygen = 3 : 1 : 4

Chapter review questions

1 Depending upon the composition. Some gases are heavier than others so will sink to the bottom of a sample of gas making a heterogneous gas. If the gases are evenly mixed, the gas sample will be homogenous.

2 A mixture can be homogenous if all parts of the mixture have uniform composition. Petrol and apple juice are examples of homogenous mixtures as each part of the liquid is the same.

3 **a** chemical

b physical

c chemical

d chemical

e physical

f physical

4 Compound. Contains a fixed proportion of aluminium, hence has constant composition. Properties did not change when subjected to heating. Formed from two elements, aluminium and sulfur, so contains more than one element.

5 coal – mixture, carbon – element, hydrogen – element, sulfur – element, carbon dioxide – compound, sulfur dioxide – compound, sulfur trioxide – compound, sulfuric acid – compound, marble – mixture, benzene – compound, naphthalene – compound, anthracene – compound

6 **a** increase as gas particles compress due to the large spaces in between them

b Number of particles would remain constant, but would increase in kinetic energy. Pressure would increase as the temperature increases.

7 **a** $C_2H_2F_4$

b $C_3H_7O_2SN$

c $C_2H_3NO_4$

8 **a** carbon – 8, hydrogen – 18

b carbon – 1, oxygen – 1, nitrogen – 2, hydrogen – 4

c carbon – 4, hydrogen – 8, oxygen – 2

d carbon – 10, hydrogen – 16, nitrogen – 2, oxygen – 8

9 **a** aluminium : chlorine = 1 : 3

b magnesium : nitrogen = 3 : 2

c zinc: nitrogen : oxygen = 1 : 2 : 6

d aluminium : sulfur : oxygen = 2 : 2 : 8

STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 2: Properties of matter

Check your understandng 2.1

1 **a** Physical properties – characteristics that can be observed or measured without changing into a different substance. Chemical properties – properties associated with the chemical changes when mixed with other substances or heated or exposed to light.

b Metals have physical properties including high melting points and good electrical conductivity, and chemical properties including reaction with acids to produce hydrogen gas. Hydrogen gas has physical properties including low boiling point and low density, and chemical properties including reaction with oxygen producing large amounts of heat.

2 **a** Melting – change from solid to liquid. Freezing – change from liquid to solid.

b Boiling and evaporating – both involve change from a liquid to a gas. Boiling – with formation of bubbles. Evaporating – without formation of bubbles.

c Boiling point – lowest temperature at which the liquid boils. Normal boiling point – boiling point at a pressure of 1000 kPa.

d Gas – physical state of matter where the particles are fast moving and spread out. Vapour – a gas that is easily liquefied or condensed (a gas close to is boiling point).

3 Density – mass of a substance per unit volume. Common units are grams per millilitre (g/mL or $g\ mL^{-}$) or kilograms per cubic metre ($kg\ m^{-3}$)

4 **a** Pure solids have sharp melting points that do not increase after undergoing further purification processes. A mixture would show a range of melting points.

b Pure liquids boil at fixed temperatures when at standard pressure. A mixture would boil over a range of temperatures.

5 **a** carbon tetrabromide, gallium, lead bromide, magnesium, sulfur

ii bromine, carbon disulfide, hydrogen peroxide, phosphorus trichloride, sulfuric acid

b carbon disulfide (liquid), neon (gas), phosphorus trichloride (liquid), sulfur dioxide (gas)

6 **a** $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{1\ 30}{0\ 50 \times 1\ 20 \times 0\ 80} = 2\ 7\ \text{g/cm}^3$

b $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{7\ 40}{5\ 0} = 1\ 5\ \text{g/mL}$

c $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{1\ 506}{\pi \times 0\ 15^2 \times 3\ 0} = 7\ 1\ \text{g/cm}^3$

7 **a** $\text{mass} = \text{density} \times \text{volume} = 0\ 79 \times 5\ 0 = 4\ 0\ \text{g}$

b $\text{mass} = \text{density} \times \text{volume} = 9\ 0 \times \left(0\ 8\ 0^3\right) = 4\ 6\ \text{g}$

c $\text{mass} = \text{density} \times \text{volume} = 0\ 66 \times 25\ 0 = 17\ \text{g}$

8 **a** $\text{volume} = \dfrac{\text{mass}}{\text{density}} = \dfrac{10}{1\ 48} = 6\ 8\ \text{mL}$

b $\text{volume} = \dfrac{\text{mass}}{\text{density}} = \dfrac{25}{1\ 11} = 23\ \text{mL}$

Check your understanding 22

1 **a** Solution – homogenous mixture in which the particles cannot be seen, even by a microscope, and never settle out. Suspension – mixture in which the particles are able to be seen in the solvent and are large enough to eventually settle out on standing.

b Solvent – liquid in which a substance is dissolved. Solute – the substance that is being dissolved.

c Immiscible liquids, when mixed, do not form a homogenous liquid but stay as drops of one liquid dispersed through the other. Miscible liquids form a homogenous liquid when mixed.

2 Sedimentation – process in which solids settle to the bottom of a container. Clay or dirt in water will form a sediment after some time. Decantation – process of pouring off the liquid and leaving the solid undisturbed at the bottom of the container. Pouring tea and leaving behind tea leaves is an example.

3 Sieving – separation based on particle size. A mesh is used that is smaller than one of the particles and larger than the other particles in the mixture. The smaller particles pass through the sieve and larger particles are trapped.

Filtration – separates a suspended solid from a liquid based on insolubility (for diagram see Figure 2.4 page 5). Liquid or solution passes through filter paper while the suspended solid is trapped in the filter paper due to its larger particle size. Liquid that passes through is called the filtrate.

Evaporation to dryness – separates and collects a solute from its solvent (for diagram see Figure 2.5 page 5). To obtain the solute in a solution heat a solution in an evaporating basin to boil off the solvent leaving the solute behind as a solid.

Distillation – separates liquids of different boiling point (for diagram see Figure 2.6 page 26). Solution or liquid mixture is heated until one of the liquid components is boiled off. The vapour is condensed back into liquid in the condenser and collected. The collected liquid is called the distillate.

Fractional distillation – separates liquids of different boiling points (for diagram see Figure 2.7 page 26). Mixture of liquids is put through successive distillations in one piece of equipment, allowing for liquids of multiple boiling points to be collected as different fractions. Also used to increase the percentage of distillate able to be recovered.

Separating funnel – separation based on solubility of different liquids (for diagram see Figure 2.9 page 8). Liquid on the bottom of the funnel is run out through the stopcock so it is collected without contamination from the top liquid.

4 Difference in particle size allows use of a sieve. The larger lead pellets would be unable to pass through, but the smaller aluminium oxide crystals would pass through the sieve.

5 Step1: add the solid crystals to water. The magnesium sulfate will dissolve; the barium sulfate will stay in solid form as a suspension.

Step2: Use filtration to separate the insoluble barium sulfate. The crystals will collect in the filter paper. Dry in an oven overnight to remove any water.

Step3: Place the magnesium sulfate solution into an evaporating dish and heat until all the water has boiled off. Dry the remaining crystals in an oven overnight

6 **a** Evaporate to dryness – the water will boil off leaving the sucrose behind.

b Distillation – similar to evaporation except the water vapour would be collected and condensed in the apparatus so the liquid water could be collected. The sucrose would be left behind in the flask.

7 **a** Ethanol has a lower boiling point (78°C) than the glycerol so it would evaporate first and be condensed and collected, hence the distillate would be ethanol.

b Acetic acid (118°C) has a higher boiling point than water so the water will boil off and condense and be collected as the distillate. The distillate would be less concentrated in acetic acid. Ideally there would be no acetic acid in the distillate.

c Chloroform (62°C) has a lower boiling point than ethanol (78°C) so the chloroform will boil off, condense and be collected. The distillate will be richer in chloroform than the starting solution.

d Undertake fractional distillation so the chloroform distillate will be much purer than by simple distillation.

e Hexane (69°C) has a lower boiling point than heptane so the distillate will be hexane that will evaporate first, condense and be collected.

Worked example .1

Total mass of two sulfates = $1.76 + 354 = 5.30$ g

As this is within 0.02 g of the mass of the sample, within experimental error, the sample contains no other substances.

silver sulfide: $\frac{1\;76}{5\;32} \times 100 = 33\;1\%$

zinc sulfide: $100 - 33.1 = 669\%$

Thus, the sample consists of 33.1% silver sulfide and 66.9% zinc sulfide.

Check your understanding 23

1 A method of determining the masses of substance present in a sample.

2 $\frac{3\;63}{500} \times 100 = 0\;726\%$

3 urea: $\frac{1\;79}{3\;67} \times 100 = 48\;8\%$

ammonium phosphate: $\frac{0\;81}{3\;67} \times 100 = 22\%$

potassium chloride: $100 - 488 - 22 = 29\%$

4 ethanol: $\frac{3.79}{18.33} \times 100 = 20.7\%$

ethylene glycol: $\frac{8.64}{18.33} \times 100 = 47.1\%$

glycerol: $100 - 20.7 - 47.1 = 32.2\%$

Check your understanding 24/25

1 Bromine and mercury

2 Oxygen, hydrogen, helium, nitrogen, fluorine, neon

3 **a** Metals are elements that are lustrous (shiny), are good conductors of heat and electricity, are malleable (rolled into sheets) and ductile (drawn into wires).

b Non–metals are elements that are usually non–lustrous, usually poor conductors of heat and electricity, are not malleable or ductile.

c Semi–metals are elements that have properties of both metals and non–metals and are not able to be classified as either metal or non–metal.

4 **a** Metals are lustrous (shiny) while non–metals are usually dull. Metals are good conductors of heat and electricity while non–metals are often (but not always) poor conductors. Metals are able to be rolled into sheets and drawn into wires (malleable and ductile) while non–metals are usually brittle or unable to change shape without breaking the structure.

b Conduction of electricity as it is easily tested. Ability to bend an element is also an easy test to perform. Both provide clear results that are not particularly subjective.

5 **a** Group – vertical columns, numbered 1 to 18.

b Period – horizontal rows, numbered 1 to 7.

c Transition elements – elements in groups 3 to 12.

6 Top and right of the table.

7 **a** Increases across a period.

b Increases down a group.

8 **a** (Examples) Gold/silver – jewellery; malleable and ductile so can be made into any shape desired. Aluminium – cookware; excellent conductor of heat. Iron – structural; high tensile strength. Copper – wiring; excellent conductor of electricity.

b (Examples) Necklaces, earrings, bracelets, rings. Made from gold, silver, copper, platinum. All are malleable and ductile so can be made into any shape required.

9 D and E – only elements that have both high melting point, around iron's melting point or higher, and good electrical conductivity.

10 a

Element	Melting point (°C)	Density (g/cm³)	Electrical conductivity (S/m)
Cadmium	321	865	14×10^{7}
Indium	157	731	12×10^{7}
Iodine	114	624	1×10^{-7}
Lithium	180	053	1.1×10^{7}
Molybdenum	2617	1022	2×10^{7}
Platinum	1772	2145	94×10^{6}
Phosphorus	44	182	1×10^{7}
Sulfur	113	207	1×10^{-15}
Tantalum	2996	1665	77×10^{6}
Titanium	1660	454	25×10^{6}
Uranium	1132	2020	36×10^{6}

b Metals (medium to high melting points and good conductivity) – cadmium, molybdenum, platinum, tantalum, titanium, uranium

Non-metals (lower melting points, often lower conductivity and lower density) – iodine, sulfur, phosphorus

Difficult to classify – lithium and indium. Should be metals but have quite low melting points. Classified as metals.

c tantalum, molybdenum, platinum, titanium, uranium, cadmium, lithium, indium, iodine, sulfur, phosphorus

ii platinum, uranium, tantalum, molybdenum, cadmium, indium, iodine, titanium, sulfur, phosphorus, lithium

iii molybdenum, cadmium indium, lithium, phosphorus, platinum, tantalum, uranium, titanium, iodine, sulfur

Chapter review questions

1 Particle size

2 (Example) Clay in water makes a suspension with dispersed clay particles. Over time the clay particles will settle and form a sediment. The water can be carefully poured off (decanted) leaving the clay sediment behind.

3 (Example) Salt water when evaporated to dryness will leave behind salt crystals that can be collected.

4 a Distillation

b Fractional distillation

5 Add both solids to water and stir until one of them is fully dissolved, the other in suspension. Use filtration where the mixture is poured through a filter paper – the insoluble solid is collected in the filter paper, the liquid that passes through contains the soluble solid. The filtrate collected is evaporated to dryness so the soluble solid crystallises.

6 Pure substances have a fixed melting point; impure substances melt over a range of temperatures.

7 **a** Top and right of the table

b Diagonal band from top of group 13 to the bottom of group 17

8 **a** berylium, calcium

b fluourne, bromine

c phosphorus, arsenic

d sodium, potassium

9 (Examples) A mixture of sand and rocks/pebbles. A mixtue of soil and rocks.

10 Use a magnet to attract th iron filings. As they are magnetic and sulfur is not, the iron filings can be collected, leaving the ulfur behind.

11 **a** 1 filtration – see igure 2.4 page 25

2 separating funnel – see Figure 2.9 page 28

3 distillation – see Figure 2.6 page 26

4 evaporating to dryness – ee Figure 2.5 page 26

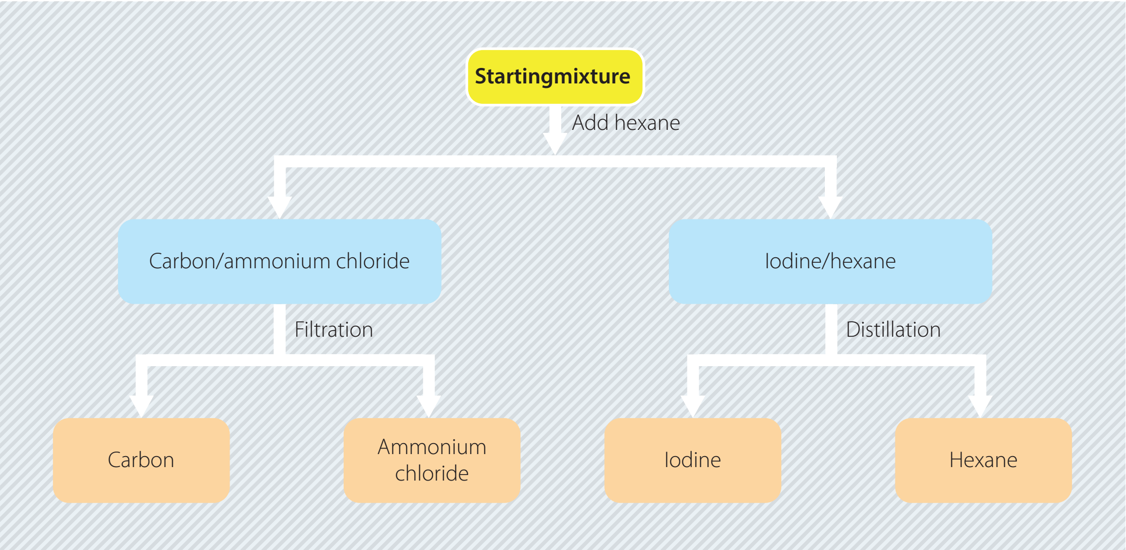
b X charcoal

Y sodium sulfate

Z kerosene

c Instead of using evaporation at step 4, use distillation so the water evaporated off could be condensed and collected separate to th sodium sulfate.

12



13 At −196°C the nitrogen would boil off, then at −186°C the argon woul boil off. At −186°C the oxygen would still be a liquid as it doesn't boil until −183°C

14 Fractional distillation could be used to separate the individual components. The mixture would be heated, and as each component evaporated at its boiling point it would rise up the apparatus, be condensed and collected prior to the next cmponent evaporating. This method separates based on boiling point.

15 Dissolve the mixture in a solvent such as hexane, then filter the mixture so the tar solids would be collected while the hexane solvent and the dissolved waxes would pass through and could be further separated to recove the waxes. This separation is done based on solubility of different compounds.

16 $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{142\ 6}{34\ 7 - 22\ 2} = 11.4\ \text{g/mL}$

17 **a** $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{092}{172} \times 100 = 53\%$

b $\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{081}{138} \times 100 = 59\%$

$\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{100}{204} \times 100 = 49\%$

$\text{density} = \dfrac{\text{mass}}{\text{volume}} = \dfrac{105}{187} \times 100 = 56\%$

Mixture. A compound would have a consistent percentage of copper, whereas a mixture has a variable composition as seen here with the varying percntages from each test.

18 Total mas: $0.324\ + 1.90 + 1.05 + 2.13 = 5404$ g

aspiin: $\dfrac{0324}{5404} \times 100 = 6.0\%$

sodium bicarbonate: $\dfrac{190}{5404} \times 100 = 35.2\%$

citric acd: $\dfrac{105}{5404} \times 100 = 19.4\%$

filler compound: $\dfrac{213}{5404} \times 100 = 39.4\%$

STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 3: Atomic structure

Check your understandng 3.1

1 Contains a small, dense nucleus that has most of the mass of theatom and is positively charged. Surrounded by a cloud ofelectrons.

2 The electrons move randomly and quickly around the nucleus, and their distance from the nucleus varies greatly. The term ‘cloud’ is used to describe the space electrons move about in rather than describing specific electron positions.

3 Protons and neutrons have approximately the same mass, an electron is approximately one–two thousandth $\left(\frac{1}{2000}\right)$ the size of a proton o neutron. The proton and electron have the same size but opposite charges: proton is positive, electron is negativ. The neutron has no charge.

4 A hydrogen atom consists of a proton and an electron, thus there is very little difference between the mass of a hydrogen atom and a proton. The approximate difference is the mass of an electron.

5 Atomic number – number of protons in the nucleus of n atom of the element. Mass number – number of protons plus neutron in the nucleus. Nucleon number – nucleons include protons and neutrons, so nucleon number is the same as the mass number.

6 Atomic number 11. Mass umber 23.

7 15 protons, 16 neutrons, 15 electrons

8 & 9

Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
Helium	2	4	2	2	2
Oxygen	8	16	8	8	8
Boron	5	11	5	6	5
Sulfur	16	32	16	16	16
Bromine	35	80	35	45	35
Iodine	53	127	53	74	53
Argon	18	40	18	22	18
Zinc	30	65	30	35	30
Scandium	21	45	21	24	21

Element	Atomic number	Mass number	Number of protons	Number of neutrons	Number of electrons
Titanium	22	48	22	26	22
Vanadium	23	51	23	28	23
Chromium	24	52	24	28	24
Manganese	25	55	25	30	25
Iron	26	56	26	30	26

Check your understandng 32

1 Isotopes of an element have the same number of protons, but different numbers of neutrons in the nucles.

Exampe 1: chlorine has 17 protons, but can have 18 or 20 neutrons, represented by chlorine-35 or chlorine-37.

Exampe 2: boron has 5 protons, but can have 5 or 6 neutrons, represented by boron-10 or boron-11.

Exampe 3: copper has 29 protons, but can have 34 or 36 neutrons, represented by copper-63 or copper-65.

2 M is the element symbol, A is the mass number, Z is the atomic umber.

3 Both have 16 protons, but sulfur-32 has 16 neutrons whie sulfur-35 has 19 neutrons.

4 Percentage of the isotope in the naturally occurring element.

5

Number of …	${}^{9}_{4}\mathrm{Be}$	${}^{20}_{0}\mathrm{Ne}$	${}^{39}_{9}\mathrm{K}$	${}^{84}_{36}\mathrm{Kr}$	${}^{37}_{56}\mathrm{Ba}$	${}^{88}_{38}\mathrm{Sr}$
… protons	4	10	19	36	56	38
… neutrons	5	10	20	48	81	50
… electrons	4	10	19	36	56	38

6 a 5 elements

b M-31 and M-33 (both atomic nmber 15). M-51 and M-53 (both atomic number 23).

c M-31 and M-33 (both atomic nmber 15). M-51 and M-53 (both atomic number 23). Atomic number is number of electrons.

d M-53 and M-58 both have 30 protons ($53 - 23 = 30$ and $58 - 28 = 30$)

7 a ${}^{24}_{12}\mathrm{Mg}$ and ${}^{27}_{12}\mathrm{Mg}$

b ${}^{206}_{82}\mathrm{Pb}$ and ${}^{208}_{82}\mathrm{Pb}$

8 Most elements have a number of different isotopes, each of which is a different type of atom, even when of the same element.

Worked examle .1

a ${}^{220}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{216}_{84}\mathrm{Po}$

b ${}^{90}_{38}\mathrm{Sr} \rightarrow {}^{90}_{39}\mathrm{Y} + {}^{0}_{-}e + \nu e$

Check your understandng 33

1 **a** Spontaneous emission of radiation from unstable isotopes.

b Isotopes of elements that spontaneously emit radiation.

2 Alpha and beta radiation are particles while gamma radiation is electromagnetic radiation. Alpha particles have a double positive charge and are heavy (2 protons + 2 neutrons). Beta particles have a single negative charge and are ver light (electron). Gamma radiation has no charge and no mass.

3 Alpha particles have very low penetrating power and can b stopped by a sheet of paper. Beta particles have medium penetrating power and can be stopped by thick aluminium or thin sheets of lead. Gamma radiation has high penetrating power and needs thick led or concrete to stop it.

4 Alpha particle – helium nucleus, composed of 2 protons and 2 neutrons, low energy, high mass particle with low penetratng power. Beta particle – electron, higher energy than alpha particle, low mass particle with reasonably high penetrating power.

5 The time required for half the atoms in a sample to undergo radioactive decay. Half-life for an element is very consistent and does not depend on the size of the sample or its age.

6 **a** $^{235}_{88}\text{Ra}$ **b** $^{222}_{86}\text{Rn}$ **c** $^{87}_{37}\text{Rb}$ **d** $^{90}_{38}\text{Sr}$

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

b $^{220}_{86}\text{Rn} \rightarrow {}^{4}_{2}\text{He} + {}^{216}_{84}\text{Po}$

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

d $^{24}_{11}\text{Na} \rightarrow {}^{0}_{-}e + {}^{24}_{12}\text{Mg}$

8 **a** Francium-214 **ii** Thallium-206

b Bismuth-214 **ii** Lead-206

9 Alpha and beta particles with a charge will be deflected by an electric field, gamma radiation with no charge will be naffected.

10 100 → 50 → 25 → 12.5 → 625 → 3.125 → 156 → 078 Hence 7 half-lives = 7 × 15 hours = 105 hours

Check your understandng 34

1 1st = 2, 2nd = 8, 3rd = 18, 4th = 32

2 B (2,3), P (2,8,5), Ar (2,8,8), Ca (2,8,8,2), Mn (2,8,13,2), Kr (2,8,18,8), Rb (2,8,18,8,1)

3 **a** Electron configurations end with an 8.

b Electron configurations ed with a 1.

c Electron configurations ed with a 7.

4 Atoms tend to lose, gain or share electrons to achieve the nearest stable electron configuration, usually that of the neares noble gas.

5 **a** Alkali metals all have 1 electron in the outer shell so will react chemically the same way and exhibit the same chemical properties.

b Halogens all have 7 electrons in the outer shell so will react the same way and have the same chemical properties.

6 **a, b** First shell must be full – 2 electrons – before the second shell begins to fill. Correct configuration: 2,1

ii Second shell must be full – 8 electrons – before third shell fills. Correct configuration: 2,8,3

iii Second shell only fills to 8 electrons Correct configuration: 2,8,3

v Second shell only fills to 8 electrons Correct configuration: 2,8,18,2

v Third shell fills to 8 electrons, then 4th shell to 2 electrons, then 3rd shell continues filling to 18 electrons. Correct configuration: 2,8,13,2

v First shell only fills to 2 electrons. Correct configuration: 2,8,9,2

7 O (2,6), S (2,8,6), Se (2,8,18,6), Te (2,8,18,18,6)

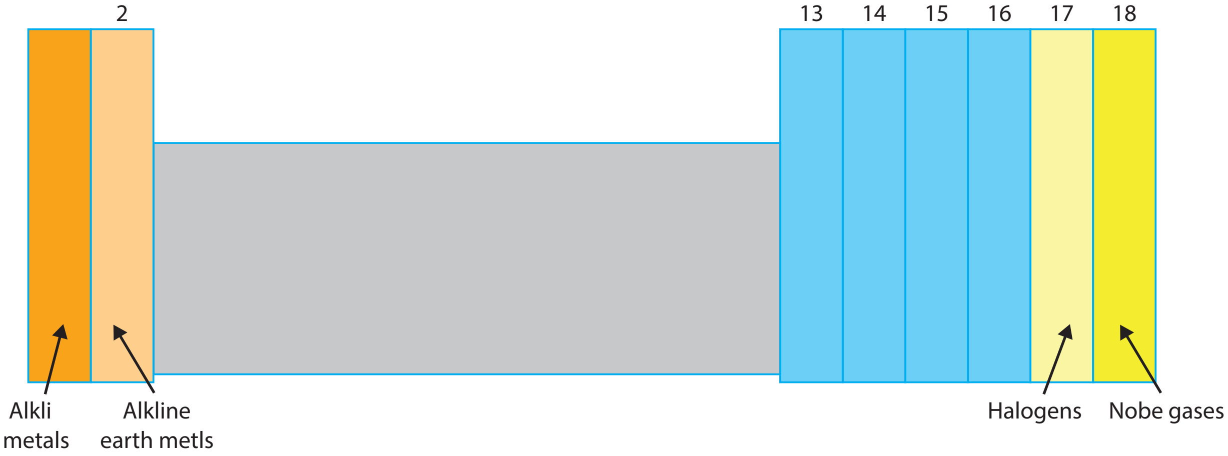
Check your understanding 35

1 Elements are listed in order of increasing atomic number across each period. Atoms with similar electron configurations are arranged vertically underneath one another in groups.

2 **a** Groups

b Periods

3



4 Transition metals or transition elements

5 Electrons in the highest energy level, or electrons in the outermost energy shell. These electrons are involved in bonding and chemical reactions.

6 The number of valence electrons – Group 1–2 and 13–18 – can be read directly from the periodic table. Group 1 and 2 have 1 and 2 valence electrons respectively. Group 13–18 can be found by subtracting 10 from the group number. E.g. Group 13 elements have 13 – 10 = 3 valence electrons.

7 An electron configuration the same as the nearest noble gas element.

8 **a** 2, 5

ii 2, 8, 3

iii 2, 8, 6

v 2, 8, 8, 2

b gain 3 electrons

ii lose 3 electrons

iii gain 2 electrons

v lose 2 electrons

9 see **8a**

10 a 2 electrons

b 4 electrons

11 a aluminium, gallium – all have 3 valence electrons

b silicon, germanium – all have 4 valence electrons

c phosphorus – both have 5 valence electrons

12 a 4 – group 14 (14 – 10 = 4 valence electrons)

b 7 – group 17 (17 – 10 = 7 valence electrons)

c 1 – group 1 = 1 valence electron

d 3 – group 13 (13 – 10 = 3 valence electrons)

13 a A (2,8), B (2,8,8), C (2,8), D (2,8,18,8), E (2,8,8)

b E is already in a stable (full) electron configuration so will not change its configuration and is stable.

14 B (2, $\underline{3}$) – group 13 (13 – 10 = 3)

Si (2, 8, $\underline{4}$) – group 14 (14 – 10 = 4)

S (2, 8, $\underline{6}$) – group 16 (16 – 10 = 6)

F (2, 7) – group 17 (17 – 10 = 7)

Check your understandng 36

1 Volume of space surrounding the nucleus of an atom in which ne or two electrons move.

2 The main energy levels are made up of sublevels which can be *s*, *p*, *d* and *f*, depending on the energy levl. Each sublevel contains between 1 to 7 orbitals. The *s* sublevel has 1 orbital, the *p* sublevel has 3 orbitals, the *d* sublevel has 5 orbitals and the *f* sublevel has 7 rbitals.

3 See Figur 3.7 pae 61.

4 9 electons: 1 s^2 $2s^2$ $2p^5$

13 electons: 1 s^2 $2s^2$ $2p^6$ $3s^2$ $3p$

33 electons: 1 s^2 $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ $3d^{10}$ $4p^3$

5 a C (2,4), $1s^2$ $2s^2$ $2p^2$

b O (2,6), $1s^2$ $2s^2$ $2p^4$

c Mg (2,8,2), $1s^2$ $2s^2$ $2p^6$ $3s^2$

d Cl (2,8,7), $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^5$

e K (2,8,8,1), $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s$

f Ca (2,8,8,2), $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$

g Co (2,8,15,2), $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^7$

h Ge (2,8,18,4), $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$

6 **a** A: $1\ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p$

b Si: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^2$

c F: $1\ \ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$

d Ga $1\ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p$

e Z: $1\ \ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}$

f S: $1\ \ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^4$

g Kr: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$

h Sr: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2$

7 Group 2: magnesium, calcium, strontium – all end in s^2

Grou 13: aluminium, gallium – both end in p

Grou 14: carbon, silicon, germanium – all end in p^2

Grou 16: oxygen, selenium – both end in p^4

Check your understandng 37/38

1 Thin, bright, coloured lines at different wavelengths o a black background.

2 **a** Electrons when heated move to a higher energy leve,where they are unstable. After a period of time, the electrons transition back to their original, lower energy level releasing energy as visible, ultraviolet or infrared ight. The difference between different energy levels varies, giving rise to lines of different colours related to the differet wavelengths.

b Electrons can move between multiple energy levels in the atom, giving rise to different energy lines produced for the same element.

3 High temperature flames decompose compounds into elements and the atoms of certain elements give the flame a distinctive colour. The colour comes from an electron transition in the atom that occurs more frequently than the others, producing liht of mostly one colour. Calcium produces a brick-red flame colour, copper a blue-green, sodium gives yellow while strontium gives a scarlet or deep red color.

4 Sample 1 – cadmium, copper

Sample 2 – chromium, strontium

5 Impurity is chromium

Check your understandng 39

1 Electrons move around the nuclus in fixed orbits. When they absorb energy they move to orbits of larger radius, and emit energy in the form of light when they move back to the smaller orbit.

2 Bohr thought electrons could only have discrete energies, related to the distinct energy levels present in an atom. Electrons moving between energy levels in the same type of atom would emit light of the same energy and wavelength, so a particular element would always give the same emission spectrum.

3 Bohr thought electrons were particles of matter. Schrödinger proposed the electron was a wave when in an atom, rather than particle. Schrödinger calculated that there is a significant probability of finding an electron in any position in a spherical volme around the nucleus. This did not match Bohr's idea of electrons in fixed radius orbits.

4 Schrödiger's equation allowed the explanation of emission spectra of atoms with multiple electrons, which Bohr's model faled to do. Schrödinger's equation also showed how sublevels and orbitals form in atoms with more thanone electron.

Chapter reiew quesions

1 **a** In the nucleus

  b Empty space

2 The electrons move randomly and quickly around the nucleus, and their distance from the nucleus varies greatly. The term 'cloud' is used to describe the space electrons move about in rather than describing specific electron positions.

3 Atoms of the same element all have the same numberof protons in the nucleus. Atoms of different elements have different numbers of protons.

4 Both isotopes have 17 protons and 17 electrons, butdifferent numbers of neutrons. Chlorine-35 has 18 neutrons; chlorine-37 ha 20 neutrons.

5 Stable atoms have a balance between the electrostatic repulsion force between positively charged protons and the mass–mass attractions that exist between all particles in the nucleus. If these forces are unbalanced so one is stronger than another, the nucleus is unstable. Sodium-24 and plutonium-239 are both unstable sotopes.

6 When an unstable isotope is neutron-rich, a neutron decays into a proton and an electron, $n \rightarrow p + e$. The proton remains in the nucleus, hanging the n : p ratio, and the electron is ejected from the nucleus as beta radiation.

7 Nuclear equations show the changes in mass and atomic numbers in a reaction where the nucleus changes either the number of protons or neutrons. (Example) $^{220}_{86}\text{Rn} \rightarrow\ ^{4}_{2}\text{He} +\ ^{216}_{84}\text{Po}$ The mass numbers add up $(220 = 216 + 4)$ as the number of nucleons dos not change. The atomic numbers also add up $(86 = 2 + 84)$

8 Simple form: 28,8,1. Orbita notation: $1\ s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s$

9 Elements in the same group have the same number of valence electrons so will have the same ending to their electron configuration.

10 A series of coloured lines on a black background corresponding to the wavelengths of light emitted when electrons transition between energy levels in an atom.

11 Bohr treated electrons as particles and their orbits as fixed radius positions around the nucleus. This allowed him to explain the emission spectrum of hydrogen. Schrödinger proposed electrons were waves and they occupied a volume of space around the nucleus called an orbital rather than Bohr›s fixed radius orbit. This allowed him to explain sublevels and orbitals and the emission spectra of multiple electron atoms.

12 18 protons, 22 neutrons, 18 electrons

13 a 29 protons, 34 neutrons, 29 electrons

b ${}^{65}_{29}\mathrm{Cu}$

14 a ${}^{241}_{95}\mathrm{Am} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{237}_{93}\mathrm{Np}$

b ${}^{137}_{55}\mathrm{Cs} \rightarrow {}^{0}_{-}e + {}^{137}_{56}\mathrm{Ba}$

15 R, V and Z in group 14, S and W in group 16, T and X in group 13,U and Y in group 2.

16 a $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$

b $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^2$

17 a berylium, magnesium, calcium

b carbon, silicon, germanium

c oxygen, sulfur, selenium

d zinc, cadmium, mercury

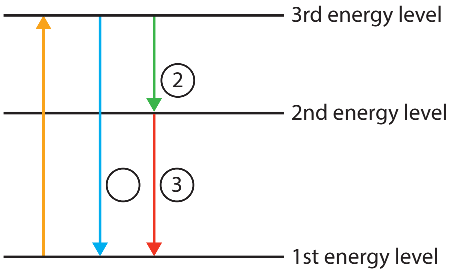
18 a 2,8,18,18,8 Noble gases in group 18 are already stable and unreactive as their outermost energy shell is already full of electrons.

b 2,8,18,18,7 Iodine has 7 valence electrons, same as chlorine and bromine, so will react chemically the same way and form bonds in the same way as they have the same number of valence electrons.

19 a W and Y

b U and Z (Group 15)

20 If a single electron was raised to the 3rd energy level then it could transition as shown (1) from 3rd to 1st level, or from (2) 3rd to 2nd then (3) from 2nd to 1st, each producing a different wavelength of light.



21 Alloy A – iron, vanadium; Alloy B – iron, cobalt, chromium; Alloy C – copper, nickel

STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 4: Periodicity

Check your understandng 41/42/43

1 Melting point increases across a period until group14 (carbon, silicon etc.) where it is at a maximum, then decreases until a minimum for the group 18 (oble gas) elements.

2 Right-hand side of the periodic table, mostly group 17 (halogens) and 18 (noble gases), and the top of group 15 and 16 (nitrogen and oxygen).

3 Group 2 elements – ionisation energies for the first and second electrons are relatively low, but the third ionisation energy is very high as this would be removing an electron from a stable ion with a noble gas configuraton. Group 3 elements – ionisation energies for first, second and third electrons are low, but the fourth ionisation energy is higher as this would remove an electron from a noble gas configuration io. Similar increases in energy are seen for any removal of an electron from a noble gas configuration, suggesting it is the most stablemfo of configuration.

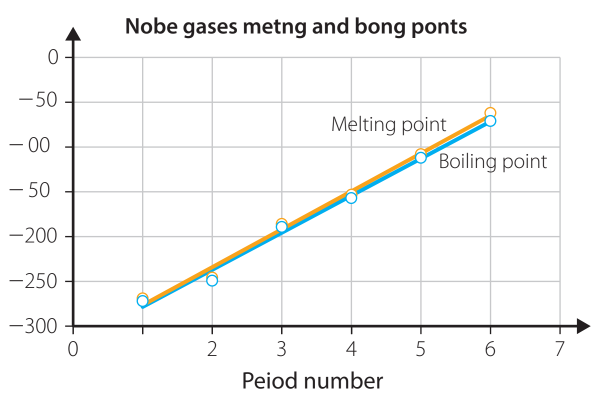
4 **a** Atomic radius decreases across a period as the force of attraction between the nucleus and the outermost shell increases as more lectrons are added. As the electrostatic attraction is stronger, it requires more energy to remove an electron from the outermost shell, thus first ionisation energy increases across aperiod.

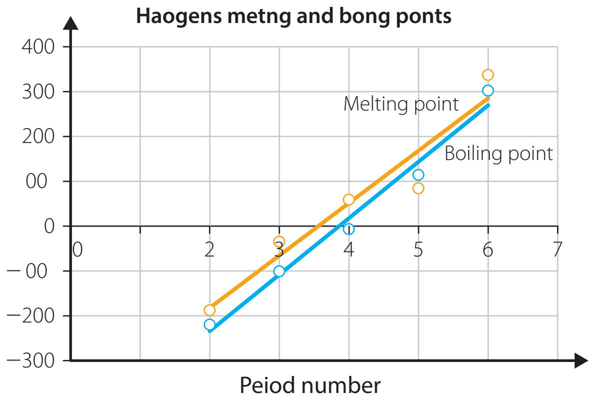
b It is easier to remove an electron from the outermost shell if it is further away from the nucleus and less subject to electrostatic forces of attraction. Thus first ionisation energy decreases as you go down a group.

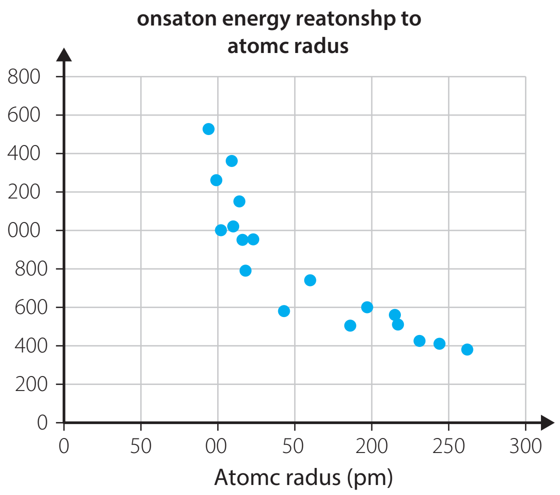
5 **a**

Halogen	Melting point (°C)	Boiling point (°C)	Noble gases	Melting point (°C)	Boiling point (°C)
Fluorine	−220	−188	Helium	−272	−269
Chlorine	−101	−35	Neon	−249	−246
Bromine	−7	588	Argon	−189	−186
Iodine	114	184	Krypton	−157	−153
Astatine	302	337	Xenon	−112	−108
			Radon	−71	−62

b







c In both groups for both melting and boiling points, there is an increase as period number increases.

6 **a, b**

Group 2 element	Melting point (°C)	Group 14 element	Melting point (°C)
Beryllium	1278	Carbon	3500
Magnesium	639	Silicon	1410
Calcium	839	Germanium	937
Strontium	769	Tin	232
Barium	725	Lead	327
Radium	700		

c For group 2 elements, there is a general decrease in melting point as you go down a group, with the exception of magnesium which has the lowest melting point despite being the 2nd element down. For group 14 there is also a general trend of the melting point decreasing as you go down the group, with the last two elements, tin and lead,being reversed.

7 Atomic number 18, group 18: this noble gas will have the highest ionisation energy of 1500 kJ mol^{-} as it is in the most stable configuration so is hardest to remove an electron from.

Atomic number 17, group 17: will have the next highest ionisation energy of 1300 kJ mol^{-} as ionisation energy increases across a group due to increased electrostatic attraction between the outer shell and the nucleus.

Atomic numbers 19and 20: will have the lowest (19) and second lowest (20) ionisation energies as they are on the left-hand side of the table and there is the lowest electrostatic attraction between the nucleus and outer shell.

8 Ionisation energy decreases going down the group so the element with the lowest atomic number will have ionisation energy of 530 kJ mol^{-} , then the next highest atomic number is 430 kJ mol^{-} , then the highest atomic number is 380 kJ mol^{-}

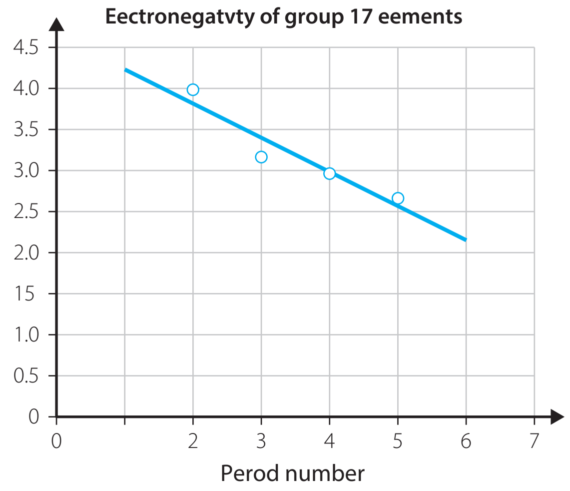
9 A group 1. The first ionisation energy is quite low, so the firstelectron is easy to remove. The higher ionisation energy of the second electron suggests it is being removed from the stable noble gas configuration, hence it has one valece electron.

B group 2. First and second ionisation energies are low, then a large jump to the third ionisation enery. Suggests two valence electrons that are relatively easy to remove, then the third electron is being removed from the stabl configuration.

C group 8. First ionisation energy is very high, so electron is being removed from the stable configuraton.

D group 1. For the same eason as A.

10 The graph below clearly shows the ionisation energy increasng as atomic radius decreases.



Check your understandng 44/45/46

1 **a** aluminium, boron, bromine, nitrogen, fluorine

b potassium, sodium, silicon, sulfur, chlorine

2 **a** Fluorine is a smaller atom than silicon as it has one less energy level occupied, so it is easier for fluorine to attract electrons to itself and form a negative ion.

b Oxygen is a smaller atom than phosphorus as it has one less energy level occupied, so attracts electrons more easily.

c Chlorine has two less energy levels occupied than iodine so is a much smaller atom, and attracts electrons moe easily.

3 vanadium (V), molybdenum (Mo), aluminium (Al)

4 **a** Group 13 and group 14

b Groups 1 and 2 are all metals, while group 17 are all non-metals hence no change in metallic character in those groups.

5 **a** Between group2 and 3. Group 2 = metal, group 3 = semi-metal

b Between group3 and 4. Group 3 = metal, Group 4 = semi-metal

c Between group6 and 7. Group 6 = metal, Group 7 = semi-metal

6 Lithium and barium

7 **a** Group 1 elements all react with water, with elements at the bottom reacting more vigorously than those at the top. Group 2 elements also increase in reactivity with water as you go down the group, except beryllium which does not react with water at all.

b In group 2, beryllium does not react with water, and magnesium only reacts with steam at high temperatures. Calcium, strontium and barium all react with cold water, with barium the most vigorous of the three, and calcium the least vigorous of the three.

Chapter review questions

1 Generally there is a trend in state from solid to gas from left to right on the periodic table. For groups 1–4, all elements are solid. For groups 15–17, there is a change from gas to solid as you go down a group. For group 18, all elements are gaseous.

2, 3 Atomic radius – increases down a group as more energy levels are occupied. Decreases across a period as there is stronger attraction between nucleus and electrons in outer shell.

First ionisation energy – decreases down a group as easier to remove electrons from energy levels further from the nucleus. Increases across a period as the atoms become smaller and the electrostatic attraction between nucleus and energy levels increases, making it harder to remove an electron from the outer shell.

Electronegativity – decreases down a group as energy levels get further away from the nucleus, and electrostatic forces are weaker so electrons are harder to attract. Increases across a period as outer shell is subject to stronger electrostatic attraction so electrons are more easily attracted by the atom.

4 Elements that react with water are found mostly in groups 1 and 2. Elements at the top of groups 1 and 2 react less vigorously, and reactivity with water increases as you go down a group.

5 First ionisation energy is the energy required to remove the first electron from the outermost shell. The lower this value, the easier it is to remove an electron. When an atom reacts it involves transfer or sharing of these electrons, so the easier they are to remove, the more reactive the element.

6 **a** magnesium, calcium, potassium

b fluorine, chlorine, silicon

c phosphorus, aluminium

d argon, chlorine, bromine, arsenic

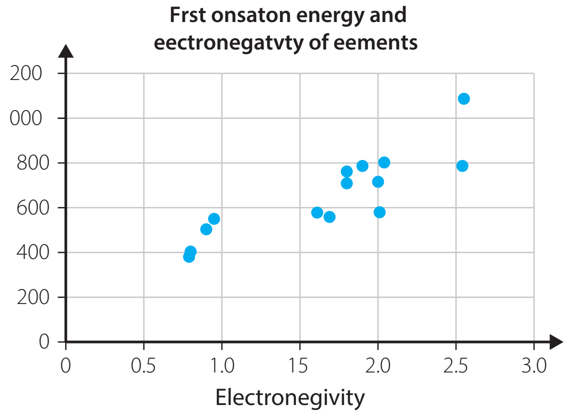
7 X – atomic radius = 230 pm, first ionisation energy = 425 kJ mol^{-}

Y – atomic radius = 160 pm, first ionisation energy = 740 kJ mol^{-}

Z – atomic radius = 100 pm, first ionisation energy = 1260 kJ mol^{-}

Atomic radius decreases across a period as electrostatic forces increase, attracting the outer shell closer to the nucleus as you go acros the period. First ionisation energy describes the ease of removing an electron from the outermost shell. For elements with large atomic radius it is easier to remove an electron as electrostatic forces are weaker. Hence first ionisation energy increases as atomic radius decreases, thus it increases aross a period.

8 From the graph, astatine (period 6) has an approximate electronegativity of 2.1–2.2.



9 **a** P (highest first ionisation energy)

b Q (lowest first ionisation energy)

c M (if P is group 18, then M is group 16)

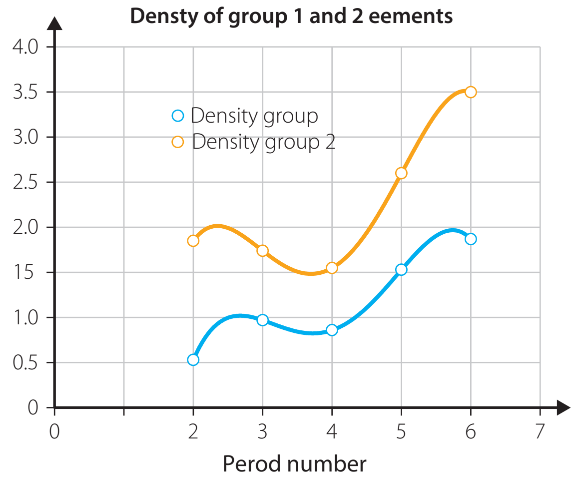
10 **a** barium, calcium, magnesium – first ionisation energy decreases down a group as the outermost electrons are further from the nucleus and require less energy to remove due to weaker electrostatic fores.

b aluminium, phosphorus, chlorine – first ionisation energy increases across a group as atomic radius decreases and electrostatic forces increase, making it more difficult to remove an electron.

c rubidium, sodium, aluminium, sulfur – first ionisation energy decreases down a group but increases across a group (as explained in **a** and **b**)

d barium, calcium, arsenic, nitrogen – first ionisation energy decreases down a group but increases across a group (as explained in **a** and **b**)

11 As seen in the graph, there is a general correlation between the electronegativity of an element and its first ionisatio energy. As one increases, th other also increases.



12 a While not perfectly consistent, there is a general increase in dnsity as you go down a group.

FPO
nnans-0406

b Francium = 21 g mL^{-} (appx.). Radium = 3.7 2.1 g mL^{-} (appx.). Accuracy of the estimates is poor as the trend is unclear and the curve could potentially goup or down after period 6 data.

c Include data from other groups in the periodic table to see if the trend continued across all groups.

STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 5: Chemical bonding

Check your understandng 51/52/53/54

1 **a** The noble gas electron configurations.

b Atoms can lose or gain electrons (ionic bonding) or share eectrons (covalent bonding).

2 **a** Atoms that have lost or gained electrons to become positively or negatively charged.

b Caton: positiveion, e.g. Mg^{2+} and Na^+

Anion: negativein, e.g. Cl^- and O^{2-}

3 Sodium atoms have 1 valence electron, so it loses an electron to form the stable electron configuration of neo. Sodium has one less electron than the atom so forms the Na^+ io.

Chlorine atoms have 7 valence electrons, so it gains an electron to form the stable electron configuration of aron. Chlorine has one extra electron than the atom so forms the Cl^- io.

4 **a** Electrostatic forces of attraction between the positive and negative ions formed through electron transfr.

b Ionic compounds consist of vast arrays of positive and negative ions arranged in an orderly fashion. This lattice has no fixed size, hence there are no separate molecules of an ionic compound.

5 The overall compound should be neutral, so the positive and negativecharges need to balance each other.

6 NH_4^+, OH^-, SO_4^{2-}, CO_3^{2-}, NO_3^-, PO_4^{3-}

7 **a** Potassium has electron configuration $1s^2\ 2s^2 2p^6\ 3s^2\ 3p^6\ 4s$ so has 1 valence elctron. Fluorine has configuration $1s^2\ 2s^2\ 2p^5$ so has 7 valence electrons. The potassium atom transfers its valence electron to the fluorine. K^+ and F^- ions form that are arranged in a lattice, held together by electrostatic attracion.

b Calcium has electron configuration $1s^2\ 2s^2 2p^6\ 3s^2\ 3p^6\ 4s^2$ so has 2 valence electrons. Bromine has configuration $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^5$ so has 7 valence electrons. The calcium atom transfers one valence electron to each of twobromine atoms. 1 × Ca^{2+} and 2 × Br^- ions form that are arranged in a lattice, held together by electrotatic attraction.

c Sodium has electron configuration $1s^2\ 2s^2 2p^6\ 3s$ so has 1 valence elctron. Oxygen has configuration $1s^2\ 2s^2\ 2p^4$ so has 6 valence electrons. Two sodium atoms transfer their one valence electron to the oxygn. 2 × Na^+ and 1 × O^{2-} ions form that are arranged in a lattice, held together by electrostatic attracion.

d Magnesium has electron configuration $1s^2\ 2s^2 2p^6\ 3s^2$ so has 2 valence electrons. Oxygen has configuration $1s^2\ 2s^2\ 2p^4$ so has 6 valence electrons. Magnesium transfers its two valence electrons to the oxygen aom. An Mg^{2+} and an O^{2-} ion form that are arranged in a lattice, held together by electrostatic attracion.

8 a KF

b $CaBr_2$

c Na_2O

d MgO

9 a Potassium has a charge of +1 as an ion, bromine is −1 as an ion, hence exist n a 1 : 1 ratio to create an overall neutral substance.

b Barium has a charge of +2 as an ion, chlorine is −1 as an ion, hence exist n a 1 : 2 ratio, thus the formula is $BaCl_2$

10 a Positiveions: Si, Rb, Mg, Ga, K, C B, Ba, Cs. egative ions I, S, r Se, P, F. No ion: Ar.

b Si = +4, Rb = +1, Mg = +2, Ga = +3, K = +1, C = +4, B = +3, Ba = +2, Cs = +1, I = −1, S = −2, Br = −1, Se = −2, P = −3

11 a magnesiumatom: $1s^2\,2s^2\,2p^6\,3s^2$, magnesiu ion: $1s^2\,2s^2\,2p^6$

ii sulfur atom: $1s^2\,2s^2\,3s^2\,3p^4$, sulfur on: $1s^2\,2s^2\,3s^2\,3p^6$

iii chlorineatom: $1s^2\,2s^2\,3s^2\,3p^5$, sulfur on: $1s^2\,2s^2\,3s^2\,3p^6$

v potassiu atom: $1s^2\,2s^2\,3s^2\,3p^6\,4s$, potassim ion: $1s^2\,2s^2\,3s^2\,3p^6$

b MgS

ii KCl

iii $MgCl_2$

v K_2O

12 a calcium hydride, lithium sulfide, aluminium bromide, zinc oxide, copper chloride, silver sulfide

b $NaNO_3$

ii Li_2CO_3

iii K_3PO_4

v $Fe(OH)_3$

v $Zn(NO_3)_2$

v $(NH_4)_2CO_3$

13 a zinc sulfate

b ammonium carbonate

c iron(II) hydroxide

d aluminium nitrate

e tin nitrate

f iron(III) sulfate

14 a +2 **ii** +1 **iii** +3 **v** +2 **v** +2 **v** +4 **v** +1 **v** +3

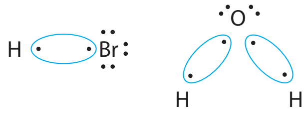
b −2 **ii** −2 **iii** −1 **v** −3 **v** −2 **v** −1 **v** −2 **v** −1

Check your understandng 55/56

1 Covalent bonds are formed between pairs of atoms due to the sharng of electrons between the atoms.

2 (Examples) methane, carbon dioxide, water

3



4 A pair of shared eectrons.

5 Molecular formulae show the number of each type of atom i a molecule of the substance. Used for covalent molecules. An example is CO_2 showing 1 atom of carbon and 2 atoms of oxygen in one carbon dioxide moleule. Empirical formulae show the ratio ofions in an ionic compound. NaCl is an empirical formula that only indicates 1 × Na^+ ion for every Cl^- io.

6 Groups 15–17 primarily as they contain non–metal atoms and atoms with 5, 6, and 7 valence electrons.

7 **a** covalent

b covalent and ionic

c covalent and ionic

d ionic

e covalent and ionic

f covalent

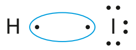
g ionic

h ionic and covalent

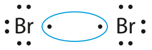
ionic

j covalent

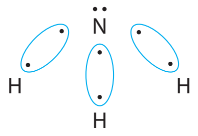
8 **a** HI



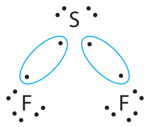
b Br_2



c NH_3



d SF_2



9 **a** Oxygen needs to share 2 valence electrons to fill its outer shell, fluorine only needs to share 1 electron as it has 7 valence elctrons. Two fluorine atoms each share 1 electron with one oxygen atom so all are stable, thus F_2O

b Nitrogen has 5 valence electrons so needs to share 3 electrons to become stable. Hydrogen only has 1 valence electron to share, so three hydrogen atoms are needed for each atom of nitrogen, thus NH_3

10 a SCl_2 – sulfur has 6 valence electrons so needs to share 2electrons for stability. Chlorine has 7 valence electrons so needs to share 1 electron. Thus two chlorine atoms are needed for one sulfur atom for all to complete their outer shell.

b HI – hydrogen has 1 valence electron so needs to share 1 electron to fill the outer shell. Iodine has 7 valence electrons so needs to share 1 electron. So one hydrogen atom is needed for each iodine atom.

c SiH_4 – silicon has 4 valence electrons so needs to share 4 electrons to fill the outer shell. Hydrogen has 1 valence electron so needs to share 1 electron. So four hydrogen atoms are needed for each silicon atom.

d PF_3 – phosphorus has 5 valence electrons so needs to share 3 elctrons to fill the outer shell. Fluorine has 7 valence electrons so needs toshare 1 electron. So three fluorine atoms are needed for each phosphorus atom.

Check your understandng 57/58/59

1 The valency of an element measures the combining power of the element when it forms compounds.

2 a In ionic compounds, for groups 1 and 2 it is the group number, for group 13 it is the group number minus 10 ($13 - 10 = 3$), for groups 14–17 it is 18 minus the roup number.

b In covalent compounds, it indicates the number of covalentbonds the element forms.

3 For any element in groups 1 and 2 it is group number. For example, group 1 has a valency of 1, and group 2 has a valency of 2. For group 13 it is the group number minus 10 ($13 - 10 = 3$) For groups 14–17 it is 18 minus the group number. For example, group 14 is ($18 - 14$) = 4.

4 The overall charge on the covalent compound is neutral, so positive nd negative charges must balance.

5 First named element is the one with the lower electronegativity.

6 a strontium = +2, arsenic = −3, selenium = −2, indium = +3, bromine = −1, caesium = +1, boron = +3 barium = +2 antimony = −3 germanium = +4, gallium = +3, iodine = −1

b strontium, indium, caesium, barium, gallium

ii germaium, boron

iii arsenic, selenium, bromine, antimony, iodine

7 a Rb_2O, CsCl, Li_2SO_4 $RbNO_3$, Cs_2SO_4, Li_2S

b $Ba(NO_3)_2$, RaO, SrS $BeSO_4$, $RaCl_2$, $SrBr_2$, BaF_2

c Cs_2O, BaO, Tl_2O_3, PbO_2, Bi_2O_3

8 a PH_3, AsH_3, SbH_3

b SeH_2 – covalent as electrons are shared between them to acieve a stable outer shell. TeH_2 – same reason.

9 a LiI, NaI, KI, RbI, CsI – all inic.

b BeO, MgO, CaO, SrO, BaO – all ionic.

c MgF_2, $MgCl_2$, $MgBr_2$, MgI_2 – all ioic.

d CCl_4, $SiCl_4$, $GeCl_4$ – covalent; $SnCl_4$, $PbCl_4$ – ioni.

e H_2O, H_2S, H_2Se, H_2Te – all covalent.

10 a SF_2 **b** NCl_3 **c** CI_4 **d** BCl_3 **e** SO_2 **f** $SiCl_4$ **g** PBr_5 **h** N_2O_3

11 a silicon tetrachloride, carbon disulfide, dinitrogen oxide, nitrogen trichloride, dihydrogen sulfide, phosphorus tribromide, silicon carbide

b diphosphorus pentoxide, dinitrogen tetroxide, dichlorine monoxide, difluorine monoxide, dichlorine trioxide, sulfur trioxide, disulfur difluoride

Check your understandng 5.10/5.11/5.12/5.13

1 They are hard, brittle and have a igh melting point.

2 Ionic compounds have high melting points due to the strong electrostatic forces of attraction between the positive and negative ion in the lattice. High amounts of energy are needed to break these bonds, so the melting poit is high. Covalent compounds have very weak forces of attraction between the molecules which are broken with low energy, so covalent compounds have very low melting points.

3 Covalent molecules are neutral so cannotconduct electricity.

4 Ionic solids have ions that are fixed in place in the lattice, thus cannot conduct electricity even though they are charge particles. When the solid is melted, the ions separate and are free to move. Free-moving charged particles conduct electricity so a molten ionic compound conducts electricity where a solid does no.

5 Covalent network solids have covalent bonding extending throughout a lattice structure such as diamond or gaphite. Covalent molecules are a defined set of atoms that form a distinct structure with a fixed number of atoms such as carbon doxide or methane.

6 They are hard due to strong covalent bonds between atoms, have high melting points due to the high energy required to break the covalent bonds between atoms, and except for graphite they do not conduct electricity as they have no free-movng charged particles.

7 Diamond is made of carbon atoms, each covalently bonded to four other carbon atoms in a tetrahedral structure. Silica is silicon dioxide, where a silicon atom is bonded to four oxygen atoms in a tetrahedral structure.

8 a Metal atoms lose their valence electrons t form positive ions. The positive ions form an ordered lattice of positve ions. The 'lost' electrons form a sea of delocalised or unattached electrons that surround the lattice of positive ions.

b The free-moving charged delocalised electrons are able to move throughout the metallic lattice and conduct electricity.

9 Molecular solids, metallic solids, ionic solids, covalent lattice solids.

10 Phosphorus trichloride is a covalent molecule so has weak forces of attraction between molecules so low energy is needed to break the bonds, resulting in a low melting point. Covalent molecules are neutral so will not conduct eectricity. Calcium chloride is ionic so has strong electrostatic forces between ions so a high melting point is needed to break these bonds and separate the ions. In solid form, the ions are fixed so do not conduct electricity until the ions are separated (melted) so they are free to move. Free-moving ions conduct electricity.

11 Lithium fluoride and beryllium fluoride are ionic compounds with strong electrostatic forces between ions, and all the others are covalent molecules with weak forces between molecules. The ionic substances have higher melting points as the energy required to break the bonds between ions is higher than the energy needed to break the weak forcs between molecules.

12 a Calcium hydride is ionic as a high melting point means high energy is required to overcome strong electrostatc forces. Arsine is a gas, thus its boiling point is low, so it has weak bonds and is covalent moleclar.

b CaH_2 and AsH_3

13 Silicon dioxide is a covalent network substance with covalent bonding between atoms throughout the substance, making it hard with a vey high melting point. Carbon dioxide is a covalent molecular substance with very weak bonds between molecules that are easily broken at room temperature, making it a gas.

14 a B, D, F

ii A, G

iii C, E

b Metals – relatively high melting points (strong bonds within lattice) and good conductivity (delocalised electrons).

Covalent molecules – low melting point (weak intermolecular bonds) and low conductivity (neutral molecules).

Covalent lattice – very high melting point (covalent bonds throughout structure) and low conductivity (no free-moving particle to conduct).

Chapter reiew quesions

1 Ionic bonding occurs whenametal (e.g. sodium) transfers its valence ectrons to a non-metal (e.g. chlorine), forming a positive metal ion (Na^+) and a negative non-metal ion (Cl^-) The ions form a lattice (NaCl – sodium chloride) with strong electrostatc forces between the ions. Covalent bonding occurs when non-metals share atoms to form a stabe electron configuration. Hydrogen shares its one valence electron with chlorine so each fills its outermost shellforming hydrogen chloride, HCl.

2 a K^+ and O^{2-}

ii Al^{3+} and S^{2-}

b potassium oxide, K_2O

ii aluminium sulfide, Al_2S_3

c They are both empirical formulae and show the ratio of ions in the ionic lattice.

3 a iron(III) oxide, tin sulfide, lead(II) oxide, copper(I) chloride, iron(II) bromide

b zinc nitrate, $Zn(NO_3)_2$

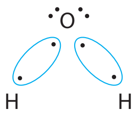
ii potassium carbonate, K_2CO_3

iii copper hydroxide, $Cu(OH)_2$

v ammonium sulfate, $(NH_4)_2SO_4$

v aluminium sufate, $Al_2(SO_4)_3$

4 **a**



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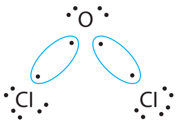


5 **a** ionic

b ionic

c ionic and covalent

d ionic and covalent

e ionic and covalent

6 **a** Covalent molecular substances have low melting and boiling points and are often found as gases or liquids at room temperature, while ionic substances are hard, brittl solids with high melting points. Covalent molecules do not condut electricity. Ionic solids do not conduct electricity but molten ionic substances do conduct lectricity.

b The hardness and high melting point of ionic solids are due to strong electrostatic forces between ions, while weaker intermolecular forces give low melting ponts of the covalent molecules. Ionic solids have charged particles fixed in place so do not conduct, but when molten the ions separate so can move freely and onduct. Covalent molecules are neutral sodo not conduct electricity.

7 **a** For diagram see igure 5.10page 108. Atoms are covalently bonded to other atoms to create a large lattice structure where all atom are fixed in place. Diamond contains carbon atoms covalently bonded to four other carbon atoms in a terahedral structure.

b For diagram see Fgure 5.12 age 109. Metal atoms lose their electrons to form positive ions that are arranged in an ordery lattice. The lattice is surrounded by the lost electrons to form a sea of delocalised electons. Magnesium loses two valence electrons and forms a lattice of Mg^{2+} ions surrounded by its 'los' elecrons.

8

Property	Ionic compounds	Covalent molecular compounds	Covalent lattices	Metals
Physical state	solid	gas/liquid	solid	solid
Melting point	variable	low	very high	variable
Conductivity	high (molten) low (solid)	low	low (except graphite)	very high
Brittleness	high	low	low	low
Malleability	low	low	low	very high

9 **a** H

b

10 a Na_2S, ionic

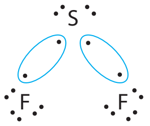
b CS_2 covalent

c SO_2 or SO_3 covalent

d SF_2 covalent

e Al_2S_3, ionic

11 a



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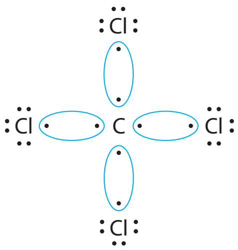

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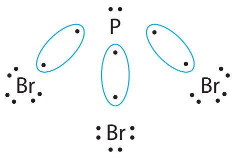

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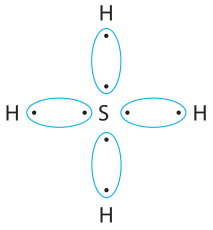


b sulfur difluoride, SF_2

ii carbon tetracloride, CCl_4

iii phosphorus tibromide, PBr_3

v silicon tetrahdride, SiH_4

c All covalent moecular.

d The number of each type of atom inone molecule.

12 Covalent; because the difference in electronegativities is less than 1.5 (B = 204, Cl = 316) However, it is not easy to decie upon. Based on the tendency to lose or gain electrons, it is easy to argue that boron tends to lose 3 electrons to achieve the helium configuration while chlorine tends to gain 1 electron to achieve the argon configuration, so $B^{3+}(Cl^{-})_3$ seems not unreasonable. In addition, the electron dot structure for covalent BCl_3 has only 6 electrons in the valence shell of B istead of the usual 8. The deciding factor is that BCl_3 is a gas at room temperature and also has a vey low melting point. These facts imply that it is covalen, not ionic. Boron is one of a few elements that form stable compounds with other than 8 electrons in theirvalence shells. This was not meant to be a simple question!

13 Silicon and carbon both form 4 covalent bonds, so carbon would form 4 covalent bonds with 4 silicon atoms, while each silicon atom would form 4 covalent bonds with arbon so the atoms alternate. They form a large lattice full ofcovalent bonds. Covalent lattice structures are strong and hard, thus used as an abraive.

14 a, b

L and Q – conducts as both solid and liquid, malleable

ii P and R – conducts as liquid but not solid, reasonably high melting points, hard substances

iii M – does not conduct at all, low melting point and a soft substance

v N – does not conduct at all, high melting point and extremely hard

STUDENT BOOK ANSWERS

Module one: Properties and structure of matter

Chapter 6: Intermolecular forces and allotropy

Check your understandng 61/62/63

1 Covalent bonds in which electrons are unequally shared.

2 A non-polar molecule has o net dipole. If the polar covalent bonds within the molecule cancel out to make a net zero dipole, then the moleculewill be non-polar.

3 Electronegativity of the element.

4 Polar: C—H, N—H, Cl—O

Non-polar: H—H, Cl—Cl, F—F or N—Cl, C—S, P—H

5 **a** Pairs of valence electrons around an atom in a molecule arrange themselves so they are at the furthest possible distance fro each other.

b The number of pairs of valence electrons that are repelling give the basic shape of a molecule as it determines where the atoms are spatially when surrounding the central element.

6 **a** Methane has four pairs of valence electrons surrounding the central carbon atom, which repel to maximum distance giving a tetrahedral shape to the molecule, with a bond angle of 109.5°

b Ammonia also has four pairs of valence electrons surrounding the central nitrogen atom, but only three of these pairs involve sharing with another atom (H) and one pair is a lone pair of electrons. The bonding pairs give the final shape hich is pyramidal.

7 $^{\delta-}\mathrm{F{—}O}^{\delta+}$, $^{\delta+}\mathrm{Cl{—}O}^{\delta-}$, N—Cl, $^{\delta-}\mathrm{N{—}H}^{\delta+}$, $^{\delta-}\mathrm{S{—}H}^{\delta+}$, C—S, $^{\delta+}\mathrm{C{—}O}^{\delta-}$, $^{\delta+}\mathrm{P{—}O}^{\delta-}$, $^{\delta+}\mathrm{P{—}Cl}^{\delta-}$, $^{\delta-}\mathrm{C{—}H}^{\delta+}$, $^{\delta+}\mathrm{C{—}Cl}^{\delta-}$

8 **a** pyramidal

ii tetrahedral

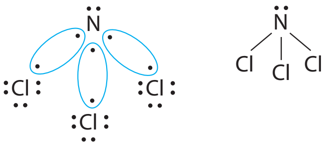
iii bent

v bent

v tetrahedral

v pyramidal

b



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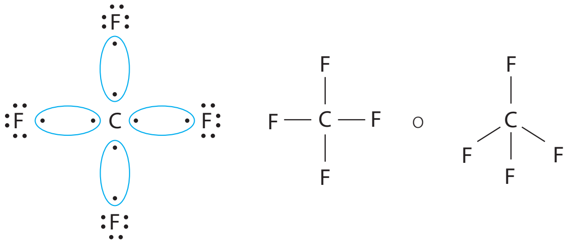

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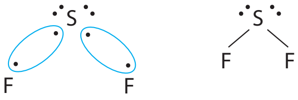

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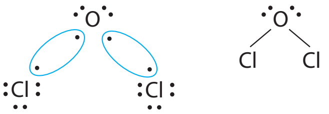

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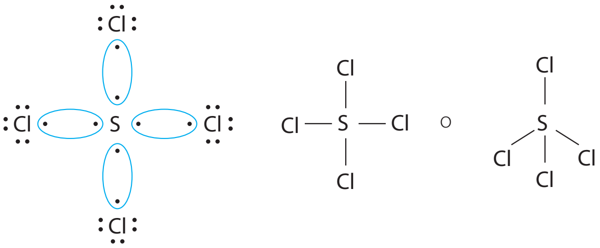

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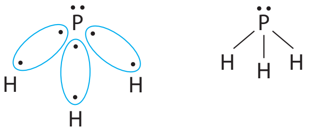


9 **a** none

ii $^{\delta+}\text{C—F}^{\delta-}$

iii $^{\delta+}\text{S—F}^{\delta-}$

v $^{\delta+}\text{Cl—O}^{\delta-}$

v $^{\delta+}\text{Si—Cl}^{\delta-}$

v none

b NCl_3, SF_2, Cl_2O, PH_3

10 **a** no

ii yes

iii yes

v yes

v yes

v yes

b non-polar: no dipole charges

ii polar: has an overall net dipole due to bent shape

iii polar: has overall net dipole as different outer atoms

v polar: overall net dipole due to bent shape

v non-polar: no overall net dipole

v polar: overall net dipole due to pyramidal shape

Check your understandng 64

1 Intramolecular forces involve the covalent bonds between atoms within a molecule; for example, the shared pair of electrons between the H and Cl atoms in hydrogen chloride. Intermolecular forces exist between molecules; for example, betwen molecules of HCl.

2 Physical state (solid, liquid or gas), and melting/boiling points.

3 **a** Dipole–dipole forces, dispersion forces and hydrogen bonds.

b Dispersion forces, dipole–dipole force, hydrogen bonds.

4 Polar molecules have positive and negative ends du to the polar bonds formed. (Example) A molecule of HCl has a slightly positive H and a slghtly negative Cl. Dipole–dipole forces are forces of attraction between the positive H on one molecule and the negative Cl on a different molecule.

5 **a** Dispersion forces form between non-polar molecules due to random motion of electrons causing temporary dipoles as the electron cloud becomes unsymmetricall distributed around the molecule. The temporary dipoles cause attraction between the positive dipole on one molecule and the negative dipole on a different molecule.

b The number of electrons in he molecule. The more electrons in the molecule, the stronger the dispersion forces

6 Hydrogen bonding occurs when molecules have H—O, H—N or H—F pola bonds within the molecule. Due to the large electronegativity differences in these bonds, the positive and negative charges on the molecules are stronger than inother molecules. The resulting attraction between the positive H in one molecule and the negative O, N or F on a different molecule is a hydrogen bond.

7 The boiling points of molecules such as H_2O (water) and NH_3 (ammonia) are higher than is expected based on their size and the trends see in their groups. This is due to them forming hydrogen bonds that are significantly stronger than dipole–dipole bonds ormed by similar molecules.

8 Hydrogen bonding: NH_3, HF, H_2O

Non-hydrogen-boning HCl, HBr, H_2S

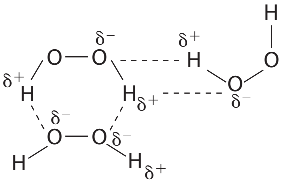
9 **a** Helium, neon and hydrogen have very weak forces between molecules as only a very small amount of energy is required to break these bonds. This is see in the low boiling point. The other gases have stronger intermolecular forces as they have higher boiling points.

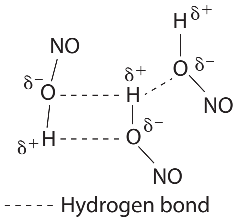
b Helium, neon and hydrogen have dispersion forces between molecules, the other molecules have dipole–dipole bonds between molecules.

10 Hydrogen fluoride has hydrogen bonding between molecules while neon has dispersion forces between atoms. Hydrogen bonds are much stronger than dispersion forces so require more energy to break, and thus hydrogen fluoride has a much higher boilng point than neon.

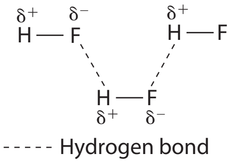
11 Chlorine gas, trichloromethan,sodium chloride. Chlorine gas is a non-polar molecule with no dipoles so has weak dispersion forces beween molecules. Trichloromethane is a polar molecule with stronger dipole–dipole forces between molecules so has a higher boilig point than chlorine gas. Sodium chloride is ionic so has significantly stronger electrostatic forces of attraction between ions giving it a much higher boiling point than the other two molecules.

12 Hydrogen peroxide and nitrous acid

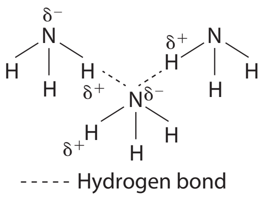




13 a



b



Check your understandng 65/66

1 Forms of the same element in the same physical state with different properties such as colour, density, hardness or conductivity.

2 Carbon, phosphorus, sulfur and oxygen

3 a Diamond is a hard, non-conductive, colourless solid while graphite is a black, soft, conductive substance that can be usedas a lubricant.

b Diamond has carbon atoms that are each covalently bonded to four other carbon atoms in a tetrahedral structure, giving a hard structure with no free-moving charged particles so it does not conduct electrcity. Graphite has each of its carbon atoms bonded to only three other carbon atoms in a planar strucure. Flat sheets of carbon atoms form that move over eachother, making it soft and slippery. The unbonded fourth electron is free to move, so allow for electricity conduction.

4 **a** Both are elemental phosphorus, but have different reactivity and colour.

b White phosphorus has P_4 molecules packed into a crystal structure. In each molecule, the P atoms are covalently bonded to four other P atoms. Red phosphorus is a chain of P_4 units where the P_4 units are covalently bonded to other P_4 unit.

5 The intermolecular forces are broken at room temperature, so are very weak.

6 **a** Forces between ozone molecules are stronger thanbetween oxygen molecules.

b Ozone has more electrons than oxygen as it contains an extra atom, so has slightly stronger dispersion forces due to its extra electrons.

7 The structure of the two phosphorus allotropes is very different. One is packed molecules in a crystal, and the other is a chain, so their properties also very different. The rhombic and monoclinic allotropes of sulfur are not very different in their structure, only the arrangement of molecules in the crystals, so their physical propertiesare quite similar.

8 All have dispersion forces between molecules which increase in strength as the number of electrons in the molecule increases (hence molecule size). Thus in order of increasing size they are H_2, O_2, Cl_2 and Br_2 Thus their boiling points (respectively) are −253 °C, −188°C, −35°C and 59°C

9 **a** Graphite has sheets of planar bonded carbon atoms with only weak forces between the planar sheets. They are easily broken so can be slid off easily by friction, thus used as a lubricant.

b Buckminsterfullerene, like graphite, contains separate structures that are only weakly bonded together so can be pulled apart through friction, thus also useful as a lubricant.

10 Elements that are gases are all non-polar molecules or atoms i the case of the noble gases. As there are no permanent dipoles in the molecule, it is only random movement of the electron cloud creating temporary dipoles that form forcesbetween molecule. This is dispersion forces.

Chapter review questions

1 Carbon and chlorine have a significant electronegativity difference so a polar bond forms. The electronegativity difference between carbon and sulfur is small to non-existent, so this bond is non-polar.

2 **a** linear

b planar

c tetrahedral

3 **a** planar

b bent

c tetrahedral

d pyramidal

4 **a** B—Cl bond is polar, but planar shape means no overall net dipole so molecule is non-polar.

b H—O bond is polar, bent shape means a net dipole forms so molecule is polar.

c C—H bond is slightly polar, but tetrahedral shape means no overall net dipole so molecule is non-polar.

d N—H bond is polar, pyramidal shape means a net dipole forms so the molecule is polar.

5 **a** hydrogen, methane, oxygen

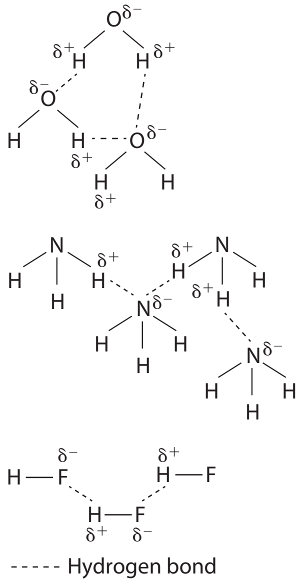
b hydrogen chloride, phosphorus trichloride, hydrogen sulfide

c water, ammonia, hydrogen fluoride

6 **a** nitrogen, oxygen, fluorine

b water, ammonia, hydrogen fluoride

c



7 **a** intramolecular – covalent; intermolecular – dispersion forces

b intramolecular – none, in atomic form; intermolecular – dispersion forces

c as an element does not form molecules – has covalent bonding between atoms but forms a lattice

d as an element does not form molecules – has metallic bonding

e for white phosphorus: intramolecular – covalent; intermolecular – molecules packed into crystalline form

8 fluorine monoxide, tetrafluoromethane, sulfur dichloride and dibromomethane

9 **a** polar: molecule has a net dipole due to its bent shape

b non-polar: molecule has no net dipole due to its linear shape

c non-polar: no net dipole due to tetrahedral shape

d polar: net dipole exists due to bent shape

e polar: net dipole exists as different atoms surround the central atom creating a net dipole

f polar: net dipole exists due to pyramidal shape

10 **a** ammonia – pyramidal, ammonium ion – tetrahedral

b boron trifluoride – planar, boron tetrafluoride ion – tetrahedral

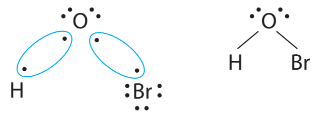
c pyramidal

11 Sulfur dioxide must be bent as a polar molecule must have a net dipole, which cannot form with a linear molecule. Sulfur trioxide must be trigonal planar as it is non-polar and thus has no net dipole, so cannot be pyramidal or a net dipole would exist.

12 a Weak: Ne, O_2, Ar. Moderate: H_2S, CO_2, HCl. Strong: HF, NH_3

b Weak = dispersion forces. Moderate = dipole–dipole forces. Strong = hydrogen bonds.

13 a

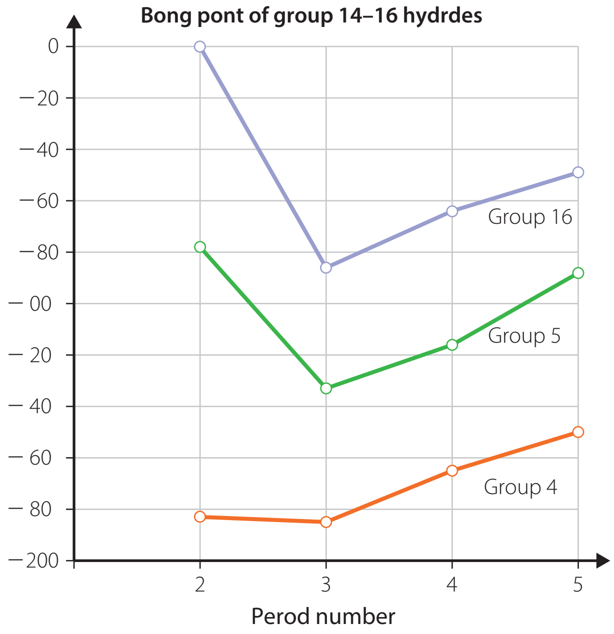


b bent

c dipole–dipole forces and hydrogen bonding

d hydrogen bonding

14 The graph shows evidence of hydrogen bonding for the compounds at the top of group 15 and 16 due to their abnormally high boiling point when considering the rest of the group trend. Thus H_2O and NH_3 have hydrogen bonding between molecules.



15 a Methane is a non-polar molecule with weak dispersion forces between molecules. Chloromethane is a polar molecule with stronger dipole–dipole forces between molecules. Dipole–dipole forces require more energy to break than dispersion forces, so chloromethane has a higher boiling point than methane.

b Sulfur hexafluoride is a non-polar covalent molecule with weak intermolecular forces. Lead fluoride is an ionic substance with very strong electrostatic forces of attraction between the ions in its lattice. More energy is needed to break ionic bonds than intermolecular bonds, so lead fluoride has a much higher boiling point than sulfur hexafluoride.

c Fluorine and bromine are both non-polar molecules with dispersion forces between molecules. Dispersion forces are stronger when more electrons are in the molecule, thus bromine has stronger dispersion forces and a higher boiling point.

d Zinc chloride and tin(IV) chloride are both ionic compounds. Ionic compounds vary in the strength of electrostatic forces between ions due to factor such as size of the ion. In this case, the tin(IV) ion is much larger than the zinc ion and will affect the strength of the electrostatic forces.

e Tetrafluoromethane is a non-polar molecule with dispersion forces between molecules, while difluoromethane is a polar molecule with a net dipole as it has different atoms surrounding the central carbon tom. It thus has dipole–dipole forces between molecules, which are stronger than dispersion forces and require more energy to break, so difluorometha has a higher boiling point.

16 Sulfur hexafluoride is a non-polar molecule as its symmetrical shpe means no net dipole exists. It has dispersion forces between molecules, and as these are the weakest intermolecular bonds it has the lowest boiling point. Sulfur dichloride has a bent shape so will have a net dipole and exhibit dipole–dipole forces between mlecules. These forces are stronger than dispersion forces so it will have a higher boiling point than sulfur hxafluoride. Sulfuric acid contains O—H bonds within the molecule so will form hydrogen bonds betwen molecules. This is the strongest type of intermolecular bonding so sulfuric acid will have the highest biling point.

17 Boron trifluoride is a planar molecule so is non–polar as it has no net dipole. It will have dispersion forces between molecules, which are weak and easily broken at room temperature, so it exists as a gas. Boric acid has O—H bonds within the molecule so forms hydroen bonds between molecules. These bonds are very strong and will not be broken at room temperature, thus boric acid is a solid.

End-of-modue reiew

1 **a** 1 = filtration, 2 = distillation

b Filtration – refer toFigure 24 page 25. Distillation – refr to Figure 2.6 page 26.

c barium sulfate

2 **a** All of them – they all revert to the stable electron configuration of neon when forming ions.

b None of hem.

3 **a** Distillation, where the mixture of liquids is heated, the liquid with the lowest boiling point evaporates, is then condensed in the condenser and collected, leaving behind the liquid with the higher boiling point (for diagram see Fiure 2.6 page 26).

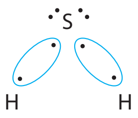
b When boiling a liquid that is pure, it will boil at a articular temperature. If it is impure, it will boil over a range of temperatures due to the different ompounds present.

c Fractional distillation, which uses multiple successive distillations to collect a pure sample when the boiling points of the two liquds are close.

4 **a** hydrogen (1), sodium (2,8,1), sulfur (2,8,6)

b Na_2S

c H_2S



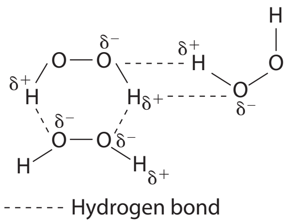
5 **a** A = metallic, B = ionic, C = covalent molecular, D = covalent molecular, E = covalent lattice

b Sodium chloride does not conduct as a solid as its ions are fixed in place and are not free to move, which is required for electrcal conduction. When melted, the ions separate out of the lattice structure and are free to move, so can now conduct electricity. Copper has delocalised electrons within its structure as a solid or when molten, so always has free moving charged particles to conduct electricity.

c Covalent molecular and covalent lattice substances have no free-moving charged particles (except graphite), so they do not conduct electricity asa solid or as a liquid. Covalent molecular substances have very weak intermolecular forces, so have low melting points as low energy is needed to break these bnds. Covalent lattices have covalent bonds between each atom, which are extremely strong and require a lot of energy to break, so they have significantly higher melting points.

6 For group 1–2 and group 13–17 elements i is very useful. Group 1 and 2 elements have a valency the same as their grop number. Group 13 element have a valence of 3. For group 14–17 elements, valency can be found by subtracting 18 fromthe group number. For example, group 16 elements have valency 16 – 18 = –2

7 **a** Hydrogen peroxide



b tetrafluoromethane

ii hydrogen sulfide

8 Non-polar molecules have weak dispersion forces between molecules, while polar molecules have stronger dipole–dipole and hydrogen bondng between molecules. Non-polar molecules have particular shapes (tetrahedral, trigonal planar) as these usuallyresult in no net dipole forming. Polar molecules have different shapes (bent, pyramidal) as these usually result in a net dipole. Hence shape can generally be used to determine polarity and hence the type of intermolecular forces. Note: there are exceptions; e.. CH_2Cl_2 is a tetrahedral molecule that is non-polar, so this is not a hard and fast rule.

Bond polarity is generally less helpful in determining the type of intermolecular bonding as non-polar molecules can contain polar bonds (CH_4) or non-polar bonds (CI_4) Polar molecules can contain polar bonds (H_2S) or non-polar bonds (PH_3)

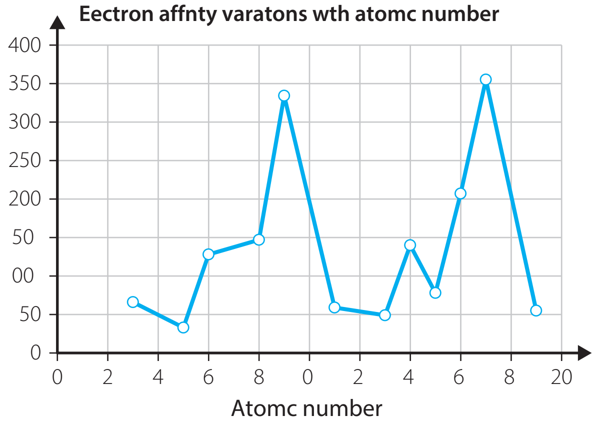
9 **a** $\frac{5\ 64-4\ 76}{5\ 64}\times 100=15\ 6\%$

b $\frac{8\ 09-6\ 82}{8\ 09}\times 100=15\ 7\%$ This nearly identical percentage suggests that the setting process involves the absorption of water to reform a solid compond similar to gypsum. The slight difference would be due to measurement and experiental error.

10 **a** A = argon, B = aluminium, C = chlorine

b argon = 94 pm, aluminium = 143 pm, chlorine = 99 pm

11 a



b The regular increase then decrease suggests there is periodic variation (varies in a systematic manner). The peaks occur at element 9 and 17 (corresponding with group 17 elements), hence verifying the systematic variation.

c F and Cl have a very small atomic radius due to increased electrostatic attraction between the nucleus and the outer shel electrons. Thus attracting an electron into the outer shell is easier as the outer shell is closer to the nucleus and subject to stronger forces of attraction. Li and Na have the largest atomic radius being in group 1, so electrons are less attracted to the nuclus than the group 17 elements.

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 7: Chemical reactions, stoichiometry and the mole

Worked example 7.1

$Fe_2O_3(s) + 3H_2SO_4(aq) \rightarrow Fe_2(SO_4)_3(aq) + 3H_2O(l)$

Check your understanding 7.1/7.2

1 **a** $MgO(s) + 2HCl\ (aq) \rightarrow MgCl_2(aq) + H_2O(l)$

b $2KBr(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + Br_2(aq)$

c $2O_3(g) \rightarrow 3O_2(g)$

d $FeS(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2S(g)$

e $4NH_3(aq) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$

f $Zn(NO_3)_2(aq) + 2NaOH(aq) \rightarrow Zn(OH)_2(aq) + 2NaNO_3(aq)$

g $Na_2CO_3(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + CO_2(g) + H_2O(l)$

h $Cu(s) + 2H_2SO_4(aq) \rightarrow CuSO_4(aq) + SO_2(g) + 2H_2O(l)$

I $Al_2O_3(s) + 6HNO_3(aq) \rightarrow 2Al(NO_3)_3(aq) + 3H_2O(l)$

j $MnO_2(s) + 4HCl(aq) \rightarrow MnCl_2(aq) + Cl_2(g) + 2H_2O(l)$

2 **a** **i** 2Al, 3S, 18O, 6Na, 6H → 2Al, 3S, 18O, 6Na, 6H

ii 8C, 20H, 18O → 8C, 26O, 20H

b **i** Balanced, as equal numbers of each atom on both sides.

ii Not balanced, as oxygen is not equal on both sides

3 **a** Ethane + oxygen → carbon dioxide + water

$2C_2H_6(g) + 7O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

b Dinitrogen pentoxide + water → nitric acid

$N_2O_5(g) + H_2O(l) \rightarrow 2HNO_3(l)$

4 **a** butene + oxygen → carbon dioxide + water

b iron(III) oxide + carbon monoxide → iron + carbon dioxide

Worked example 7.2

$$\text{Average mass} = \frac{(69 \times 63) + (31 \times 65)}{100}$$

Relative atomic mass of copper is 63.62.

Worked example 7.3

$M = (9 \times 12.01) + (8 \times 1.008) + (4 \times 16.00) = 180.154$

Worked example 7.4

$M = (2 \times 26.98) + (3 \times 32.07) + (12 \times 16.00) = 342.17$

Check your understanding 7.3/7.4/7.5/7.6

1 **a** Both compounds are ionic so do not form molecules. Relative molecular mass is the weight of one molecule compared to a carbon atom, and only applies to covalent molecular compounds.

b ‘Formula mass’ is used for ionic compounds.

2 $7 \times 12.01 = 84.07$

3 $\frac{96}{12.01} = 8$, thus molybdenum is 8 times heavier than carbon.

4 A calcium atom is twice as heavy as a neon atom.

5 $\text{Average mass} = \frac{(60 \times 69) + (40 \times 71)}{100} = 69.8$

6 $\text{Average mass} = \frac{(80 \times 11) + (20 \times 10)}{100} = 10.8$

7 $\text{Average mass} = \frac{(78.7 \times 24) + (10.1 \times 25) + (11.2 \times 26)}{100} = 24.3$

8 **a** $M = 39.10 + 79.90 = 119.00$

b $M = 12.01 + (2 \times 16.00) = 44.01$

c $M = 24.31 + (2 \times 35.45) = 95.21$

d $M = (2 \times 22.99) + 32.07 = 78.05$

e $M = (2 \times 55.85) + (3 \times 16.00) = 159.7$

f $M = (2 \times 30.97) + (5 \times 16.00) = 141.94$

g $M = 40.08 + 32.07 + (4 \times 16.00) = 136.15$

h $M = (2 \times 22.99) + 12.01 + (3 \times 16.00) = 105.99$

i $M = (2 \times 14.01) + (4 \times 1.008) + (3 \times 16.00) = 80.052$

j $M = (6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00) = 180.156$

k $M = 40.08 + (2 \times 16.00) + (2 \times 1.008) = 74.096$

l $M = 63.55 + (2 \times 14.01) + (6 \times 16.00) = 187.57$

m $M = (2 \times 14.01) + (8 \times 1.008) + 32.07 + (4 \times 16.00) = 132.154$

n $M = 55.85 + (3 \times 14.01) + (9 \times 16.00) = 241.88$

o $M = (2 \times 22.99) + 12.01 + (13 \times 16.00) + (20 \times 1.008) = 286.15$

Worked example 7.5

1 $MM\ (Na_2SO_4) = (2 \times 22.99) + 32.07 + (4 \times 16.00) = 142.05\ \text{g mol}^{-1}$

$$n = \frac{m}{MM} = \frac{2.46}{142.05} = 0.0173\ \text{mol}$$

2 $MM\ (CaCO_3) = 40.08 + 12.01 + (3 \times 16.00) = 100.09\ \text{g mol}^{-1}$

$m = n \times MM = 34.2 \times 100.09 = 3420\ \text{g}$

Worked example 7.6

1 $$n = \frac{m}{MM} = \frac{0.213}{32.07} = 6.64 \times 10^{-3}\ \text{mol}$$

$N = n \times N_A = 6.64 \times 10^{-3} \times 6.022 \times 10^{23} = 4.00 \times 10^{21}$ atoms

2 $$n = \frac{N}{N_A} = \frac{4.6 \times 10^{18}}{6022\ \times 10^{23}} = 7.64 \times 10^{-6}\ \text{mol}$$

$MM = (3 \times 1.008) + 10.81 + (3 \times 16.00) = 61.834\ \text{g mol}^{-1}$

$m = n \times MM = 7.64 \times 10^{-6} \times 61.834 = 4.7 \times 10^{-4}\ \text{g}$

Check your understanding 7.7/7.8

1 63.6 g

2 Relative molecular mass has no units as it is a comparison to other elements rather than a specific measurement.

3 It could be a mole of nitrogen atoms or nitrogen molecules. Nitrogen molecules have two nitrogen atoms in them, so the wording needs to include 'atoms' or 'molecules'.

4 $n = \frac{m}{MM}$ where n = number moles (mol), m = mass (g), MM = molecular mass (g mol^{-1})

5 **a** Divide by the molecular mass of the substance.

b Multiply by the Avogadro constant, 6.022×10^{23}.

c Divide by the molecular mass to find the moles, then multiply by the Avogadro constant.

6 **a** $$n = \frac{m}{MM} = \frac{13}{65.38} = 0.20\ \text{mol}$$

b $$n = \frac{m}{MM} = \frac{0.50}{197.0} = 0.0025\ \text{mol}$$

c $$n = \frac{m}{MM} = \frac{15}{32.07} = 0.47\ \text{mol}$$

d $$n = \frac{m}{MM} = \frac{42}{12.01} = 3.5\ \text{mol}$$

e $n = \frac{m}{MM} = \frac{5.0 \times 10^{3}}{55.85} = 90 \text{ mol}$

f $n = \frac{m}{MM} = \frac{1 \times 10^{6}}{207.2} = 4826 \text{ mol}$

7 **a** $m = n \times MM = 3.0 \times 63.55 = 191 \text{ g}$

b $m = n \times MM = 0.044 \times 32.07 = 1.4 \text{ g}$

c $m = n \times MM = 2.3 \times 10^{-3} \times 24.31 = 0.056 \text{ g}$

8 **a** $N = n \times N_A = 0.20 \times 6.022 \times 10^{23} = 1.2 \times 10^{23}$ atoms

b $N = n \times N_A = 0.0025 \times 6.022 \times 10^{23} = 1.3 \times 10^{23}$ atoms

c $N = n \times N_A = 0.47 \times 6.022 \times 10^{23} = 2.8 \times 10^{23}$ atoms

d $N = n \times N_A = 3.5 \times 6.022 \times 10^{23} = 2.1 \times 10^{24}$ atoms

e $N = n \times N_A = 90 \times 6.022 \times 10^{23} = 5.4 \times 10^{25}$ atoms

f $N = n \times N_A = 4826 \times 6.022 \times 10^{23} = 2.9 \times 10^{27}$ atoms

9 **a** $n = \frac{N}{N_A} = \frac{1}{6022 \times 10^{23}} = 1.66 \times 10^{-24} \text{ mol}$

$m = n \times MM = 1.66 \times 10^{-24} \times 65.38 = 1.09 \times 10^{-22} \text{ g}$

b $n = \frac{N}{N_A} = \frac{1}{6022 \times 10^{23}} = 1.66 \times 10^{-24} \text{ mol}$

$m = n \times MM = 1.66 \times 10^{-24} \times 197.0 = 3.27 \times 10^{-22} \text{ g}$

c $n = \frac{N}{N_A} = \frac{1}{6022 \times 10^{23}} = 1.66 \times 10^{-24} \text{ mol}$

$m = n \times MM = 1.66 \times 10^{-24} \times 4.003 = 6.64 \times 10^{-24} \text{ g}$

10 $n = \frac{m}{MM} = \frac{0.10}{12.01} = 0.0083 \text{ mol}$

$N = n \times N_A = 0.0083 \times 6.022 \times 10^{23} = 5.0 \times 10^{21}$ atoms

11 **a** $MM\ (\text{Na}_2\text{SO}_4) = (2 \times 22.99) + 32.07 + (4 \times 16.00) = 142.05 \text{ g mol}^{-1}$

$n = \frac{m}{MM} = \frac{20}{142.05} = 0.14 \text{ mol}$

b $MM\ (\text{P}_2\text{O}_5) = (2 \times 30.97) + (5 \times 16.00) = 141.94 \text{ g mol}^{-1}$

$n = \frac{m}{MM} = \frac{20}{141.94} = 0.14 \text{ mol}$

c $MM\ (\text{C}_6\text{H}_{12}\text{O}_6) = (6 \times 12.01) + (12 \times 1.008) + (6 \times 16.00) = 180.156 \text{ g mol}^{-1}$

$n = \frac{m}{MM} = \frac{20}{180.156} = 0.11 \text{ mol}$

d $MM\ (\text{Cu(NO}_3)_2) = 63.55 + (2 \times 14.01) + (6 \times 16.00) = 187.57 \text{ g mol}^{-1}$

$n = \frac{m}{MM} = \frac{20}{187.57} = 0.11 \text{ mol}$

12 a $MM\ (\mathrm{MgCl_2}) = 24.31 + (2 \times 35.45) = 95.21\ \mathrm{g\ mol^{-1}}$

$m = n \times MM = 4.6 \times 95.21 = 440\ \mathrm{g}$

b $MM\ (\mathrm{CaSO_4}) = 40.08 + 32.07 + (4 \times 16.00) = 136.15\ \mathrm{g\ mol^{-1}}$

$m = n \times MM = 0.037 \times 136.15 = 5.0\ \mathrm{g}$

c $MM\ (\mathrm{NH_4NO_3}) = (2 \times 14.01) + (4 \times 1.008) + (3 \times 16.00) = 80.052\ \mathrm{g\ mol^{-1}}$

$m = n \times MM = 6.5 \times 10^{-4} \times 80.052 = 0.052\ \mathrm{g}$

13 a $MM\ (\mathrm{Cl_2}) = 2 \times 35.45 = 70.9\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{70.9} = 0.141\ \mathrm{mol}$$

$N = n \times N_A = 0.141 \times 6.022 \times 10^{23} = 8.5 \times 10^{22}$ molecules

b $MM\ (\mathrm{PCl_5}) = 30.97 + (5 \times 35.45) = 208.22\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{208.22} = 0.048\ \mathrm{mol}$$

$N = n \times N_A = 0.048 \times 6.022 \times 10^{23} = 2.9 \times 10^{22}$ molecules

c $MM\ (\mathrm{HNO_3}) = 1.008 + 14.01 + (3 \times 16.00) = 63.018\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{63.018} = 0.159\ \mathrm{mol}$$

$N = n \times N_A = 0.159 \times 6.022 \times 10^{23} = 9.6 \times 10^{22}$ molecules

d $MM\ (\mathrm{H_2O_2}) = (2 \times 1.008) + (2 \times 16.00) = 34.016\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{34.016} = 0.294\ \mathrm{mol}$$

$N = n \times N_A = 0.294 \times 6.022 \times 10^{23} = 1.8 \times 10^{23}$ molecules

e $MM\ (\mathrm{C_3H_8}) = (3 \times 12.01) + (8 \times 1.008) = 44.094\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{44.094} = 0.227\ \mathrm{mol}$$

$N = n \times N_A = 0.227 \times 6.022 \times 10^{23} = 1.4 \times 10^{23}$ molecules

f $MM\ (\mathrm{C_2H_5OH}) = (2 \times 12.01) + (6 \times 1.008) + 16.00 = 46.068\ \mathrm{g\ mol^{-1}}$

$$n = \frac{m}{MM} = \frac{10}{46.068} = 0.217\ \mathrm{mol}$$

$\mathrm{N} = n \times \mathrm{N_A} = 0.217 \times 6.022 \times 10^{23} = 1.3 \times 10^{23}$ molecules

14 a $m = n \times MM = 5.18 \times (2 \times 14.01) = 145\ \mathrm{g}$

b $m = n \times MM = 0.082 \times 39.95 = 3.3\ \mathrm{g}$

c $m = n \times MM = 6.24 \times 10^{-2} \times (2 \times 1.008) = 0.13\ \mathrm{g}$

15 a $$n = \frac{m}{MM} = \frac{7.0}{(2 \times 14.01)} = 0.2498\ \mathrm{mol}$$

$N = n \times N_A = 0.2498 \times 6.022 \times 10^{23} = 1.5 \times 10^{23}$ molecules

b $2 \times 1.5 \times 10^{23} = 3.0 \times 10^{23}$ atoms

16 a $MM\ (\text{AlCl}_3) = 26.98 + (3 \times 35.45) = 133.33\ \text{g mol}^{-1}$

$$n = \frac{m}{MM} = \frac{0.374}{133.33} = 0.00281\ \text{mol}$$

b AlCl_3 contains 1 part Al to 1 part AlCl_3, so $n = 0.00281$ mol

c AlCl_3 contains 3 parts Cl to 1 part AlCl_3, so $n = 3 \times 0.00281 = 0.00842$ mol

17 a $n = \frac{N}{N_A} = \frac{1}{6022\ \times 10^{23}} = 1.66 \times 10^{-24}\ \text{mol}$

$MM = 14.01 + (3 \times 1.008) = 17.034\ \text{g mol}^{-1}$

$m = n \times MM = 1.66 \times 10^{-24} \times 17.034 = 2.828 \times 10^{-23}\ \text{g}$

b $n = \frac{N}{N_A} = \frac{1}{6022\ \times 10^{23}} = 1.66 \times 10^{-24}\ \text{mol}$

$MM = 2 \times 35.45 = 70.9\ \text{g mol}^{-1}$

$m = n \times MM = 1.66 \times 10^{-24} \times 70.9 = 1.18 \times 10^{-22}\ \text{g}$

c $n = \frac{N}{N_A} = \frac{1}{6022\ \times 10^{23}} = 1.66 \times 10^{-24}\ \text{mol}$

$MM = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00) = 342.296\ \text{g mol}^{-1}$

$m = n \times MM = 1.66 \times 10^{-24} \times 342.296 = 5.682 \times 10^{-22}\ \text{g}$

18 a $MM = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00) = 342.296\ \text{g mol}^{-1}$

$$n = \frac{m}{MM} = \frac{1.4 \times 10^{-3}}{342.296} = 4.09 \times 10^{-6}\ \text{mol}$$

$N = n \times N_A = 4.09 \times 10^{-6} \times 6.022 \times 10^{23} = 2.5 \times 10^{18}$ molecules

b $12 + 22 + 11 = 45$

$45 \times 2.5 \times 10^{18} = 1.1 \times 10^{20}$ atoms

Worked example 7.7

$MM = (2 \times 14.01) + (8 \times 1.008) + 32.07 + (4 \times 16.00) = 132.154\ \text{g mol}^{-1}$

$$\%N = \frac{2 \times 14.01 \times 100}{132.154} = 21.2\%$$

Worked example 7.8

Mass of Cl = 1.75 − 0.84 = 0.91 g

$$n(\text{Cl}) = \frac{m}{MM} = \frac{0.91}{35.45} = 0.0257\ \text{mol},\ n(\text{Zn}) = \frac{m}{MM} = \frac{0.84}{65.38} = 0.0128\ \text{mol}$$

$\text{Cl} = \frac{0.0257}{0.0128} = 2$, $\text{Zn} = \frac{0.0128}{0.0128} = 1$, thus formula of the compound formed is ZnCl_2.

Worked example 7.9

silver : sulfur : oxygen

69.2 : 10.3 : 20.5

$$\frac{69.2}{107.9} : \frac{10.3}{32.07} : \frac{20.5}{16.00} \rightarrow 0.641 : 0.321 : 1.28$$

$$\rightarrow \frac{0.641}{0.321} : \frac{0.321}{0.321} : \frac{1.28}{0.321} \rightarrow 2.00 : 1.00 : 3.98 \rightarrow 2 : 1 : 4 \rightarrow Ag_2SO_4$$

Check your understanding 7.9/7.10

1 Empirical formula indicates the simplest ratio of atoms in the substance, molecular formula indicates the actual number of atoms in a single molecule of a substance. Ethane is a molecule with molecular formula C_2H_6. Its empirical formula is CH_3.

2 Work out the masses of each substance → convert all masses to moles by dividing by the molecular mass → divide each by the smallest number of moles to get a ratio → convert to whole numbers (if necessary) → round off to whole numbers and write empirical formula.

3 **a** $MM = (2 \times 22.99) + 32.07 + (4 \times 16.00) = 142.05 \text{ g mol}^{-1}$

$$\%S = \frac{3207 \times 100}{142.05} = 22.58\%$$

b $MM = (2 \times 26.98) + (3 \times 16.00) = 101.96 \text{ g mol}^{-1}$

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

c $MM = 137.3 + (2 \times 16.00) + (2 \times 1.008) = 171.316 \text{ g mol}^{-1}$

$$\%Ba = \frac{1373 \times 100}{171.316} = 80.14\%$$

4 **a** $MM = 107.9 + 14.01 + (3 \times 16.00) = 169.91 \text{ g mol}^{-1}$

$$\%Ag = \frac{1079 \times 100}{169.91} = 63.5\%, \%N = \frac{1401 \times 100}{169.91} = 8.24\%, \%O = \frac{3 \times 16.00 \times 100}{169.91} = 28.3\%$$

b $MM = 26.98 + (3 \times 16.00) + (3 \times 1.008) = 78.004 \text{ g mol}^{-1}$

$$\%Al = \frac{2698 \times 100}{78.004} = 34.6\%, \%O = \frac{3 \times 16.00 \times 100}{78.004} = 61.5\%, \%H = \frac{3 \times 1.008 \times 100}{78.004} = 3.88\%$$

c $MM = (3 \times 14.01) + (12 \times 1.008) + 30.97 + (4 \times 16.00) = 149.096 \text{ g mol}^{-1}$

$$\%N = \frac{3 \times 14.01 \times 100}{149.096} = 28.2\%, \%H = \frac{12 \times 1.008 \times 100}{149.096} = 8.11\%, \%P = \frac{3097 \times 100}{149.096} = 20.8\%,$$

$$\%O = \frac{4 \times 16.00 \times 100}{149.096} = 42.9\%$$

d $MM = (2 \times 55.85) + (3 \times 32.07) + (12 \times 16.00) = 399.91 \text{ g mol}^{-1}$

$$\%Fe = \frac{2 \times 55.85 \times 100}{399.91} = 27.9\%, \%S = \frac{3 \times 32.07 \times 100}{399.91} = 24.1\%, \%O = \frac{12 \times 16.00 \times 100}{399.91} = 48.0\%$$

5 **a** $MM = (2 \times 14.01) + (4 \times 1.008) + (3 \times 16.00) = 80.052 \text{ g mol}^{-1}$

$$\%N = \frac{2 \times 14.01 \times 100}{80.052} = 35.0\%$$

b $MM = 12.01 + (4 \times 1.008) + (2 \times 14.01) + 16.00 = 60.062 \text{ g mol}^{-1}$

$$\%N = \frac{2 \times 14.01 \times 100}{60.062} = 46.7\%$$

Urea has a higher percentage of nitrogen so would be better per gram.

6 $MM = (7 \times 12.01) + (5 \times 1.008) + (3 \times 14.01) + (6 \times 16.00) = 227.14 \text{ g mol}^{-1}$

$$\%N = \frac{3 \times 14.01 \times 100}{227.14} = 18.5\%$$

7 copper : chlorine

$1.00 - 0.524 = 0.476$ g copper

$$0.476 : 0.524 \rightarrow \frac{0.476}{63.55} : \frac{0.524}{35.45} \rightarrow 0.00749 : 0.0148 \rightarrow \frac{0.00749}{0.00749} : \frac{0.0148}{0.00749} \rightarrow 1.00 : 1.98 \rightarrow 1 : 2 \rightarrow CuCl_2$$

8 calcium : nitrogen : oxygen

$$24.5 : 17.2 : 58.3 \rightarrow \frac{24.5}{40.08} : \frac{17.2}{14.01} : \frac{58.3}{16.00} \rightarrow 0.611 : 1.23 : 3.64 \rightarrow \frac{0.611}{0.611} : \frac{1.23}{0.611} : \frac{3.64}{0.611} \rightarrow 1.00 : 2.01 : 5.95$$

$$\rightarrow 1 : 2 : 6 \rightarrow CaN_2O_6 = Ca(NO_3)_2$$

9 magnesium : nitrogen

$3.51 - 2.55 = 0.96$ g nitrogen

$$2.55 : 0.96 \rightarrow \frac{2.55}{24.31} : \frac{0.96}{14.01} \rightarrow 0.105 : 0.0685 \rightarrow \frac{0.105}{0.0685} : \frac{0.0685}{0.0685} \rightarrow 1.53 : 1.00 \rightarrow 3 : 2 \rightarrow Mg_3N_2$$

10 aluminium : oxygen

$$52.9 : 47.1 \rightarrow \frac{52.9}{26.98} : \frac{47.1}{16.00} \rightarrow 1.96 : 2.94 \rightarrow \frac{1.96}{1.96} : \frac{2.94}{1.96} \rightarrow 1.00 : 1.50 \rightarrow 2 : 3 \rightarrow Al_2O_3$$

Worked example 7.10

1 **a** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, nitrogen : hydrogen = 1 : 3, so if 0.10 mol nitrogen, $0.10 \times 3 = 0.30$ mol hydrogen

b nitrogen : ammonia = 1 : 2, so if 0.10 mol nitrogen, $0.10 \times 2 = 0.20$ mol ammonia formed

2 **a** aluminium : sulfuric acid = 2 : 3, so if 0.35 mol aluminium, $0.35 \times \frac{3}{2} = 0.525$ mol

b aluminium : aluminium sulfate = 2 : 1, so if 0.35 mol aluminium, $0.35 \times \frac{1}{2} = 0.175$ mol

Worked example 7.11

$$n(HNO_3) = \frac{m}{MM} = \frac{17.4}{1.008 + 14.01 + (3 \times 16.00)} = 0.2761 \text{ mol}$$

calcium carbonate : nitric acid = 1 : 2

$n(CaCO_3) = 0.5 \times 0.2761 = 0.1381$ mol

$m(CaCO_3) = n \times MM = 0.1381 \times (40.08 + 12.01 + (3 \times 16.00)) = 13.82 = 13.8$ g (3sf)

Worked example 7.12

$2HCl(aq) + Mg(s) \rightarrow H_2(g) + MgCl_2(aq)$

$$n(Mg) = \frac{m}{MM} = \frac{0.84}{24.31} = 0.03455 \text{ mol}$$

hydrochloric acid : magnesium = 2 : 1

$n(HCl) = 2 \times 0.03455 = 0.06911$ mol

$m(HCl) = n \times MM = 0.06911 \times (1.008 + 35.45) = 2.520 = 2.5$ g (2sf)

Check your understanding 7.11/7.12

1 Write a balanced chemical equation → calculate the number of moles of the given substance using $n = \frac{m}{MM}$ → write the ratio of moles of the required substance and the given substance → use the mole ratio to determine the moles of the required substance → calculate the mass of the required substance using $m = n \times MM$

2 **a** $2C_2H_6(g) + 5O_2(g) \rightarrow 4CO_2(g) + 6H_2O(l)$

- **i** 2.5 moles of oxygen for each mole of ethane
- **ii** 2 moles of carbon dioxide for each mole of ethane
- **iii** 3 moles of water for each mole of ethane

b $C_8H_{18}(g) + 12.5O_2(g) \rightarrow 8CO_2(g) + 9H_2O(l)$

- **i** $12.5 \times 10 = 125$ moles
- **ii** 9 moles water for each mole of octane

3 **a** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

hydrogen : ammonia = 3 : 2, thus $40 \times 1.5 = 60$ mol hydrogen

b $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$

oxygen : sulfur trioxide = 1 : 2, thus $0.6 \times 0.5 = 0.3$ mol oxygen

4 **a** zinc oxide : zinc = 1 : 1, thus 0.2 mol zinc required

$m(Zn) = n \times MM = 0.2 \times 65.38 = 13.1 = 10$ g (1sf)

b $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

magnesium sulfate : magnesium hydroxide = 1 : 1

$n(Mg(OH)_2) = 0.050$ mol

$m(Mg(OH)_2) = n \times MM = 0.050 \times (24.31 + (2 \times 16.00) + (2 \times 1.008)) = 2.916 = 2.9$ g (2sf)

5 $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$$n(CaCO_3) = \frac{m}{MM} = \frac{2.0}{40.08 + 12.01 + (3 \times 16.00)} = 0.01998 \text{ mol}$$

mole ratio = 1 : 1, thus $n(CaO) = 0.01998$ mol

$m(CaO) = 0.01998 \times (40.08 + 16.00) = 1.120 = 1.1$ g (2sf)

6 $2\text{Na(s)} + 2\text{H}_2\text{O(l)} \rightarrow 2\text{NaOH(aq)} + \text{H}_2\text{(g)}$

a $n(\text{Na}) = \frac{m}{MM} = \frac{1.32}{22.99} = 0.05742 \text{ mol}$

mole ratio = 1 : 1, $n(\text{H}_2\text{O}) = 0.05742$ mol

$m(\text{H}_2\text{O}) = n \times MM = 0.05742 \times ((2 \times 1.008) + 16.00) = 1.034 = 1.03 \text{ g (3sf)}$

b sodium : hydrogen = 2 : 1, $n(\text{H}_2) = 0.5 \times 0.05742 = 0.02871$ mol

$m(\text{H}_2) = n \times MM = 0.02871 \times (2 \times 1.008) = 0.05788 = 0.0579 \text{ g (3sf)}$

7 $\text{HCl(aq)} + \text{NaHCO}_3\text{(s)} \rightarrow \text{CO}_2\text{(g)} + \text{NaCl(aq)} + \text{H}_2\text{O(l)}$

$n(\text{HCl}) = \frac{m}{MM} = \frac{18.6}{1008\ +\qquad 35.45} = 0.5102 \text{ mol}$

mole ratio = 1 : 1, $n(\text{NaHCO}_3) = 0.5102$ mol

$m(\text{NaHCO}_3) = n \times MM = 0.5102 \times (22.99 + 1.008 + 12.01 + (3 \times 16.00)) = 42.86 = 42.9 \text{ g (3sf)}$

Worked example 7.13

If 12 mol ammonia used, then $\frac{5}{4} \times 12 = 15$ mol oxygen required. As 14 mol oxgyen present, the ammonia is in excess.

If 14 mol oxygen used, then $\frac{4}{5} \times 14 = 11.2$ mol ammonia required.

$12 - 11.2 = 0.8$ mol ammonia left over.

Worked example 7.14

$\text{Al(NO}_3)_3\text{(aq)} + 3\text{NaOH(aq)} \rightarrow \text{Al(OH)}_3\text{(s)} + 3\text{NaNO}_3\text{(aq)}$

$n(\text{Al(NO}_3)_3) = \frac{m}{MM} = \frac{7.0}{26.98 + (3 \times 14.01) + (9 \times 16.00)} = 0.0329 \text{ mol}$

$n(\text{NaOH}) = \frac{m}{MM} = \frac{4.6}{2299\ +\qquad +\ \ 16.0\ \ 1.008} = 0.1150 \text{ mol}$

If 0.0329 mol aluminium nitrate used then $3 \times 0.0329 = 0.0987$ mol NaOH required. Thus NaOH is in excess.

aluminium nitrate : aluminium hydroxide = 1 : 1, thus $n(\text{Al(OH)}_3) = 0.0329$ mol

$m(\text{Al(OH)}_3) = n \times MM = 0.0329 \times (26.98 + (3 \times 1.008) + (3 \times 16.00)) = 2.566 = 2.6 \text{ g (2sf)}$

Check your understanding 7.13

1 The reagent that is completely used up in a reaction, hence limiting the final amount of product.

2 Calculate the moles of the two reactants. Use the mole ratio from the balanced equation to determine which substance is the limiting reagent. Use the moles of the limiting reagent to determine the moles of the product. Calculate the mass of the product.

3 hydrogen sulfide : oxygen = 2 : 3

a 4 mol oxygen reacts then $\frac{2}{3} \times 4 = 2.67$ mol hydrogen sulfide required. $3 - 2.67 = 0.33$ mol hydrogen sulfide in excess.

b 4 mol hydrogen sulfide reacts then $\frac{3}{2} \times 4 = 6$ mol oxygen required. 1 mol oxygen in excess.

c 0.2 mol hydrogen sulfide reacts then $\frac{3}{2} \times 0.2 = 0.3$ mol. 0.05 mol oxygen in excess.

d 0.20 mol oxygen reacts then $\frac{2}{3} \times 0.20 = 0.13$ mol hydrogen sulfide required. 0.03 mol hydrogen sulfide in excess.

4 copper : nitric acid = 3 : 8, copper : nitric oxide = 3 : 2, nitric acid : nitric oxide = 8 : 2

a 15 mol nitric acid reacts then $\frac{3}{8} \times 15 = 5.625$ mol copper required, so copper is in excess.
$\frac{2}{8} \times 15 = 3.75$, thus 3.75 mol nitric oxide formed.

b 0.25 mol copper reacts then $\frac{8}{3} \times 0.25 = 0.667$ mol nitric acid required, so nitric acid is in excess.
$\frac{2}{3} \times 0.25 = 0.167$ mol nitric oxide formed.

c 0.5 mol nitric acid reacts then $\frac{3}{8} \times 0.5 = 0.1875$ mol copper required, so copper is in excess.
$\frac{2}{8} \times 0.5 = 0.125$ mol nitric oxide formed.

5 $H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

$$n(\text{NaOH}) = \frac{m}{MM} = \frac{24}{2299\ + \qquad + 1.016800} = 0.600 \text{ mol}$$

$$n(H_2SO_4) = \frac{m}{MM} = \frac{19.6}{(2 \times 1.008) + 32.07 + (4 \times 16.00)} = 0.200 \text{ mol}$$

If 0.200 mol sulfuric acid reacts then 2 × 0.200 = 0.400 mol sodium hydroxide required, so sodium hydroxide is in excess.

a $n(H_2SO_4) = n(Na_2SO_4) = 0.200$ mol

$m(Na_2SO_4) = n \times MM = 0.200 \times ((2 \times 22.99) + 32.07 + (4 \times 16.00)) = 28.4$ g (3sf)

b Sodium hydroxide is in excess by 0.200 mol.

$m(\text{NaOH}) = n \times MM = 0.200 \times (22.99 + 16.00 + 1.008) = 8.0$ g (2sf)

6 $AlCl_3(aq) + 3NaOH(aq) \rightarrow Al(OH)_3(s) + 3NaCl$

$$n(\text{NaOH}) = \frac{m}{MM} = \frac{12}{2299\ + \qquad + \quad 16.0 \quad 1.008} = 0.300 \text{ mol}$$

$$n(AlCl_3) = \frac{m}{MM} = \frac{15}{26.98 + (3 \times 35.45)} = 0.113 \text{ mol}$$

If 0.300 mol sodium hydroxide reacts, then $\frac{1}{3} \times 0.300 = 0.100$ mol aluminium chloride required. Thus aluminium chloride is in excess by 0.113 mol.

Sodium hydroxide : aluminium hydroxide = 1 : 3, so $\frac{1}{3} \times 0.300 = 0.100$ mol aluminium hydroxide formed.

$m(Al(OH)_3) = n \times MM = 0.100 \times (26.98 + (3 \times 16.00) + (3 \times 1.008)) = 7.8$ g (2sf)

7 $CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

$$n(\text{HCl}) = \frac{m}{MM} = \frac{1.0}{1008\ + \qquad 35.45} = 0.0274 \text{ mol}$$

$$n(CaCO_3) = \frac{m}{MM} = \frac{2.5}{40.08 + 12.01 + (3 \times 16.00)} = 0.0250 \text{ mol}$$

0.0274 moles of HCl reacts, then $\frac{1}{2} \times 0.0274 = 0.0137$ moles of calcium carbonate required.

0.0250 moles of $CaCO_3$ reacts, then $0.0250 \times 2 = 0.0500$ moles HCl; therefore HCL is the limiting agent.

HCl : calcium chloride = 2 : 1, thus $n(CaCl_2) = 0.0274/2$ mol

$m(CaCl_2) = n \times MM = 0.0137 \times (40.08 + (2 \times 35.45)) = 1.5$ g

Chapter review questions

1 Matter can neither be created nor destroyed, but changed from one form to another. In a chemical reaction, mass is conserved and the number of atoms of each element is conserved. Experiments you may perform vary – check with your teacher for the experiment you conducted in your classroom.

2 Equations must be balanced for each type of atom.

3 Weight is a measure of force, not of mass. It is the convention to keep using atomic weight as it has been used for a long period of time. The preferred name is atomic mass or relative atomic mass.

4 Molecular weight

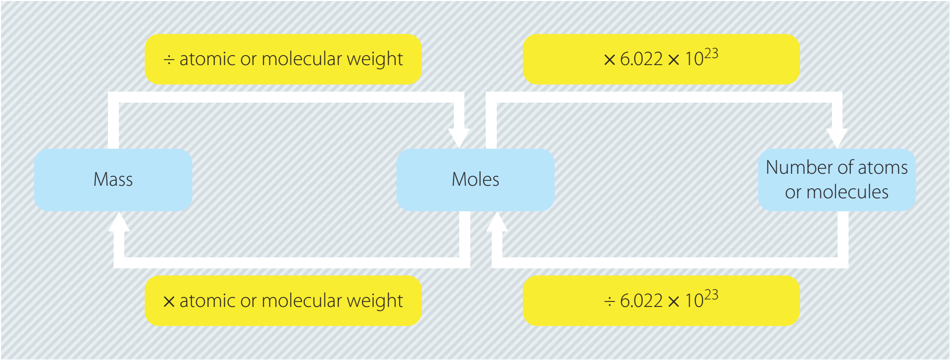
5 A molecule is a distinct particle that has a fixed number of atoms in its structure held together by covalent bonds. A mole is a fixed number of particles (ions, atoms, molecules) and is 6.022×10^{23}, also known as the Avogadro constant.

6 $n = \frac{m}{MM}$ where n = number moles (mol), m = mass (g), MM = molecular mass (g mol^{-1})

7 Multiply by the Avogadro constant (6.022×10^{23}).

8 It could refer to a mole of chlorine atoms (Cl) or a mole of chorine molecules (Cl_2).

9



10 Relative molecular mass – a number that expresses the relative mass of the molecule, relative to carbon. Molar mass (sometimes molecular mass) is the mass of one mole of the substance.

11 (Example) C_8H_{16} (molecular formula) has 8 carbons and 16 hydrogens in one molecule. The simplest ratio of carbon to hydrogen is CH_2 (empirical formula).

12 Work out the masses of each substance → convert all masses to moles by dividing by the molecular mass → divide each by the smallest number of moles to get a ratio → convert to whole numbers (if necessary) → round off to whole numbers and write empirical formula

13 Refer to diagram on page 165.

14 Write a balanced equation for the reaction. Calculate the moles of each of the two reactants. Use the mole ratio from the balanced equation to determine which substance is the limiting reagent. Use the moles of the limiting reagent to determine the moles of the product. Calculate the mass of the product.

15 a $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

b $Na_2S_2O_3(s) + 2HCl(aq) \rightarrow S(s) + SO_2(g) + 2NaCl(aq) + H_2O(l)$

c $Pb(s) + PbO_2(s) + 2H_2SO_4(aq) \rightarrow 2PbSO_4(aq) + 2H_2O(l)$

16 Relative atomic mass $= \dfrac{(88 \times 83) + (87 \times 7) + (86 \times 10)}{100} = 87.73$

17 $20.2 = \dfrac{(20 \times X) + (22 \times (100 - X))}{100}$

$20.2 = \dfrac{20X + 2200 - 22X}{100}$

$2020 = -2X + 2200$

$2X = 180$, so $X = 90$

Thus 90% Ne-20, 10% Ne-22

18 a i RMM $= (4 \times 12.01) + (6 \times 1.008) + (6 \times 16.00) = 150.088$

ii RMM $= (19 \times 12.01) + (28 \times 1.008) + (2 \times 16.00) = 288.414$

b i RFM $= (2 \times 55.85) + (3 \times 32.07) + (12 \times 1.008) = 220.006$

ii RFM $= 40.08 + (4 \times 1.008) + (2 \times 30.97) + (8 \times 16.00) = 234.052$

19 a 39

b $N = n \times N_A = 0.10 \times 6.022 \times 10^{23} = 6.0 \times 10^{22}$ molecules

c $5 \times 6.022 \times 10^{23} = 3.011 \times 10^{23}$ atoms of oxygen

20 $MM = 63.55 + 55.85 + (2 \times 32.07) = 183.54\ g\ mol^{-1}$

$\%Cu = \dfrac{6355\ \times 100}{183.54} = 34.6\%$

21 Chalcocite $MM = (2 \times 63.55) + 32.07 = 159.17\ g\ mol^{-1}$

$\%Cu = \dfrac{6355\ \times 100}{159.17} = 39.9\%$

Malachite $MM = (2 \times 63.55) + (5 \times 16.00) + (2 \times 1.008) + 12.01 = 221.126\ g\ mol^{-1}$

$\%Cu = \dfrac{2 \times 63.55 \times 100}{221.126} = 57.5\%$

Azurite $MM = (3 \times 63.55) + (2 \times 1.008) + (8 \times 16.00) + (2 \times 12.01) = 281.136\ g\ mol^{-1}$

$\%Cu = \dfrac{3 \times 63.55 \times 100}{281.136} = 67.8\%$

In order of decreasing %Cu: azurite, malachite, chalcocite, chalcopyrite

22 carbon : hydrogen : oxygen : chlorine : nitrogen

$$40.99 : 3.7 : 24.8 : 22.0 : 8.7 \rightarrow \frac{40.99}{12.01} : \frac{3.7}{1.008} : \frac{24.8}{16.00} : \frac{22.0}{35.45} : \frac{8.7}{14.01} \rightarrow 3.41 : 3.67 : 1.55 : 0.62 : 0.62$$

$$\rightarrow \frac{3.41}{0.62} : \frac{3.67}{0.62} : \frac{1.55}{0.62} : \frac{0.62}{0.62} : \frac{0.62}{0.62} \rightarrow 5.5 : 5.91 : 2.50 : 1.00 : 1.00 \rightarrow 11 : 12 : 5 : 2 : 2$$

$\rightarrow C_{11}H_{12}O_5Cl_2N_2$

23 Compound 1 – copper : sulfur

$$0.84 : 0.21 \rightarrow \frac{0.84}{63.55} : \frac{0.21}{32.07} \rightarrow 0.013 : 0.0065 \rightarrow \frac{0.013}{0.0065} : \frac{0.0065}{0.0065} \rightarrow 2.00 : 1.00 \rightarrow 2 : 1$$

Compound 2 – copper : sulfur

$$1.26 : 0.32 \rightarrow \frac{1.26}{63.55} : \frac{0.32}{32.07} \rightarrow 0.0198 : 0.00998 \rightarrow \frac{0.0198}{0.00998} : \frac{0.00998}{0.00998} \rightarrow 1.98 : 1.00 \rightarrow 2 : 1$$

Compound 3 – copper : sulfur

$$1.81 : 0.45 \rightarrow \frac{1.81}{63.55} : \frac{0.45}{32.07} \rightarrow 0.0285 : 0.0140 \rightarrow \frac{0.0285}{0.0140} : \frac{0.0140}{0.0140} \rightarrow 2.04 : 1.00 \rightarrow 2 : 1$$

Compound 4 – copper : sulfur

$$2.53 : 0.65 \rightarrow \frac{2.53}{63.55} : \frac{0.65}{32.07} \rightarrow 0.0398 : 0.0203 \rightarrow \frac{0.0398}{0.0203} : \frac{0.0203}{0.0203} \rightarrow 1.96 : 1.00 \rightarrow 2 : 1$$

As all have a 2 : 1 ratio for copper : sulfur, the hypothesis was supported.

24 $$n(Al_2O_3) = \frac{m}{MM} = \frac{12.7}{(2 \times 26.98) + (3 \times 16.00)} = 0.1246 \text{ mol}$$

$$n(NaOH) = \frac{2}{1} \times 0.1246 = 0.2491 \text{ mol}$$

$$m(NaOH) = n \times MM = 0.2491 \times (22.99 + 16.00 + 1.008) = 9.964 = 9.96 \text{ g (3sf)}$$

25 Compound 1 – copper : sulfur

$$0.84 : 0.21 \rightarrow \frac{0.84}{63.55} : \frac{0.21}{32.07} \rightarrow 0.013 : 0.0065 \rightarrow \frac{0.013}{0.0065} : \frac{0.0065}{0.0065} \rightarrow 2.00 : 1.00 \rightarrow 2 : 1$$

Compound 2 – copper : sulfur

$$1.26 : 0.32 \rightarrow \frac{1.26}{63.55} : \frac{0.32}{32.07} \rightarrow 0.0198 : 0.00998 \rightarrow \frac{0.0198}{0.00998} : \frac{0.00998}{0.00998} \rightarrow 1.98 : 1.00 \rightarrow 2 : 1$$

Compound 3 – copper : sulfur

$$1.81 : 0.45 \rightarrow \frac{1.81}{63.55} : \frac{0.45}{32.07} \rightarrow 0.0285 : 0.0140 \rightarrow \frac{0.0285}{0.0140} : \frac{0.0140}{0.0140} \rightarrow 2.04 : 1.00 \rightarrow 2 : 1$$

Compound 4 – copper : sulfur

$$2.53 : 0.65 \rightarrow \frac{2.53}{63.55} : \frac{0.65}{32.07} \rightarrow 0.0398 : 0.0203 \rightarrow \frac{0.0398}{0.0203} : \frac{0.0203}{0.0203} \rightarrow 1.96 : 1.00 \rightarrow 2 : 1$$

As all have a 2 : 1 ratio for copper : sulfur, the hypothesis was supported.

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

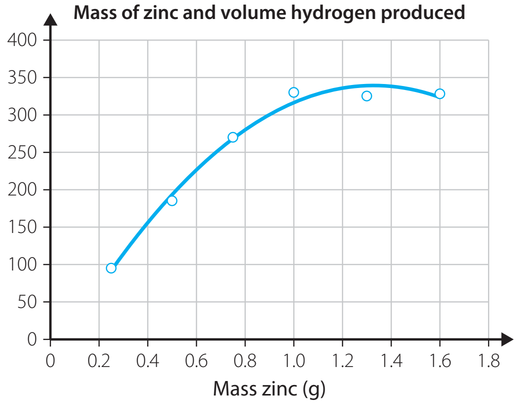
i $\frac{13}{2} \times 4.0 = 26$ mol

ii $\frac{13}{2} \times 0.067 = 0.4355$ mol

b i $\frac{8}{2} \times 4.0 = 16$ mol

ii $\frac{8}{2} \times 0.067 = 0.268$ mol

27 a



b The hydrochloric acid is the limiting reagent, so once it is used up the amount of hydrogen produced is reasonably constant.

c Graph shows approximately 1.0 g zinc.

$$n(\mathrm{HCl}) = \frac{m}{MM} = \frac{1.00}{1008\ +\ 35.45} = 0.0274 \text{ mol}$$

$$n\ (Zn) = \frac{m}{MM} = \frac{1.00}{65.38} = 0.0153 \text{ mol}$$

$\frac{0.0274}{0.0153} = 1.79$, thus hydrochloric acid : zinc = 1.79 : 1.00

According to the equation $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ the ratio should be 2 : 1. Given the spread of data points on the graph and the difficulty in reading the mass that sees the volume of hydrogen level off, the ratio is approximately 2 : 1 and this matches the expected. More data points would give a more confident answer.

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 8: Solutions – concentration and molarity

Worked examle .1

1 $221\ \text{g}\ 100\ \text{mL}^{-} \times 10 = 221\ \text{g}\ \text{L}^{-}\ (100\ \text{mL} \times 10 = 1000\ \text{mL}/1\text{L})$

2 $136\ \text{g}/250\ \text{g} \rightarrow 136/25 = 544\ \text{g}/100\ \text{g} = 544\%(\text{w}/\text{w})$

Worked examle .2

1 $25\%(\text{w}/\text{w}) = 25\ \text{g}/100\ \text{g} = 125\ \text{g}/500\ \text{g}$. Thus ned 12.5 g ofiodine.

2 $60\%(\text{v}/\text{v}) = 60\ \text{mL}/100\ \text{mL} = 30\ \text{mL}/500\ \text{mL}$ Thus have 30 mL of acetc acid.

3 $23\ \text{p pm} = 23\ \text{g}/10^{6}\ \text{g} = 23 \times 10^{-4}\ \text{g}/100\ \text{g} = 23 \times 10^{-4}\ \%(\text{w}/\text{w})$

Check your understandng 81/82/83

1 In a solution, the particles are fully dissolved in the solvent and do not settle out. Particles in a suspension are visible, and will settle out s a sediment over time.

2 A solution in which the solvent is water.

3 Mass of solute per 100 mLor per L.

Volume of solute per 100 L or per L.

%(w/w) – percent weight/weight – the mass of soluteper 100 mL solution.

%(w/v) – percent weight/volume – the volume of solute per 100 mL solution.

ppm – parts per million – grams of solid or liquid solute per million grams of solution.

4 **a** The number of grams of solid or liquid solute per millin grams of solution.

b Divide by 10 000 (to convert from 10^{6} mL to 100 mL.

5 **a** measuring cylinder

b pipette or burette

6 **a** $\dfrac{7\ 3 \times 100}{75} = 973\ \text{g}/100\ \text{mL}$

ii $973 \times 10 = 973\ \text{g/L}$

iii 973%(w/v)

b $\frac{15\times 100}{95} = 1579$ mL/100 mL

ii $\frac{11.9\times 100}{95} = 1253$ g/100 mL

iii $1579\ \times 10 = 1579$ mL/L

v %(v/v) = volume solute per 100 mL solution = 1579 mL

7 **a** 25%(w/v) = 25 g/100 mL Thus, to make 50mL, 2.5/2 = 1.25 g

b 70%(w/w) = 70 g/100 g. Thus, to make 200 g solution 7.0 × 2 = 14.0 g

c 750 ppm = 750 g/10^6 g. Thus, to make 500 g, 750/2000 = 0375 g

8 **a** Add .5 g ammonium chloride, measured on an electronic balance, to a 100 mL measuring cyliner. Add water to dissolve all the ammonium chloride and then fill to the 100 mL mark with watr.

ii Add .5 g ammonium chloride, measured on an electronic balance, to a 100.0 mL volumetric flask, dissolve in water, fill the flask to the graduation mark so the bottom of the meniscus is on the mark.

b Use a measuring cylinder to measure out (15.0 × 5 =) 750 mL of ethanol. Measure out (85 × 5 =) 425 mL of water in a different measurig cylinder. Use a beaker to mix the two together.

ii Measure 75.0 mL of ethanol using a pipette and place into a 00.0 mL volumetric flask. Make up to the graduation mark with water so the bottom of the meniscus is on the mark.

9 85%(w/v) = 85 g/100 mL Measure on an electronc balance (8.5 × 25 =) 2125 g of sodium carbonate and transfer to a 250 mL volumetric flask. Add enough water to dissolve all the solid, then make up the flask to the graduation mark with water so the bottom o the meniscus is on the mark.

Worked examle .3

1 **a** $AgNO_3\ n = \frac{m}{MM} = \frac{10\ 6}{107\ 9 + 14\ 01 + (3\times 16\ 00)} = 00624$ mol

$c = \frac{n}{V} = \frac{0\ 0624}{0\ 2500} = 02495\ = 0250$ mol L $^{-}$ (3sf)

b $K_2SO_4\ n = \frac{m}{MM} = \frac{34\ 8}{(2\times 39\ 10) + 32\ 07 + (4\times 16\ 00)} = 01997$ mol

$c = \frac{n}{V} = \frac{0\ 1997}{0\ 1000} = 1997\ = 200$ mol L $^{-}$ (3sf)

2 **a** $(NH_4)_2SO_4\ n = c\times V = 0304\ \times 05000\ = 0152$ mol

$m = n\times MM = 0152\ \times ((2\times 14.01) + (8\times 1008)\ + 3207\ + (4\times 1600))\ = 2009\ = 201$ g (3sf)

b $Pb(NO_3)_2\ n = c\times V = 00500\ \times 02500\ = 00125$ mol

$m = n\times MM = 00125\ \times (2072\ + (2\times 14.01) + (6\times 1600))\ = 4140\ = 414$ g (3sf)

3 **a** $Ca(OH)_2\ n = c\times V = 00126\ \times 00457\ = 576\ \times 10^{-4}$ mol (3sf)

b $m = n\times MM = 576\ \times 10^{-4}\times (4008\ + (2\times 1008)\ + (2\times 1600))\ = 004268\ = 00427$ g (3sf)

4 a $n = c \times V = 1.55 \times 0150\ = 0233\ \text{mol}$

b $m = n \times MM = 02325\ \times (2431\ + (2 \times 3545))\ = 2214\ = 221\ \text{g (3sf)}$

Worked examle .4

1 $c\ V\ = c_2V_2$

$25 \times 1.2 = c_2 \times 500$

$c_2 = 0060\%(\text{w/w})$

2 $c\ V\ = c_2V_2$

$106\ \times 25 = 0053\ \times V_2$

$V_2 = 500\ \text{mL}$

Check your understandng 84/85

1 a $n = c \times V$ where $n =$ number moles (mol), $c =$ concentration of solution (mol L^{-}), $V =$ volume (L)

b $V = \frac{n}{c}$ where $V =$ volume (L), $n =$ number moles (mol), $c =$ concentration (mol L^{-})

2 a The amount of solute stays the same.

b $c\ V\ = c_2V_2$ where $c =$ concentration (mol L^{-}), $V =$ volume (L)

c As long as the units of volume are the same on both sides of the equation, it doesn't matter what units are ued.

3 a ±5%

b ±02% to ±05%

c ±02% to ±05%

d ±02% to ±05%

e ±10% (at least)

4 a $c = \frac{n}{V} = \frac{0\ 20}{0\ 100} = 20\ \text{mol L}^{-}\ \text{(2sf)}$

b $c = \frac{n}{V} = \frac{1\ 5}{2\ 00} = 075\ \text{mol L}^{-}\ \text{(2sf)}$

c $c = \frac{n}{V} = \frac{0\ 51}{0\ 250} = 20\ \text{mol L}^{-}\ \text{(2sf)}$

d $\text{Na}_2\text{CO}_3\quad n = \frac{m}{MM} = \frac{31\ 8}{(2 \times 22\ 99) + 12\ 01 + (3 \times 16\ 00)} = 03000\ \text{mol}$

$c = \frac{n}{V} = \frac{0\ 3000}{0\ 250} = 120\ \text{mol L}^{-}\ \text{(3sf)}$

e Ba(OH)$_2$ $n = \frac{m}{MM} = \frac{5\ 04}{137\ 3 + (2 \times 16\ 00) + (2 \times 1\ 008)} = 002941$ mol

$c = \frac{n}{V} = \frac{0\ 02941}{0\ 500} = 00588$ mol L^{-} (3sf)

f AgNO$_3$ $n = \frac{m}{MM} = \frac{11.9}{107\ 9 + 14\ 01 + (3 \times 16\ 00)} = 00700$ mol

$c = \frac{n}{V} = \frac{0\ 0700}{0\ 100} = 0700$ mol L^{-} (3sf)

5 **a** $n = c \times V = 0400 \times 200 = 0800$ mol (3sf)

b $n = c \times V = 0500 \times 0250 = 0125$ mol (3sf)

c $n = c \times V = 250 \times 10^{-3} \times 0250 = 625 \times 10^{-4}$ mol (3sf)

6 **a** $n = c \times V = 0200 \times 500 = 100$ mol

$m = n \times MM = 1.00 \times (1008 + 3545) = 36458 = 365$ (3sf)

b $n = c \times \mathrm{V} = 1.50 \times 0250 = 0375$ mol

$m = n \times MM = 0375 \times (2299 + 1600 + 1008) = 14999 = 150$ g (3sf)

c $n = c \times \mathrm{V} = 725 \times 10^{-2} \times 20 = 0145$ mol

$m = n \times MM = 0145 \times ((2 \times 2299) + 12.01 + (3 \times 1600)) = 15.37 = 154$ g (3sf)

7 **a** $n = c \times V = 0106 \times 0025 = 00027$ mol (2sf)

b $n = c \times V = 0048 \times 00072 = 3.5 \times 10^{-4}$ mol (2sf)

c $n = c \times V = 0055 \times 0050 = 00028$ mol (2sf)

d $n = c \times V = 036 \times 1.83 = 066$ mol (2sf)

8 **a** $m = n \times MM = 00027 \times (2299 + 1600 + 1008) = 0.11$ g (2sf)

c $m = n \times MM = 00028 \times (2431 + (2 \times 3545)) = 027$ g (2sf)

9 **a** $n = c \times V = 00330 \times 00177 = 584 \times 10^{-4}$ mol (3sf)

ii CaCl$_2$, thus $n(\mathrm{CaCl_2}) = n(\mathrm{Ca^{2+}}) = 584 \times 10^{-4}$ mol (3sf)

iii $n(\mathrm{Cl^-}) = 2 \times n(\mathrm{CaCl_2}) = 2 \times 584 \times 10^{-4} = 000117$ mol (3sf)

b $n = c \times V = 0.110 \times 00384 = 000422$ mol (3sf)

ii Na$_2$SO$_4$, thus $n(\mathrm{SO_4^{2-}}) = n(\mathrm{Na_2SO_4}) = 000422$ mol (3sf)

iii $n(\mathrm{Na^+}) = 2 \times n(\mathrm{Na_2SO_4}) = 2 \times 000422 = 000844$ mol (3sf)

10 **a** $c\ V = c_2 V_2$

$0106 \times 50 = c_2 \times 500$

$c_2 = 00106$ mol L^{-} (3sf)

b $c\ V = c_2 V_2$

$0223 \times 0025 = c_2 \times 1.00$

$c_2 = 00558$ mol L^{-} (2sf)

c $c\ V\ = c_2V_2$

$0304\ \times 100 = c_2 \times 250$

$c_2 = 0122 \text{ mol L}^{-}\ \ (3\text{sf})$

d $c\ V\ = c_2V_2$

$152\ \times 10 = c_2 \times 500$

$c_2 = 0030 \text{ mol L}^{-}\ \ (2\text{sf})$

11 a $c\ V\ = c_2V_2$

$15 \times 25 = c_2 \times 100$

$c_2 = 375\%(\text{w/v})$

b $c\ V\ = c_2V_2$

$106\ \times 5 = c_2 \times 100$

$c_2 = 0053 \text{ g/100 mL}$

12 a Measuring cylinder

b Volumetric flask

13 a 20 ppm $= 2.0 \text{ g} / 10^6$ g, thus 1.0 g in 10^3 g. (10^3 g = 1kg)

$$n\left(\text{Pb}^{2+}\right) = \frac{m}{MM} = \frac{1}{207\ 2} = 000483 \text{ mol}$$

$\text{Pb(NO}_3)_2$, thus $n(\text{Pb(NO}_3)_2) = n(\text{Pb}^{2+}) = 000483$ mol

$m(\text{Pb(NO}_3)_2) = n \times MM = 000483\ \times (2072\ + (2 \times 14.01) + (6 \times 1600))\ = 160 \text{ g } (3\text{sf})$

b Adding a solid to the required volume of water may cause the fina volume to increase past 1 L. A better method would be to add the solid to a 1.00 L volumetric flask, add water to dissolve then make up to the graduated mark to ensure exactly 1.00 of final solution.

Worked examle .5

1 $\text{H}_2\text{SO}_4\text{(aq)} + 2\text{NaOH(aq)} \rightarrow \text{Na}_2\text{SO}_4\text{(aq)} + \text{H}_2\text{O(l)}$

$n(\text{NaOH}) = c \times \text{V} = 1.07 \times 0125\ = 01338 \text{ mol}$

$n(\text{H}_2\text{SO}_4) = \frac{1}{2} \times n(\text{NaOH}) = 05\ \times 01338\ = 006688\ = 00669 \text{ mol } (3\text{sf})$

2 $2\text{HNO}_3\text{(aq)} + \text{BaO(s)} \rightarrow \text{Ba(NO}_3)_2\text{(aq)} + \text{H}_2\text{O(l)}$

$$n\left(\text{BaO}\right) = \frac{m}{MM} = \frac{2\ 30}{137\ 3 + 16\ 00} = 001500 \text{ mol}$$

$n(\text{HNO}_3) = 2 \times n(\text{BaO}) = 2 \times 001500\ = 00300 \text{ mol}$

$$V\left(\ \ \text{NO}_3\right) = \frac{n}{c} = \frac{0\ 0300}{0\ 423} = 00709 \text{ mol L}^{-}\ \ (3\text{sf})$$

3 $2\text{NaOH(aq)} + \text{CuSO}_4\text{(aq)} \rightarrow \text{Cu(OH)}_2\text{(s)} + \text{Na}_2\text{SO}_4$

$n(\text{NaOH}) = c \times V = 0425\ \times 0025\ = 001063 \text{ mol}$

$n(\text{Cu(OH)}_2) = \frac{1}{2} \times n(\text{NaOH}) = 05\ \times 001063\ = 0005315 \text{ mol}$

$m(\text{Cu(OH)}_2) = n \times MM = 0005315\ \times (6355\ + (2 \times 1600)\ + (2 \times 1008))\ = 05186\ = 052 \text{ g } (2\text{sf})$

Check your understandng 86

1 Use the volume and molarity to calculate the number of moles using $n = c \times V$ where n = number moles (mo), c = molarity (mol L^{-}), V = volume(L). Then use $m = n \times MM$ where m = mass (g), n = number moles (mol), MM = molecular mass (g mol^{-}) to calculate he mass.

2 That enough of that reactant is present to allow the seond reactant to fully react.

3 $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

$n(Ba(OH)_2) = c \times V = 087 \times 0075 = 006525$ mol

$n(HNO_3) = 2 \times n(Ba(OH)_2) = 2 \times 006525 = 01305 = 013$ mol (2sf)

4 $2NaOH(aq) + CuSO_4(aq) \rightarrow Cu(OH)_2(aq) + Na_2SO_4(aq)$

$n(CuSO_4) = c \times V = 065 \times 0050 = 00325$ mol

$n(NaOH) = 2 \times n(CuSO_4) = 2 \times 00325 = 0065$ mol

$$V(NaOH) = \frac{n}{c} = \frac{0\ 065}{2\ 33} = 0028 \text{ mol L}^{-} \ (3sf)$$

5 **a** $Na_2SO_4(aq) + BaCl_2(aq) \rightarrow 2NaCl(aq) + BaSO_4(s)$

$n(BaCl_2) = c \times V = 0326 \times 0050 = 00163$ mol

$n(BaSO_4) = n(BaCl_2) = 0016$ mol (2sf)

b $Pb(NO_3)_2(aq) + 2NaBr(aq) \rightarrow 2NaNO_3(aq) + PbBr_2(s)$

$n(NaBr) = c \times V = 0509 \times 0025 = 00127$ mol

$n(PbBr_2) = 2 \times n(NaBr) = 2 \times 0127 = 0025$ mol (2sf)

c $Fe(NO_3)_3(aq) + 3NaOH \rightarrow Fe(OH)_3(aq) + 3NaNO_3(aq)$

$n(NaOH) = c \times V = 0087 \times 0075 = 0006525$ mol

$$n(Fe(OH)_3) = \frac{1}{3} \times n(NaOH) = \frac{1}{3} \times 0006525 = 00022 \text{ mol (2sf)}$$

6 **a** $m = n \times MM = 0016 \times (137.3 + 3207 + (4 \times 1600)) = 37$ g (2sf)

c $m = n \times MM = 00022 \times (5585 + (3 \times 1600) + (3 \times 1008)) = 024$ g (2sf)

7 **a** $NaCl(aq) + Ag^{+}(aq) \rightarrow AgCl(s) + Na^{+}$

$$n(AgCl) = \frac{m}{MM} = \frac{1\ 87}{107\ 9 + 35\ 45} = 00130 \text{ mol}$$

$n(AgCl) = n(Ag^{+}) = 00130$ mol

$$c = \frac{n}{V} = \frac{0\ 013}{0\ 050} = 026 \text{ mol L}^{-} \ (2sf)$$

b $Pb^{2+}(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2K^{+}(aq)$

$$n(PbI_2) = \frac{m}{MM} = \frac{3\ 19}{207\ 2 + (2 \times 126\ 9)} = 0006920 \text{ mol}$$

$n(PbI_2) = n(Pb^{2+}) = 0006920$ mol

$$c = \frac{n}{V} = \frac{0\ 006920}{0\ 025} = 028 \text{ mol L}^{-} \ (2sf)$$

8 $m(\text{Pb}^{2+}) = n \times MM = 0006920 \times 2072 = 1.43$ g

Thus1.44 g in 25 mL → $144 \times 4 = 576$ g/100 mL

9 $\text{AgNO}_3\text{(aq)} + \text{NaCl(aq)} \rightarrow \text{AgCl(s)} + \text{NaNO}_3\text{(aq)}$

$n(\text{AgNO}_3) = c \times V = 0.13 \times 0040 = 00052$ mol

$n(\text{NaCl}) = c \times V = 022 \times 0020 = 00044$ mol

As mole ratio = 1 1, the silver nitrate is in excess and the sodium chloride is the limiting reagent.

$n(\text{AgCl}) = n(\text{NaCl}) = 00044$ mol (2sf)

Chapter review questions

1 Concentration is the amount of solute (liquid or solid) in a volume of solvent/solution. Solubility is the ability of a solute to dissolve in a solvent.

2 **a** Volume of solute per 100 mL of solution – percent weight volume.

b Grams of solid or liquid solute per million grams of solution or (for gases) molecules per million molecules.

c Concentration of a solution – the amount of substance (in moles) per volume of solution/solvent.

3 $c = \dfrac{n}{V}$

a $n = c \times V$

b $V = \dfrac{n}{c}$

4 **a** measuring cylinder

b pipette/burette/volumetric flask

5 Calculations involving quantitative aspects of chemical reactions including mass, concentration, volume and moles.

6 Calculate the number of moles of the substance.

7 Calculate the number of moles of each substance, determine which is the limiting reagent, then use the moles of the limiting reagent to determine the moles of the product formed.

8 6 ppm = 6 g/10^6 g

%(w/w) = g/100 g

Thus 6×10^{-4} g/100 g

9 **a** 15%(w/v) = 15 g/100 mL, thus 1.5/4 = 0375 g/25 mL Thus0.375 g of iodine.

b 15 %(w/w) = 15 g/100 g, thus 10/19.8 = 5051, thus 1./5.051 = 0297 g/198 g. Thus 0.297 g iodine

10 **a** $c = \dfrac{n}{V} = \dfrac{0\ 56}{0\ 250} = 22 \text{ mol L}^{-}$ (2sf)

b $c = \dfrac{n}{V} = \dfrac{7\ 3 \times 10^{-2}}{0\ 100} = 073 \text{ mol L}^{-}$ (2sf)

c $n = \frac{m}{MM} = \frac{23\ 4}{107\ 9 + 14\ 01 + (3 \times 16\ 00)} = 0138\text{ mol}$

$c = \frac{n}{V} = \frac{0\ 138}{0\ 500} = 0276\text{ mol L}^{-}\ \ (3\text{sf})$

d $n = \frac{m}{MM} = \frac{4\ 22 \times 10^{-3}}{207\ 2 + (2 \times 14\ 01) + (6 \times 16\ 00)} = 1274\ \times 10^{-5}\text{ mol}$

$c = \frac{n}{V} = \frac{1\ 274 \times 10^{-5}}{0\ 250} = 5.10 \times 10^{-5}\text{ mol L}^{-}\ \ (3\text{sf})$

11 a $n = c \times V = 00306\ \times 0025\ = 77\ \times 10^{-4}$ mol (2sf)

b $n = c \times V = 0425\ \times 0250\ = 0106$ mol (3sf)

c $n = c \times V = 855\ \times 10^{-3} \times 0050\ = 43\ \times 10^{-4}$ mol (2sf)

12 $c\ V\ = c_2V_2$

$0236\ \times 25 = c_2 \times 500$

$c_2 = 0012\text{ mol L}^{-}\ \ (2\text{sf})$

13 a $Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3$

$c(KI) = \frac{n}{V} = \frac{0\ 44}{0\ 250} = 176\text{ mol L}^{-}\ \ (3\text{sf})$

$n(\text{KI in 25 mL}) = c \times V = 1.76 \times 0025\ = 0044\text{ mol}$

$n(Pb(NO_3)_2) = \frac{1}{2} \times n(KI) = 05\ \times 0044\ = 0022\text{ mol}$

$V = \frac{n}{c} = \frac{0\ 022}{0\ 644} = 0034\text{ L}$

b $HCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + HNO_3(aq)$

$n(AgNO_3) = n(HCl) = c \times V = 0.13 \times 0050\ = 00065\text{ mol}$

$V = \frac{n}{c} = \frac{0\ 0065}{0\ 021} = 031\text{ L}\ (2\text{sf})$

14 a $c\ V\ = c_2V_2$

$00177\ \times 25 = c_2 \times 500$

$c_2 = 89\ \times 10^{-4}\text{ mol L}^{-}$

ii $\text{ppm} = \text{g}/10^6\ \text{g}/10^6\ \text{mL}/10^3\ \text{L}$

Thus $1\text{ L} \rightarrow 10^3\text{ L} = \times 1000$

$m(Pb(NO_3)_2) = n \times MM = 89\ \times 10^{-4} \times (2072\ + (2 \times 14.01) + (6 \times 1600))\ = 0295\text{ g}$

Thus$0.295\ \times 1000 = 295\text{ g}/10^3\text{ L} = 295\text{ ppm}$

b Mole ratio $Pb(NO_3)_2 : Pb^{2+} = 1\ \ 1$ Thus lead ions will have the same concentratin as the lead nitrate.

$89\ \times 10^{-4}\text{ mol L}^{-}$

ii 295 ppm

15 400%(v/v) $= 400$ mL/100 mL

Using a pipette, measure out $(4.0 \times 5 =)$ 200 mL of acetone and transfer to a 500 mL volumetric flask. Add water until the meniscus of the solution is on the bottom of the graduated mark. Shake the solution until homoenous.

16 a $130 \times 079 = 1027$ g. Thus $c = 1027$ g/100 mL

b $1027 \times 10 = 1027$ g/L

c %(w/v) = g/100 mL Thus 0.27%(w/v)

17 $c\ V = c_2V_2$

$107 \times V = 0054 \times 1.00$

$V = 0050$ L

Using a pipette, measure 50mL of the 1.07 mol L^{-} zinc nitrate solution,add to a 1.00 L volumetric flask, make up to the graduated mark with water, until the bottom of th meniscus is on the graduated mark.

18 $Al_2(SO_4)_3$ thus aluminium slfate : aluinium ions : sulfate ions $= 1 : 2 : 3$

a $2 \times 25 \times 10^{-4} = 50 \times 10^{-4}$ mol L^{-}

b $3 \times 25 \times 10^{-4} = 75 \times 10^{-4}$ mol L^{-}

19 050 p pm = 050 g/10^{6} g or 10^{6} mL or 10^{3} L = 050/1000 $= 50 \times 10^{-4}$ g/L

Measure out 5.0×10^{-4} g of chromium chloride on a very sensitive lectronic balance. (Admittedly this is very difficult – could increase masses and volumes, but will end up with either a small mass or a very large volum – approx. 1000L). Add to a 1.00 L volumetric flask, add enough water to dissolve the chromium chloride and shake until the solid is fully dissolved. Add water until the bottom of the meniscus of the solution is on the graduated mark. Shake the solution to make homogenous.

20 a Dilution is by a factor of10, so 10.0 mL of Solution A made up in a 100. mL volumetric flask.

b Dilution is by a factor of40, so 25.0 mL of SolutionA, made up in a 1.0 L volumetric flask.

c Dilution is by a factor of5, so 100.0 mL of Solution A made up in a 500. mL volumetric flask.

21 $Ca(OH)_2(aq) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(aq)$

$n(HCl) = c \times V = 0102 \times 00813 = 0008293$ mol

$n(Ca(OH)_2) = 1/2 \times n(HCl) = 05 \times 0008293 = 0004146$ mol

$$c = \frac{n}{V} = \frac{0004146}{00250} = 0166 \text{ mol L}^{-} \text{ (3sf)}$$

$m = n \times M = 0166 \times (4008 + (2 \times 1600) + (2 \times 1008)) = 123$ g/L (3sf)

22 $CuSO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + Cu(OH)_2(s)$

Assume the copper sulfate is the imiting reagent.

$n(CuSO_4) = c \times V = 1.21 \times 00250 = 00302$ mol

$n(CuSO_4) = n(Cu(OH)_2) = 00302$ mol

$m = n \times MM = 00302 \times (6355 + (2 \times 1600) + (2 \times 1008)) = 295$ g (3sf)

23 $NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq)$

$$n(AgCl) = \frac{m}{MM} = \frac{0\ 58}{107\ 9 + 35\ 45} = 0004046\text{ mol}$$

$$n(AgNO_3) = n(AgCl) = 0004046\text{ mol}$$

$$c(\text{dilute}) = \frac{n}{V} = \frac{0\ 004046}{0\ 0500} = 008092\text{ mol L}^{-}$$

$$c\ V\ = c_2V_2$$

$$c\ \times 25 = 008092\ \times 500$$

$$c\ = 1.618 = 16\text{ mol L}^{-}\ \ (2\text{sf})$$

$$m = n \times MM = 1.618 \times (1079\ + 14.01 + (3 \times 1600))\ = 2749\text{ g L}^{-}\ = 270\text{ g/100 mL}^{-}\ \ (2\text{sf})$$

24 $MgSO_4(aq) + 2NaOH(aq) \rightarrow Mg(OH)_2(s) + Na_2SO_4(aq)$

a $n\left(Mg(OH)_2\right) = \dfrac{m}{MM} = \dfrac{1\ 72}{24\ 31 + (2 \times 16\ 00) + (2 \times 1\ 008)} = 002949\text{ mol}$

$$n(MgSO_4) = n(Mg(OH)_2) = n(Mg^{2+})\ 002949\text{ mol}$$

$$c\left(Mg^{2+}\right) = \frac{n}{V} = \frac{0\ 02949}{0\ 0500} = 05898\ = 059\text{ mol L}^{-}\ \ (2\text{sf})$$

b $m = n \times MM = 05898\ \times 2431\ = 1434\ = 14\text{ g L}^{-}\ \ (2\text{sf})$

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 9: Gas laws

Worked example .1

1 $n = \frac{V}{V_M} = \frac{233}{2271} = 01026 \text{ mol}$

$m = n \times MM = 01026 \times (12.01 + (4 \times 1008)) = 1646 = 165 \text{ g (3sf)}$

2 $n = \frac{m}{MM} = \frac{23\ 0}{1401 + (3 \times 1\ 008)} = 1350 \text{ mol}$

$V = n \times V_M = 1.350 \times 2479 = 3347 = 335 \text{ L (3sf)}$

Worked example .2

1 $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

$n(Na) = \frac{m}{MM} = \frac{187}{2299 + 1600 + 1008} = 004675 \text{ mol}$

$n(H_2) = \frac{1}{2} \times n(Na) = 05 \times 004675 = 002338 \text{ mol}$

$V = n \times V_M = 002338 \times 2479 = 05795 = 0580 \text{ L (3sf)}$

2 $n(HCl) = \frac{V}{V_M} = \frac{13\ 4}{2271} = 05900 \text{ mol}$

$2KCl(aq) + H_2SO_4(aq) \rightarrow 2HCl(g) + K_2SO_4(aq)$

$n(KCl) = n(HCl) = 05900 \text{ mol}$

$m(KCl) = n \times MM = 05900 \times (3910 + 3545) = 4398 = 440 \text{ g (3sf)}$

Check your understanding 9.1/9.2

1 The volume of a gas changes with temperature and pressure, both of which compress or expand the gas particles.

2 Use the formula $n = \frac{V}{V_M}$ to convert the volume of gas to a number of moles. V_M depends upon the temperature of the gas. Use $m = n \times MM$ to convert the number of moles (n) to a mass (m) by multiplying by the molecular mass (MM)

3 **a** $2 \times 250 = 500$ mL

ii 250 mL

b 500 mL

ii Nothing, as the law of combining volumes only applies to gases.

4 **a** $n = \dfrac{V}{V_M} = \dfrac{2\ 042}{22\ 71} = 008992\ \text{mol}$

$m = n \times MM = 008992\ \times (3207\ + (2 \times 1600))\ = 5761\ \text{g (4sf)}$

b $n = \dfrac{m}{MM} = \dfrac{1\ 32}{12\ 01 + (2 \times 16\ 00)} = 002999\ \text{mol}$

$V = n \times V_M = 002999\ \times 2479\ = 07435\ = 0744\ \text{L (3sf)}$

5 $\text{Mg(s)} + 2\text{HCl(aq)} \rightarrow \text{H}_2\text{(g)} + \text{MgCl}_2\text{(aq)}$

$n(\text{Mg}) = \dfrac{m}{MM} = \dfrac{1\ 0}{24\ 31} = 004114\ \text{mol}$

$n(\text{Mg}) = n(\text{H}_2) = 004114\ \text{mol}$

$V = n \times V_M = 004114\ \times 2271\ = 09342\ = 093\ \text{L (2sf)}$

6 $\text{CaCO}_3\text{(s)} \rightarrow \text{CaO(s)} + \text{CO}_2\text{(g)}$

$n(\text{CaCO}_3) = \dfrac{m}{MM} = \dfrac{0\ 33}{40\ 08 + 12\ 01 + (3 \times 16\ 00)} = 0003297\ \text{mol}$

$V = n \times V_M = 0003297\ \times 2479\ = 008173\ = 0082\ \text{L (2sf)}$

7 $\text{SO}_2\text{(g)} + \text{CaO(s)} \rightarrow \text{CaSO}_3\text{(s)}$

$n(\text{SO}_2) = \dfrac{V}{V_M} = \dfrac{1\ 00 \times 10^3}{24\ 79} = 4034\ \text{mol}$

$n(\text{SO}_2) = n(\text{CaO}) = 4034\ \text{mol}$

$m(\text{CaO}) = n \times MM = 4034\ \times (4008\ + 1600)\ = 2262 = 2260\ \text{g (3sf)}$

8 $\text{Na}_2\text{CO}_3\text{(aq)} + 2\text{HCl(aq)} \rightarrow 2\text{NaCl(aq)} + \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$

$n(\text{Na}_2\text{CO}_3) = c \times V = 233\ \times 0025\ = 005825\ \text{mol}$

$n(\text{HCl}) = c \times V = 1.42 \times 0100\ = 0142\ \text{mol}$

If .05825 mol sodium carbonate reacts, then 2 × 005825 = 01165 mol hydrochloric aid reacts. Thus sodium carbonate is the limting reagent.

$n(\text{Na}_2\text{CO}_3) = n(\text{CO}_2) = 005825\ \text{mol}$

$V = n \times V_M = 005825\ \times 2479\ = 1444\ = 14\ \text{L (2sf)}$

9 $\text{HCl(aq)} + \text{NH}_3\text{(g)} \rightarrow \text{NH}_4\text{Cl(aq)}$

$n(\text{NH}_3) = \dfrac{V}{V_M} = \dfrac{5\ 33}{24\ 79} = 02150\ \text{mol}$

$n(\text{NH}_3) = n(\text{HCl}) = 02150\ \text{mol}$

$V = \dfrac{n}{c} = \dfrac{0\ 2150}{1\ 37} = 01569\ = 0157\ \text{mol L}^{-}\ \text{(3sf)}$

10 $\text{FeS(s)} + 2\text{HCl(aq)} \rightarrow \text{H}_2\text{S(g)} + \text{FeCl}_2\text{(aq)}$

$n(\text{FeS}) = \dfrac{m}{MM} = \dfrac{1\ 24}{55.85 + 32\ 07} = 00141\ \text{mol}$

$n(\text{HCl}) = c \times V = 1.04 \times 0030 = 00312$ mol

If 00141 mol iron sulfide react, then $0.0141 \times 2 = 00282$ mol hydrochloric aid needed. Thus iron sulfide is the limitingreagent.

$n(\text{FeS}) = n(\text{H}_2\text{S}) = 00141$ mol

$V = n \times V_M = 00141 \times 2271 = 03202 = 032$ L (2sf)

Worked examle .3

1 $P V = P_2V_2$

$101 \times 23 = 245 \times V_2$

$V_2 = 09482 = 095$ L (2sf)

2 $P V = P_2V_2$

$550 \times 500 = 3600 \times P_2$

$P_2 = 7639 = 764$ kPa (3sf)

Worked examle .4

1 $\frac{V}{T} = \frac{V_2}{T_2}$

$$\frac{250}{25 + 27315} = \frac{V_2}{155 + 27315}$$

$V_2 = 359$ L (3sf)

2 $\frac{V}{T} = \frac{V_2}{T_2}$

$$\frac{125}{18 + 27315} = \frac{0750}{T_2}$$

$T_2 = 1747 - 27315 = -9846 = -98$°C (2sf)

Check your understandng 93a

1 **a** Use any units of pressure or volume, as long as they are the same on both sides.

b Temperature must be in Kelvin where 0°C = −27315 K

2 **a** $64 + 27315 = 33715$ K

b $3715 + 27315 = 310.3$ K

c $-186 + 27315 = 8715$ K

d $550 + 27315 = 82315$ K

e $-1036 + 27315 = 16955$ K

3 $P V = P_2V_2$

$733 \times 250 = P_2 \times 663$

$P_2 = 2764 = 276$ kPa (3sf)

4 **a** $P\ V\ = P_2V_2$

$344 \times 505 = 83.1 \times V_2$

$V_2 = 2090$ mL (3sf)

b $P\ V\ = P_2V_2$

$87 \times 46\ = P_2 \times 65$

$P_2 = 6157\ = 62$ kPa (2sf)

c $P\ V\ = P_2V_2$

$042\ \times 327 = 0846\ \times V_2$

$V_2 = 1623\ = 160$ mL (2sf)

d $P\ V\ = P_2V_2$

$231 \times 455\ = P_2 \times 0653$

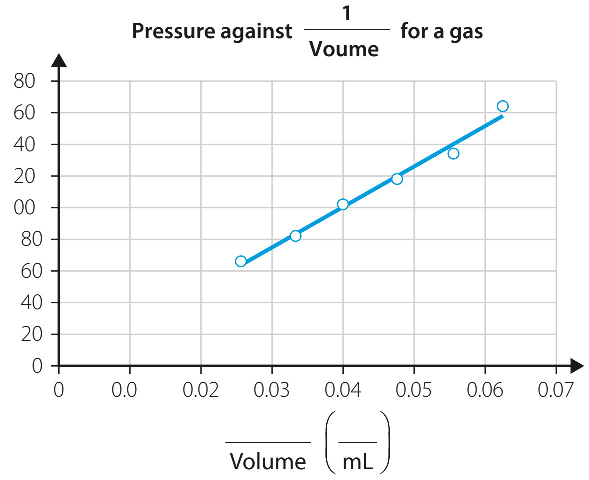
$P_2 = 1610$ Pa $= 1610$ kPa $\ = 161$ kPa (3sf)

5 **a** $25 \times 102 = 2550$, $21 \times 118 = 2478$, $18 \times 134 = 2412$, $16 \times 164 = 2624$, $30 \times 82 = 2460$, $39 \times 66 = 2574$

Within experimentl error, $P \times V$ is reasonably onsistent. Thus these results are consistent with Boyle's law.

b Pressure is linear when plotted against $\frac{1}{V}$, or pressure is linear against the iverse of volume.

c



6 **a** $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{1\ 7}{25 + 273\ 15} = \frac{V_2}{250 + 273\ 15}$

$V_2 = 30$ L (2sf)

b $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{1\ 7}{25 + 273\ 15} = \frac{0\ 900}{T_2}$

$T_2 = 157.8 - 27315\ = -1153\ = -120$°C (2sf)

7 **a** $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{500}{200} = \frac{V_2}{500}$

$V_2 = 1250 \text{ mL (3sf)}$

b $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{1\ 5}{350} = \frac{0450}{T_2}$

$T_2 = 105 - 27315 \ = -16815 \ = -170°\text{C (2sf)}$

c $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{250}{22 + 27315} = \frac{V_2}{85 + 27315}$

$V_2 = 3034 \ = 300 \text{ mL (2sf)}$

d $\frac{V}{T} = \frac{V_2}{T_2}$

$\frac{550}{180 + 27315} = \frac{1350}{T_2}$

$T_2 = 1112 - 27315 \ = 839.1 = 839°\text{C (3sf)}$

Worked examle .5

1 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$\frac{101 \times 0750}{18 + 27315} = \frac{P_2 \times 167}{92 + 27315}$

$P_2 = 5689 \ = 57 \text{ kPa (2sf)}$

2 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$\frac{175 \times 1175}{32 + 27315} = \frac{V_2 \times 46}{143}$

$V_2 = 2095 \ = 21 \text{ L (2sf)}$

Check your understandng 93b

1 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

2 Units for the same quantity need to be the same on oth sides of the equation.

3 When measured at the same temperature and pressure, equal volumes of different gases contain the same number o moles.

A mole of any gas has the same volume at constant temperature and pressure as a mole of any other gas.

At constant pressure and temperature, the volume of a gas is proportional to the number of moles presnt; $V = kn$

4 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{90\ 2 \times 15\ 1}{292} = \frac{P_2 \times 182\ 2}{428}$$

$P_2 = 1096\ \ = 110$ kPa (3sf)

5 **a** $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{101\ 3 \times 1\ 5\ 0}{300} = \frac{V_2 \times 53\ 3}{450}$$

$V_2 = 4276\ \ = 43$ L (2sf)

b $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{131\ 7 \times 500}{280} = \frac{V_2 \times 1200}{505}$$

$V_2 = 9897\ \ = 990$ kPa (3sf)

c $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{101 \times 650}{25 + 273\ 15} = \frac{87 \times 1350}{T_2}$$

$T_2 = 5334\ \ - 27315\ \ = 260$°C (2sf)

d $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{101 \times 1\ 30}{80 + 273\ 15} = \frac{V_2 \times 350}{271 + 273\ 15}$$

$V_2 = 05780\ \ = 058$ L (2sf)

e $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{53 \times 23}{17 + 273\ 15} = \frac{P_2 \times 1650}{344}$$

$P_2 = 08759\ \ = 088$ kPa (3sf)

f $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{0\ 086 \times 5\ 67}{298} = \frac{4\ 23 \times 0\ 158}{T_2}$$

$T_2 = 4084\ \ -27315\ \ = 140$°C (2sf)

6 $$\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$$

$$\frac{5000 \times 25}{295} = \frac{P_2 \times 1000}{282}$$

$P_2 = 119.5 = 120$ kPa (2sf)

7 Avogadro's law says at constant pressure and temperature; the volume of a gas is proportional to the number of moles. If the mass of carbon dioxide is tripled, then the number of moles will triple, and thus the volume will be tripled. $3 \times 750 = 2250$ mL

8 $$\frac{P}{T} = \frac{P_2}{T_2}$$

$$\frac{53}{25 + 273\ 15} = \frac{P_2}{250 + 273\ 15}$$

$P_2 = 930\ = 93$ kPa (2sf)

9 a $$\frac{P}{T} = \frac{P_2}{T_2}$$

$$\frac{100}{17 + 273\ 15} = \frac{P_2}{90 + 273\ 15}$$

$P_2 = 1252\ = 130$ kPa (2sf)

b $$\frac{P}{T} = \frac{P_2}{T_2}$$

$$\frac{100}{17 + 273\ 15} = \frac{70}{T_2}$$

$T_2 = 203105\ - 27315\ = -70$°C (2sf)

10 a By Avogadro's law, equal volumes at the same temperature and pressure have equal moles, so 37.5 mL/100 = 375 mL Thus $n(CO_2) = 375\ \times 42\ \times 10^{-4} = 000158$ mol (3sf)

b $\frac{3\ 8}{1\ 4} = 2\ 7\ 1$, thu $0.50\ \times 2.71 = 1.36$ L

Worked examle .6

1 $PV = nRT$

$360 \times V = 062\ \times 8314\ \times (295 + 27315)$

$V = 8.135 = 8.1$ L (2sf)

2 $PV = nRT$

$235 \times 345\ = n \times 8314\ \times (35 + 27315)$

$n = 03165\ = 032$ mol (2sf)

3 $PV = nRT$

$P \times 0500\ = 0050\ \times 8314\ \times (18 + 27315)$

$P = 2421\ = 240$ kPa (2sf)

Check your understandng 94

1 Combining oyle's, Charles and Avogadro's law to get $\frac{PV}{nT}$ = constant. This constant is the same for all gases and is calculated from the molar volume of a gas at 0° and 100 kPa, which is 22.71 L. Thus

$$\frac{100 \times 22\ 71}{1 \times 273\ 15} = 8314 \text{ kPa L K}^{-} \text{ mol}^{-}$$

2 **a** $PV = nRT$

$20 \times 50 = n \times 8314 \times 300$

$n = 00401 = 0040 \text{ mol (2sf)}$

b $PV = nRT$

$304 \times V = 020 \times 8314 \times 450$

$V = 2461 = 25 \text{ L (2sf)}$

c $PV = nRT$

$P \times 12 = 0055 \times 8314 \times (303 + 27315)$

$P = 2195 = 22 \text{ kPa (2sf)}$

d $PV = nRT$

$533 \times 0500 = 0033 \times 8314 \times T$

$T = 9713 - 27315 = -176\text{°C (3sf)}$

e $PV = nRT$

$2026 \times 0505 = n \times 8314 \times (26 + 27315)$

$n = 4.114 \times 10^{-4} = 4.1 \times 10^{-4} \text{ mol (2sf)}$

f $PV = nRT$

$0350 \times V = 0010 \times 8314 \times (35 + 27315)$

$V = 7320 = 73 \text{ L (2sf)}$

3 $PV = nRT$

$221 \times 10^{2} \times V = 045 \times 8314 \times (22 + 27315)$

$V = 4997 = 50 \text{ L (2sf)}$

4 $PV = nRT$

$$P \times 47 = \frac{3\ 51 \times 10^{3}}{12\ 01 + (2 \times 16\ 00)} \times 8314 \times 290$$

$P = 4091 = 4100 \text{ kPa (2sf)}$

5 $PV = nRT$

$83 \times 1.5 = n \times 8314 \times 288$

$n = 005200 \text{ mol}$

$m = n \times MM = 005200 \times (12.01 + (4 \times 1008)) = 08342 = 083 \text{ g (2sf)}$

6 $PV = nRT$

$194 \ \times 10.0 = n \times 8314 \ \times (25 + 27315)$

$n = 007826$ mol

$MM = \frac{m}{n} = \frac{5\ 03}{0\ 07826} = 6427 \ = 64$ g mol^{-}

7 $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$n(Zn) = \frac{m}{MM} = \frac{3\ 6}{65\ 38} = 005506$ mol

$n(H_2) = n(Zn) = 005506$ mol

PV = nRT

$150 \times V = 005506 \ \times 8314 \ \times (18 + 27315)$

$V = 08885 \ = 089$ L (2sf)

8 $PV = nRT$

$110 \times 0250 \ = n \times 8314 \ \times (30 + 27315)$

$n = 00109$ mol

$2NaOH(aq) + CO_2(g) \rightarrow Na_2CO_3(aq) + H_2O(l)$

$n(NaOH) = 2 \times n(CO_2) = 2 \times 00109 \ = 002182$ mol

$V = \frac{n}{c} = \frac{0\ 02182}{2\ 05} = 00106 \ = 0011$ L (2sf)

Chapter review questions

1 1 volume of nitrogen gas reacts with 3 volumes of hydrogen gas to produce 2 volumes of ammonia gas.

2 Use $n = \frac{m}{MM}$ to calculate the number of moles (n) of copper carbonate after determining the molar mass (MM) copper carbonate. → Use the mole ratio in the balanced equation to determine the number of moles of carbon dioxide that form. → Use $V = n \times V_M$ where V_M is the molar volume of a gas at a given temperature.

3 Use the equation $P\ V \ = P_2V_2$ and substitute the initial and final pressure and initial volume, then determine V_2

4 **a** $\frac{V}{T} = \frac{V_2}{T_2}$

b Any units can be used as long as the same units are used for the same quantity on both sides of the equation.

5 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

6 $PV = nRT$, where P = pressure (kPa), V = volume (L), n = number moles (mol), R = gas constant, T = temperature (K)

7 $FeS(s) + 2HCl(aq) \rightarrow FeCl_2(aq) + H_2S(g)$

$n(FeS) = \frac{m}{MM} = \frac{0\ 38}{55\ 85 + 32\ 07} = 0004322$ mol

$V = n \times V_M = 0004322 \ \times 2271 \ = 00982 \ = 0098$ L (2sf)

8 $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

$$n(CO_2) = \frac{V}{V_M} = \frac{0\ 025}{24\ 79} = 000101 \text{ mol}$$

$$n(Ca(OH)_2) = n(CO_2) = 000101 \text{ mol}$$

$$V = n \times V_M = 000101 \ \times 2479 \ = 0025 \text{ L (2sf)}$$

9 **a** Hydrogen gas and chlorine gas must have both two atoms in them (H_2 and Cl_2) Avogado's law states 1 molecule + 1 molecule → 2 molecles. In order to form two molecules at the end, each containing H and Cl atoms, there must be two atoms of hydrogen and two atoms of chlorine to start with.

b If two molecules of steam are formed from one molecule of oxygen and two molecules of water, each molecule of steam must have 1 oxygen atom and 2 hydrogen molecules – hence the formula is H_2O. As each molecule of steam has 1 oxygen atom, the oxygen molecule must have 2 atoms.

10 $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

$$n(Mg) = \frac{m}{MM} = \frac{4\ 51}{24\ 31} = 01855 \text{ mol}$$

$$n(H_2SO_4) = c \times V = 2.51 \times 0075 \ = 01883 \text{ mol}$$

As mole ratio = 1 1 the sulfuric acid is in excess and the magnesium is the limiting reagent.

$$n(H_2) = n(Mg) = 01855 \text{ mol}$$

$$V = n \times V_M = 01855 \ \times 2479 \ = 459 \ = 46 \text{ L (2sf)}$$

11 $P\ V \ = P_2V_2$

$$1733 \ \times 650 = P_2 \times 3250$$

$$P_2 = 3466 \ = 347 \text{ kPa (3sf)}$$

12 $\frac{V}{T} = \frac{V_2}{T_2}$

$$\frac{50}{(22 + 273\ 15)} = \frac{V_2}{-35 + 273\ 15}$$

$$V_2 = 4034 \ = 40 \text{ L (2sf)}$$

13 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{101 \times 10}{20 + 273\ 15} = \frac{V_2 \times 5\ 0}{-50 + 273\ 15}$$

$$V_2 = 154 = 150 \text{ m}^3 \text{ (2sf)}$$

14 $PV = nRT$

$$P \times 0500 \ = \frac{2\ 0}{39\ 95} \times 8314 \ \times (18 + 27315)$$

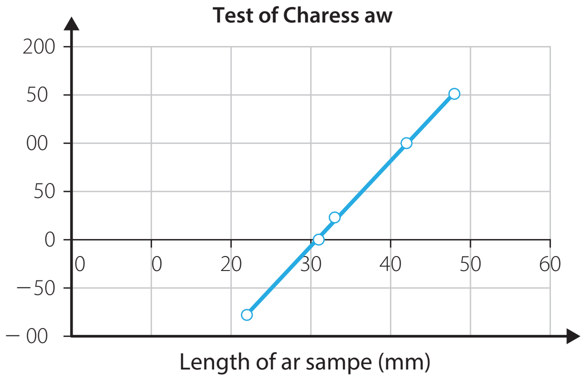
$$P = 242 = 240 \text{ kPa (2sf)}$$

15 $PV = nRT$

$$P \times 0001 \ = \frac{10^{12}}{3\ 022 \times 10^{23}} \times 8314 \ \times (20 + 27315\ )$$

$$P = 8 \times 10^{-6} \text{ kPa}$$

16 a The linear relationship between temperature and pressure verifies Charles's law.



b $\frac{22}{19515} = 0113$, $\frac{31}{27315} = 0113$, $\frac{33}{29615} = 0111$, $\frac{42}{37315} = 0115$, $\frac{48}{42415} = 0113$

As all the numbers found by $\frac{P}{T}$ are the same, Charles's law is verified.

17 $\frac{P_1V_1}{T} = \frac{P_2V_2}{T_2}$

$$\frac{101 \times 100}{17 + 27315} = \frac{475 \times 23}{T_2}$$

$T_2 = 314 - 27315 = 41°\text{C}$

18 $PV = nRT$

$667 \times 25 = n \times 8314 \times 310$

$n = 00647$ mol

$m = n \times MM = 00647 \times (14.01 + 1600) = 1.94 = 1.9$ g (2sf)

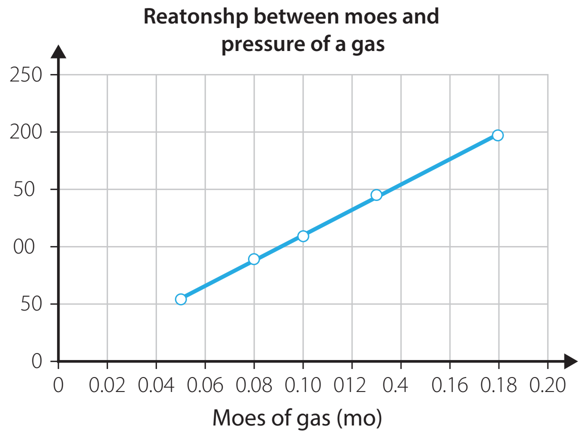
19 $n(\text{SO}_2) = \frac{m}{MM} = \frac{321}{3207 + (2 \times 1600)} = 00501$ mol

$n(\text{NO}) = \frac{m}{MM} = \frac{240}{1401 + 1600} = 007997$ mol

$n(\text{HCl}) = \frac{m}{MM} = \frac{365}{1008 + 3545} = 01001$ mol

$n(\text{CO}_2) = \frac{m}{MM} = \frac{572}{1201 + (2 \times 1600)} = 01300$ mol

$n(\text{NH}_3) = \frac{m}{MM} = \frac{306}{1401 + (3 \times 1\,008)} = 01796$ mol



Graph shows that for constant volume and temperature, the pressure is proportional (shown by the linear trendline) to the number of moles which obeys the ideal gas law.

20 a $\frac{P}{T} = \frac{P_2}{T_2}$

$\frac{755}{45 + 273\ 15} = \frac{455}{T_2 + 273\ 15}$

$T_2 = -814°\text{C} = -81°\text{C (2sf)}$

b $\frac{P}{T} = \frac{P_2}{T_2}$

$\frac{347}{10 + 273\ 15} = \frac{P_2}{85 + 273\ 15}$

$P_2 = 4387 = 440\ \text{kPa (2sf)}$

21 a $3072 \times 0998 = 30659$ g water in the can

Mass can $= 3880 - 30659$

$= 81.41$ g

b $\text{m(He)} = 8166 - 81.41 = 0253\ \text{g}$

$n(\text{He}) = \frac{\text{m}}{\text{MM}} = \frac{0\ 253}{4\ 033} = 00631\ \text{mol}$

$\text{m}(\text{CH}_4) = 8252 - 81.41 = 1.11\ \text{g}$

$\text{n}(\text{CH}_4) = \frac{\text{m}}{\text{MM}} = \frac{1\ 11}{12\ 01 + (4 \times 1\ 008)} = 00691\ \text{mol}$

$\text{m}(\text{O}_2) = 8366 - 81.41 = 225\ \text{g}$

$n(\text{O}_2) = \frac{2\ 25}{32\ 00} = 00702\ \text{mol}$

$\text{m}(\text{CO}_2) = 8441 - 81.41 = 300\ \text{g}$

$\text{n}(\text{CO}_2) = \frac{3\ 00}{12\ 01 + (2 \times 16\ 00)} = 00682\ \text{mol}$

$$m(SO_2) = 8562 - 81.41 = 4.21 \text{ g}$$

$$n(SO_2) = \frac{421}{3206 + (2 \times 1600)} = 00657 \text{ mol}$$

As the moles present in the same volume (can) are similar, within experimental error, molar volume of all gases is the same so the statment is confirmed.

End-of-modue reiew

1 $n(\text{Fe}) = \frac{m}{MM} = \frac{100}{5585} = 179 \text{ mol}$

$n(\text{C}) = \frac{3}{2} \times n(\text{Fe}) = \frac{3}{2} \times 1.79 = 2685 \text{ mol}$

$m(\text{C}) = n \times MM = 2685 \times 12.01 = 3225 = 323 \text{ g (3sf)}$

2 **a** $c(\text{Al(NO}_3)_3) = \frac{n}{V} = \frac{020}{0500} = 0400 \text{ mol}$

$\text{Al(NO}_3)_3 : \text{Al}^{3+}\ \text{NO}_3^{-} = 1 : 1 : 3$

$n(\text{Al}^{3+}) = 0400$

b $n(\text{NO}_3^{-}) = 3 \times 0400 = 120 \text{ mol}$

3 $PV = nRT$

$144 \times 10^3 \times 0500 = n \times 8314 \times (32 + 27315)$

$n = 02838 \text{ mol}$

$m = n \times MM = 02838 \times (2 \times 14.01) = 7952 = 80 \text{ g (2sf)}$

4 $\text{CaCO}_3\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$

$n(\text{HCl}) = c \times V = 1.12 \times 0025 = 002800 \text{ mol}$

$n(\text{CaCO}_3) = \frac{m}{MM} = \frac{2\ 5}{4008 + 1201 + (3 \times 1600)} = 02498 \text{ mol}$

If 002800 mol hydrochloric acid reacts then $\frac{1}{2} \times n(\text{HCl}) = 05 \times 002800 = 00140$ mol calcium carbonate required, thus the hydrochloric acid is he limiting reagent.

$n(\text{CaCl}_2) = \frac{1}{2} \times n(\text{HCl}) = 05 \times 002800 = 0125 \text{ mol}$

$m(\text{CaCl}_2) = n \times MM = 0125 \times (4008 + (2 \times 3545)) = 13.87 = 14 \text{ g (2sf)}$

5 **a** $c = \frac{n}{V} = \frac{120 \times 10^{-3}}{01000} = 00120 \text{ mol L}^{-}$

$c\ V = c_2 V_2$

$500 \times 00120 = c_2 \times 1000$

$c_2 = 600 \times 10^{-5} \text{ mol L}^{-}$

b pipette (see Figure 8.4b on page 179)

ii volumetric flask (se Figure 8.4d on page 179)

6 $MM = 4008\ + (4 \times 1008)\ + (2 \times 3097)\ + (8 \times 1600)\ = 193972$ g mol $^{-}$

$$\%P = \frac{2 \times 30\ 97 \times 100}{193\ 172} = 3206\%$$

Thus $\frac{5\ 0}{0\ 3206} = 15\ 6$ kg required (as 32 or 5.0 kg will be P)

7 $P\ T\ = P_2T_2$

$172 \times (20 + 27315)\ = P_2 \times (55 + 27315)$

$P_2 = 153.0 = 150$ kPa (2sf)

8 **a** $\text{RMM} = 5869\ + (8 \times 12.01) + (14 \times 1008)\ + (4 \times 14.01) + (4 \times 1600)\ = 288922$ g mol $^{-}$

b 31

c $\%N = \frac{4 \times 14\ 01 \times 100}{288\ 922} = 1940\%$

$10\ \times 01940\ = 01940$ g is nitrogen toms.

$$n = \frac{m}{MM} = \frac{0\ 1940}{14\ 01} = 001385\ = 00139 \text{ mol}$$

$N = n \times N_A = 00139\ \times 6022\ \times 10^{23} = 8339\ \times 10^{21}$ atoms

9 **a** $n = c \times V = 0200\ \times 0250\ = 00500$ mol $\text{Pb(NO}_3)_2$

$m = n \times MM = 00500\ \times (2072\ + (2 \times 14.01) + (6 \times 1600))\ = 1656\ = 166$ g (3sf)

b $n = c \times V = 0247\ \times 0025\ = 0006175$ mol

$$V = \frac{n}{c} = \frac{0\ 006175}{0\ 200} = 003088\ = 0031 \text{ L (2sf)}$$

10 $PV = nRT$

$505\ \times 00257\ = n \times 8314\ \times (272\ + 27315)$

$n = 5.198 \times 10^{-4}$ mol

$$MM = \frac{m}{n} = \frac{0\ 0772}{5\ 198 \times 10^{-4}} = 1485 \text{ g mol}^{-}$$

$MM = 3207\ + (y \times 1900)\ = 1485$

$y = 61$ Thus SF_6

11 $\text{CuCO}_3\text{(s)} + 2\text{HNO}_3\text{(aq)} \rightarrow \text{Cu(NO}_3)_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$

$$n(\text{CuCO}_3) = \frac{m}{MM} = \frac{0\ 18}{63\ 55 + 12\ 01 + (3 \times 16\ 00)} = 0001457 \text{ mol}$$

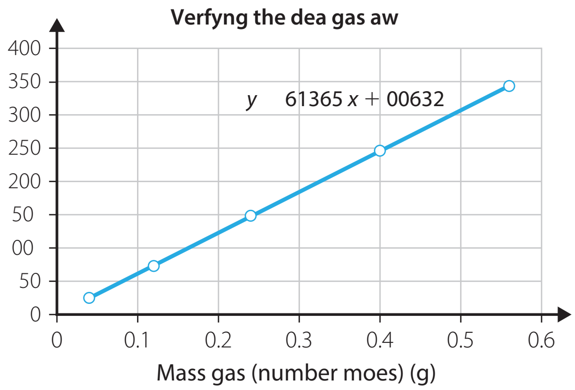
$$n(\text{HNO}_3) = \frac{m}{MM} = \frac{0\ 18}{1\ 008 + 14\ 01 + (3 \times 16\ 00)} = 0002856 \text{ mol}$$

If .002856 mol nitric acid react then $\frac{1}{2} \times 0002856\ = 0001428$ mol copper carbonate required, so nitric acid is the limiting reagent.

$n(\text{Cu(NO}_3)_2) = \frac{1}{2} \times n(\text{HNO}_3) = 05\ \times 0002856\ = 0001428$ mol

$m(\text{Cu(NO}_3)_2) = n \times MM = 0001428\ \times (6355\ + (2 \times 14.01) + (6 \times 1600))\ = 02678\ = 027$ g (2sf)

12 a Ideal gas law is $PV = nRT$, thus $P \times V$ is proportional to n at constant temperature. This should give a linear graph when PV is plotted against n In this case, mass will be used for n as they are proportional as they are all the same gas. This is verified by the graph.



b Slope of graph $= \frac{PV}{m} = \frac{PV}{n \times MM} = \frac{PV}{n} \times \frac{1}{MM} = 61365$ (from equation on graph)

As PV = nR, $\frac{PV}{n} = RT$

Thus $RT \times \frac{1}{MM} = 61365$

$R \times (22 + 27315) \times \frac{1}{4\ 003} = 61365$

$R = 8323$

This value is very close to the known value of 8.314, hence this experiment is valid.

13 a The hydrochloric acid becomes the limiting reagent after a certain amount of calcium carbonate is reacted, thus any extra calcium carbonate added does not result in any further products forming.

b 090 g

c $n(\text{CaCO}_3) = \frac{m}{MM} = \frac{0\ 90}{40\ 08 + 12\ 01 + (3 \times 16\ 00)} = 0008991$ mol

$\text{CaCO}_3\text{(s)} + 2\text{HCl(aq)} \rightarrow \text{CaCl}_2\text{(aq)} + \text{H}_2\text{O(l)} + \text{CO}_2\text{(g)}$

$n(\text{HCl}) = 2 \times n(\text{CaCO}_3) = 2 \times 0008991 = 001798$ mol

$c = \frac{n}{V} = \frac{0\ 01798}{0\ 050} = 03597 = 036\ \text{mol L}^{-}$ (2sf)

14 a $P\ V = P_2V_2$

$101 \times 68 = 14.9 \times 10^3 \times V_2$

$V_2 = 00461 = 0046\ \text{m}^3$ (2sf)

b $\frac{P}{T} = \frac{P_2}{T_2}$

$\frac{14\ 9}{20 + 273\ 15} = \frac{P_2}{38 + 273\ 15}$

$P_2 = 15.8 = 16$ MPa (2sf)

15 a $PV = nRT$

$780 \ \times 0505 \ = n \times 8314 \ \times (31.2 + 27315)$

$n = 001557$ mol

$$MM = \frac{m}{n} = \frac{1\ 34}{0\ 01557} = 8606 \ = 861\ (3sf)$$

b $MM(C_6H_{\ 4}) = (6 \times 12.01) + (14 \times 1008) \ = 86172 \text{ g mol}^{-}$

Hence the empirical formula is also the molecular formula.

STUDENT BOOK ANSWERS

Module three: Reactive chemistry

Chapter 10: Types of chemical reaction

Worked examle 1.1

$2C_6H_6(l) + 15O_2(g) \rightarrow 12CO_2(g) + 6H_2O(l)$

Check your understandng 10.1/10.2/10.3/10.4

1 Synthesis reactions involve two or more substances combining to form a new substance. A direct combination reaction involves elements combning to form compounds.

2 **a** Thermal decomposition – heating of calcium carbonate to form calcium oxide and carbon dioxide.

Electrolysis – passing a current through a liquid or solid to break a compound down – electrolysis of copper bromide to form copperand bromine gas.

Light – silver chloride when exposed to light decomposes tosilver and chlorine gas.

b $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$CuBr_2(aq) \rightarrow Cu(s) + Br_2(aq)$

$2AgCl(s) \rightarrow 2Ag(s) + Cl_2(g)$

3 Students' own responses

4 **a** zinc chloide: $Zn(s) + Cl_2(g) \rightarrow ZnCl_2(s)$

b magnesium sulfide: $Mg(s) + S(s) \rightarrow MgS(s)$

c aluminium fluoride: $2Al(s) + 3F_2(g) \rightarrow 2AlF_3(s)$

d potassiumoxide: $4\,K(s) + O_2(g) \rightarrow 2K_2O(s)$

5 **a** $S(s) + O_2(g) \rightarrow SO_2(g)$

b $2P(s) + 3Cl_2(g) \rightarrow 2PCl_3(s)$

c $O_2(g) + 2F_2(g) \rightarrow 2OF_2(g)$

d $4P(s) + 5O_2(g) \rightarrow 2P_2O_5(s)$

6 **a** $ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

b $Cu(OH)_2 \rightarrow CuO(s) + H_2O(l)$

c $2Fe(OH)_3(s) \rightarrow Fe_2O_3(s) + 3H_2O(l)$

d $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

7 $2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(l)$

8 N^{3-} $2Na(s) + 3N_2(g) \rightarrow 2NaN_3(s)$

9 Magnesium burning in oxygen to form magnesium oxide. Iron burning i oxygen to form iron oxide.

10 $4Cr(NO_3)_3(s) \rightarrow 2Cr_2O_3(s) + 12NO_2(g) + 3O_2(g)$

Worked examle 1.2

1 Could form potassium nitrate (soluble) and silver carbonate (insoluble).

$K_2CO_3(aq) + AgNO_3(aq) \rightarrow KNO_3(aq) + Ag_2CO_3(s)$

2 Could form sodium nitrate (soluble) and lead sulfate (soluble).

$Na_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow NaNO_3(aq) + PbSO_4(aq)$

3 Could form potassium chloride (soluble) and barium hydroxide (soluble), so no precipitate formed.

Check your understandng 10.5/10.6

1 $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$, neutral speciesequation.

$Ba^{2+}(aq) + 2Cl^{-}(aq) + 2Na^{+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2Na^{+}(aq) + 2Cl^{-}(aq)$, complete ionic equation.

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$, net ionic equtions.

2 **a** Sulfates – soluble except Ag^{+} Pb^{2+} Ba^{2+} Sr^{2+} Ca^{2+}

b Carbonates – insoluble except group 1 and NH_4^{+}

c Nitrates – soluble

d Hydroxides – insoluble except group 1 NH_4^{+} Ba^{2+} Sr^{2+} Ca^{2+}

e Chlorides – soluble except Ag^{+} Pb^{2+}

f Ammonium compounds – soluble

g Potassium compounds – soluble

3 copper hydroxide, iron(II) oxide, zinc carbonate, iron(III) hydroxide, lead sulfate, magnesium hydroide, silver hloride, barium carbonate

4 **a** $Cu(OH)_2$

b $MgSO_4$

c $AlCl_3$

d FeO

e $ZnCO_3$

f $Cu(NO_3)_2$

g $CaCl_2$

h K_2SO_4

$Fe(OH)_3$

j $PbSO_4$

k $(NH_4)_2SO_4$

Na_2CO_3

- **m** $Mg(OH)_2$
- **n** AgCl
- **o** $BaCO_3$

5 **a** Lead sulfate: $PbSO_4$
- **b** Zinc carbonate: $ZnCO_3$
- **c** No precipitate formed
- **d** Iron(III) hydroxide: $Fe(OH)_3$
- **e** No precipitate formed
- **f** Aluminium hydroxide: $Al(OH)_3$

6 **a** (**5a**) $Pb^{2+}(aq) + 2NO_3^-(aq) + 2NH_4^+(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s) + 2NO_3^-(aq) + 2NH_4^+(aq)$

(**5d**) $Fe^{3+}(aq) + 3Cl^-(aq) + 3Na^+(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s) + 3Cl^-(aq) + 3Na^+(aq)$

b $Pb^{2+}(aq) + SO_4^{2-}(aq) \rightarrow PbSO_4(s)$

$Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

7 **a** (**5b**) $ZnSO_4(aq) + K_2CO_3(aq) \rightarrow ZnCO_3(s) + K_2SO_4(aq)$

(**5e**) $CuSO_4(aq) + 2NaCl(aq) \rightarrow CuCl_2(aq) + Na_2SO_4(aq)$

(**5f**) $Al_2(SO_4)_3(aq) + 6KOH(aq) \rightarrow 2Al(OH)_3(s) + 3K_2SO_4(aq)$

b (**5b**) $Zn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow ZnCO_3(s)$

(**5e**) No net ionic equation as all substances stay soluble.

(**5f**) $Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$

8 **a** barium nitrate and sodium carbonate
- **ii** silver nitrate and sodium sulfate
- **iii** Iron(III) nitrate and potassium hydroxide

b $Ba^{2+}(aq) + 2NO_3^-(aq) + 2Na^+(aq) + CO_3^{2-}(aq) \rightarrow 2Na^+(aq) + 2NO_3^-(aq) + BaCO_3(s)$
- **ii** $2Ag^+(aq) + 2NO_3^-(aq) + 2Na^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s) + 2NO_3^-(aq) + 2Na^+(aq)$
- **iii** $Fe^{3+}(aq) + 3NO_3^-(aq) + 3K^+(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s) + 3NO_3^-(aq) + 3K^+(aq)$

c $Ba^{2+}(aq) + CO_3^{2-}(aq) \rightarrow BaCO_3(s)$
- **ii** $2Ag^+(aq) + SO_4^{2-}(aq) \rightarrow Ag_2SO_4(s)$
- **iii** $Fe^{3+}(aq) + 3OH^-(aq) \rightarrow Fe(OH)_3(s)$

Worked example 1.3

- **a** Li_2CO_3
- **b** $Al(NO_3)_3$

Check your understandng 10.7/10.8/10.9

1 **a** Hydrochloric acid, HCl; sulfuric acid, H_2SO_4; carbonic acid, H_2CO_3

b Sodium hydroxide, NaOH; iron(III) oxide, Fe_2O_3 ammonia, NH_3

c Sodium chloride, NaCl; potassium carbonate, K_2CO_3; iron(II) sulfate, $FeSO_4$

2 **a** Chemical reactions in which acids react with bass to form ionic salts.

b $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

$H_2SO_4(aq) + 2NH_3(aq) \rightarrow (NH_4)_2SO_4(aq)$

3 **a** Sodium hydroxid, NaOH; potassium oxide, K_2O

b Magnesium oxid, MgO; copper hydroxide, $Cu(OH)_2$

4 Bubble the gas through clear limewater (calciu hydroxide solution). If carbon dioxide is present, then the solution turns milky as insoluble calcium carbonate forms.

$Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

5 **a** $HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$

ii $2HNO_3(aq) + ZnO(s) \rightarrow Zn(NO_3)_2(aq) + H_2O(l)$

iii $HNO_3(aq) + NH_3(g) \rightarrow NH_4NO_3(aq)$

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

ii $H_2SO_4(aq) + Cu(OH)_2(s) \rightarrow CuSO_4(aq) + 2H_2O(l)$

6 **a** $2H^+(aq) + K_2O(aq) \rightarrow 2K^+(aq) + H_2O(l)$

ii $3H^+(aq) + Fe(OH)_3(s) \rightarrow Fe^{3+}(aq) + 3H_2O(l)$

iii $H^+(aq) + NH_3(aq) \rightarrow NH_4^+(aq)$

b $OH^-(aq) + HNO_3(aq) \rightarrow NO_3^-(aq) + H_2O(l)$

ii $2OH^-(aq) + H_2CO_3(aq) \rightarrow CO_3^{2-}(aq) + 2H_2O(l)$

iii $3OH^-(aq) + H_3PO_4(aq) \rightarrow PO_4^{3-}(aq) + 3H_2O(l)$

c nitrate ion

ii carbonate ion

iii phosphate ion

7 **a** LiCl

b $ZnSO_4$

c $Cu(NO_3)_2$

d $FeCl_3$

e $(NH_4)_2CO_3$

f $Al_2(SO_4)_3$

8 **a** calcium fluoride, sodium chloride, potassium bromide, silver iodide, potassium sulfate, sodium sulfite, lead nitrate, sodium nitrite, calcium carbonate, sodium phosphate, silver acetate, potassium cyanide, zinc sulfide

b silver iodide

ii magnesium chloride

iii ammonium carbonate

v iron(III) chloride

v aluminium nitrate

v calcium acetate

9 **a** (**7c**) copper nitrate – add copper hydroxide and nitric acid together

(**7f**) aluminium sulfate – add aluminium oxide and sulfuric acid

(**8b iii**) ammonium carbonate – add ammonium hydroxide and carbonic acid

b (**7c**) copper nitrate

$\mathrm{Cu(OH)_2(aq) + 2HNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2H_2O(l)}$

ii $\mathrm{Cu^{2+}(aq) + 2OH^-(aq) + 2H^+(aq) + NO_3^-(aq) \rightarrow Cu^{2+}(aq) + 2NO_3^-(aq) + 2H_2O(l)}$

iii $\mathrm{OH^-(aq) + H^+(aq) \rightarrow H_2O(l)}$

(**7f**) aluminium sulfate

$\mathrm{Al_2O_3(s) + H_2SO_4 \rightarrow Al_2SO_4(aq) + H_2O(l)}$

ii $\mathrm{Al_2O_3(s) + 2H^+(aq) + SO_4^{2-}(aq) \rightarrow 2Al^{3+}(aq) + SO_4^{2-}(aq) + H_2O(l)}$

iii $\mathrm{Al_2O_3(s) + 2H^+(aq) \rightarrow 2Al^{3+}(aq) + H_2O(l)}$

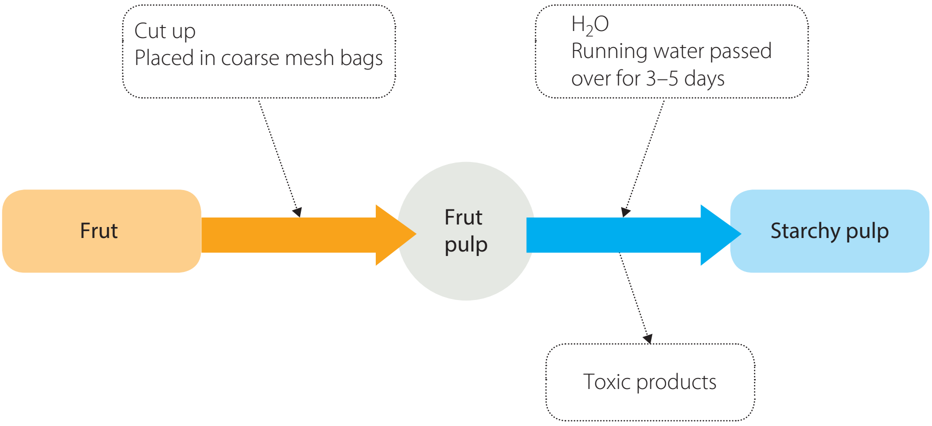
(**8b iii**) ammonium carbonate

$\mathrm{2NH_4OH(aq) + H_2CO_3(aq) \rightarrow (NH_4)_2CO_3(aq) + 2H_2O(l)}$

ii $\mathrm{2NH_4^+(aq) + 2OH^-(aq) + 2H^+(aq) + CO_3^{2-}(aq) \rightarrow 2NH_4^+(aq) + CO_3^{2-}(aq) + 2H_2O(l)}$

iii $\mathrm{OH^-(aq) + H^+(aq) \rightarrow H_2O(l)}$

10 Student flowcharts will vary depending o process sele ted. An example is shown.



Chapter review questions

1 Synthesis reactions – magnesium burning in oxygen to form magnesium oxide.

Decomposition reactions – heating of calcium carbonate to for calcium oxide and carbon dioxide.

Combustion reactions – burning of octane in oxygen (asa fuel) to produceenergy, carbon dioxide and water.

Precipitation reactions – formation of lead iodide solid by addition of lead nitrate and potassium iodide.

Acid–base (neutralisation) reactions – addition of sodium hydroxide and hydrochloric acid to form the neutral salt sodium chloride, and water.

Acid–carbonate reactions – sulfuric acid added to copper carbonate forming copper sulfate, carbon dioxide and water.

2 **a** magnesium–oxygen, magnesium–chlorine magnesium–sulfur,zinc–oxygen, zinc–chlorine, zinc–sulfur, aluminiumoxygen, aluminium–chlorine aluminium–sulfur, sulfur–chlorine, sulfur–oxygen, chlorine–oxygen

b magnesium–oxygen: $2Mg(s) + O_2(g) \rightarrow 2MgO(s)$

aluminium–chlorine: $2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

sulfur–oxygen: $S(s) + O_2(g) \rightarrow SO_2(g)$

3 $O_2(g) + 2H_2(g) \rightarrow 2H_2O(l)$

$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

$H_2(g) + S(s) \rightarrow H_2S(g)$

4 **a** all hydroxides, most carbonates, many nitrates, no sulfates, no chlorides

b $Cu(OH)_2(s) \rightarrow CuO(s) + H_2O(l)$

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

$4LiNO_3(s) \rightarrow 2Li_2O(s) + 4NO_2(g) + O_2(g)$

5 $2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

6 All group 1, potassium/sodium/ammonium/nitrate compounds are soluble.

Chlorides (except silver and lead)are soluble.

Sulfates (except silver, lead, barium, strontium and calcium) are soluble.

Carbonates (except group 1 and ammonium/potassium/sodium) are insoluble.

Hydroxides and oxides (except group 1 and ammonium/potassium/sodium/barium/strontium and calcium) are insoluble.

7 $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2NaCl(aq)$, neutral species equation

$Ba^{2+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s)$, net ionic equation

$ZnSO_4(aq) + K_2CO_3(aq) \rightarrow ZnCO_3(s) + K_2SO_4(aq)$, neutral species equation

$Zn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow ZnCO_3(s)$, net ionic equation

$Al_2(SO_4)_3(aq) + 6KOH(aq) \rightarrow 2Al(OH)_3(s) + 3K_2SO_4(aq)$, neutral species equation

$Al^{3+}(aq) + 3OH^-(aq) \rightarrow Al(OH)_3(s)$, net ionic equation

8 Hydrochloric acid, HCl; sulfuric acid, H_2SO_4; carbonic acid, H_2CO_3; nitric acid, HNO_3

9 $2HCl(aq) + CaCO_3(s) \rightarrow CaCl_2(aq) + CO_2(g) + H_2O(l)$

10 a Bubble the gas through clear limewater (calciu hydroxide solution). If carbon dioxide is present, then the solution turns milky as insoluble alcium carbonate forms.

b $Ca(OH)_2(aq) + CO_2(g) \rightarrow CaCO_3(s) + H_2O(l)$

11 a The cut-up fruit is heated to decompose toxins, then leaching (soaking in running water for a number of days) is performed to remove an remaining toxins.

b The heating is a decomposition reaction, converting the toxins into less harmful products by breaking up the toxin compounds.

12 calcium carbonate, zinc hydroxide, aluminium hydroxide, lead carbonate, zinc nitrate

13 a carbon dioxide and water (for both)

b $2C_6H_{4}(l) + 13O_2(g) \rightarrow 6CO_2(g) + 14H_2O(l)$

$C_2H_6O(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

14 a $2C_3H_8(g) + 7O_2(g) \rightarrow 6CO(g) + 8H_2O(l)$

ii $2C_3H_8(g) + 9O_2(g) \rightarrow 4CO_2(g) + 2CO(g) + 8H_2O(l)$

b $C_6H_{4}(l) + 8O_2(g) \rightarrow 3CO_2(g) + 3CO(g) + 7H_2O$

c $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

15 a & b

no precipitate formed

ii precipitate = silver chloride

$KCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + KNO_3(aq)$

iii no precipitate formed

v precipitate = barium sulfate

$H_2SO_4(aq) + BaCl_2(aq) \rightarrow BaSO_4(s) + 2HCl(aq)$

v precipitate = magnesium sulfate

$MgSO_4(aq) + Na_2CO_3(aq) \rightarrow MgCO_3(s) + Na_2SO_4(aq)$

v precipitate = iron(II) hydroxide

$FeSO_4(aq) + 2KOH(aq) \rightarrow K_2SO_4(aq) + Fe(OH)_2(s)$

v precipitate = lead chloride

$2NH_4Cl(aq) + Pb(NO_3)_3(aq) \rightarrow 2NH_4NO_3(aq) + PbCl_2(s)$

16 a $ZnO(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

b $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

c $2Al(OH)_3(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 6H_2O(l)$

d $2HCl(aq) + MgO(s) \rightarrow MgCl_2(aq) + H_2O(l)$

e $Fe(OH)_3(s) + 3HCl(aq) \rightarrow FeCl_3(aq) + 3H_2O(l)$

f $2HNO_3(aq) + K_2O(s) \rightarrow 2KNO_3(aq) + H_2O(l)$

17 a zinc chloride, water and carbon dioxide

$ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$

b potassium sulfate, water and carbon dioxide

$K_2CO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + CO_2(g) + H_2O(l)$

c silver nitrate, water and carbon dioxide

$2HNO_3(aq) + Ag_2CO_3(s) \rightarrow 2AgNO_3(aq) + CO_2(g) + H_2O(l)$

d lead nitrate, water and carbon dioxide

$PbCO_3(s) + 2HNO_3(aq) \rightarrow Pb(NO_3)_2(aq) + CO_2(g) + H_2O(l)$

e barium chloride, water and carbon dioxide

$2HCl(aq) + BaCO_3(s) \rightarrow BaCl_2(aq) + CO_2(g) + H_2O(l)$

18 Acids have a sour taste, conduct electricity in solution, hange blue litmus dye to red. They all produce hydrogen ions in slution.

Bases contain hydroxide ions, oxide ions or produce hydoxide ions in solution. Bases that are soluble in water are calledalkalis. Alkalis have a soapy feel, have a bitter taste, conduct electricity in solution and change red litms to blue.

19 a & b

add nitric acid

$ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

ii add nitric acid

$Ag_2O(s) + HNO_3(aq) \rightarrow AgNO_3(aq) + H_2O(l)$

iii add sulfuric acid

$K_2CO_3(aq) + H_2SO_4(aq) \rightarrow K_2SO_4(aq) + H_2O(l)$

v add hydrochloric acid

$2HCl(aq) + Ba(OH)_2(aq) \rightarrow BaCl_2(aq) + H_2O(l)$

20 a Add 5 mL each of solutions of aluminium nitrate and sodium hydroxide in a large test tube. Solid precipitate of aluminium hydroxide should form. Filter out the solid and dry overnight to recover the solid aluminium hyroxide.

b Add 5 mL each of solutions of zinc chlorideand lead nitrate. Solid lead chloride will form that can be filtered out, leaving behind a filtrate of zinc nitrate. Evaporate to dryness, the zinc nitrate crystals should remain behind after the wate has evaporated. Dry overnight to recover the solid zinc nitrate.

21 Net ionic equations show the reacting species, which is useful, but they do not show what chemical is used to do his. For example, in the fowing reaction:

$Zn^{2+}(aq) + CO_3^{2-}(aq) \rightarrow ZnCO_3(s)$

a solution of zinc is used, but this equation does not indicate what solution is used so this reaction would be impossible to replicate from the netionic equation alone. Thus, these equations are useful for demonstrating the reacting species but not useful for allowing relication of the chemical reaction.

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 11: Reactions of metals

Check your understandng 11.1/11.2/11.3

1 **a** lithim, soium, potassium

b hydrogen gas

c $Li(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$

$K(s) + H_2O(l) \rightarrow KOH(aq) + H_2(g)$

2 **a** ionic salt and hydrogen gas

b Do react with dilute acid – zinc, magnesium, calcium, aluminium

Do not react with dilute acid – silver, gold

c $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

3 Burn – magnsium, aluminium

React slowly – tin, copper

Do not react at all – gold, platinum

4 They react with oxygen at roo temperature. Storage under oil keeps the oxygen away from the metal so it can be stored safely.

5 **a** Zinc metal into a solution of copper ions will displace the opper ions out of solution. Copper metal into a solution of silver ions will displace the silver ions out of solution.

b $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

$Cu(s) + 2Ag^{+}(aq) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

6 Predicting reactions of various metals with water and hydrogen. Predicting metal activity for displacement reactios to occur.

7 **a** $2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

b $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

8 **a** $Mg(s) + HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

ii $Mg(s) + 2H^{+}(aq) + 2Cl^{-}(aq) \rightarrow Mg^{2+}(aq) + 2Cl^{-}(aq) + H_2(g)$

iii $Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

b $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

ii $2Al(s) + 6H^+(aq) + 3SO_4^{2-}(aq) \rightarrow 2Al^{3+}(aq) + 3SO_4^{2-}(aq) + 3H_2(g)$

iii $2Al(s) + 6H^+(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

9 **a** $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

b $Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)$

c $2Al(s) + 3Pb(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Pb(s)$

10 a (any metals above lead in activity) zinc, magnesium, iron

ii (any metals above iron in activity) calcium, magnesium, aluminium

b $Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

ii $Ca(s) + Fe^{2+}(aq) \rightarrow Ca^{2+}(aq) + Fe(s)$

11 a & b

reaction occurs

$2Al(s) + 3Pb(NO_3)_2(aq) \rightarrow 2Al(NO_3)_3(aq) + 3Pb(s)$

$2Al(s) + 3Pb^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Pb(s)$

ii no reaction occurs

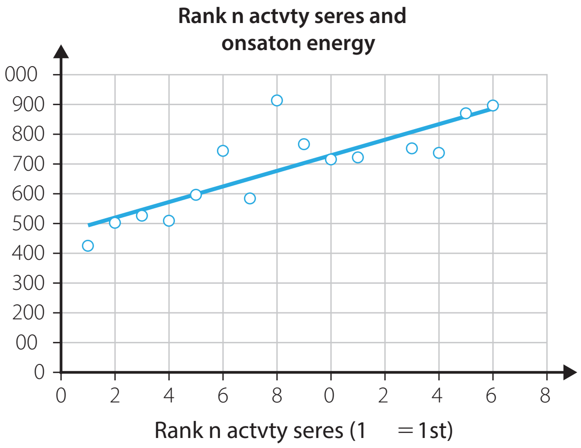
iii no reaction occurs

v reaction occurs

$Mg(s) + CuSO_4(aq) \rightarrow Cu(s) + MgSO_4(aq)$

$Mg(s) + Cu^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cu(s)$

12 There is a general correlation between rank in activity series and ionisation energy, but certain elements (ranks of 6 and 8 in particular) do not folow the trend well. A linear trend line can be drawn to show the general relationship that as rank in activity series increases the ionisation energy of the metal increases.



Check your understandng 11.4/11.5

1 The most reactive metals are on the upper left of the periodictable (excluding hydrogen).

2 Ionisation energy is the ease with which an electron an be removed from the atom. The lower the ionisation energy, the easier it is to remove an eectron and hence react. Thus, the metals with lower ionisation energy are mre reactive.

3 **a** Oxidation – metal atom losing electrons to form positive metal ons, an element gaining oxygen.

Reduction – non-metal atom gaining electrons to form negative non-metal ions, an element losing oxygn.

b Electrons must be transferred from one substance to another, thus oxidation (loss of electrons) must be accompanied by reduction (gain f electrons).

4 **a/d** Before potassium (at the start) – it is further down group 1, so is more reactive than potassium.

b/d Between calcium and barium – it is one period earlier than barium and one later than calcium, and reactivity increases up the groupso fits between them.

c/d Around the area of Zn > Fe > Cu – part of the transition metals, but less pattern is obvious with transition metals than with main group elements.

5 **a/b** $Li \rightarrow Li^+ + e^-$ (lithium is oxidised)

$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$ (water is reduced)

ii $Al \rightarrow Al^{3+} + 3e^-$ (aluminium is oxidised)

$2H^+ + 2e^- \rightarrow H_2$ (hydrogen ions are reduced)

iii $Pb \rightarrow Pb^{2+} + 2e^-$ (lead is oxidised)

$Cu^{2+} + 2e^- \rightarrow Cu$ (copper ions are reduced)

6 **a–c** Yes – redox reaction

Iron = oxidised, oxygen = reduced

$Fe \rightarrow Fe^{3+} + 3e^-$, $O_2 + 2e^- \rightarrow 2O^{2-}$

ii Yes – redox reaction

Zinc = oxidised, chlorine = reduced

$Zn \rightarrow Zn^{2+} + 2e^-$, $Cl_2 + 2e^- \rightarrow 2Cl^-$

iii No – not a redox reaction

The calcium ion starts and finishes as a calcium ion, so has no gain or loss of electrons. Likewise, the oxygen and hydrogen do not gain or ose electrons.

v Yes – redox reaction

Magnesium = oxidised, lead ions = reduced

$Mg \rightarrow Mg^{2+} + 2e^-$, $Pb^{2+} + 2e^- \rightarrow Pb$

v Yes – redox reaction

Silver = oxidised, hydrogen = reduced

$Ag \rightarrow Ag^+ + e^-$, $H_2 + 2e^- \rightarrow 2H^+$

v No – not a redox reaction

Zinc and chlorine both begin and end as ions, so do notlose or gain electrons.

7 **a** Equation 1: $Pb^{2+} + 2e^{-} \rightarrow Pb$ (reduction)

Equation 2: $O_2 + 2e^{-} \rightarrow 2O^{2-}$ (reduction)

b Equation 1: lead reduced, carbon oxidised

Equation 2: sulfur oxidised, oxygen reduced

Worked examle 11.1

1 $H = +1$ and $O = -2$ Let P be z

$3(+1) + z + 4(-2) = 0$

$z = +5$

2 $O = -2$ Let $S = z$

$2(z) + 3(-2) = -2$

$z = +2$

3 $Na = +1$ and $O = -2$ Let $Br = z$

$+1 + z + 3(-2) = 0$

$z = +5$

Check your understandng 11.6/11.7

1 **a** $CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$, oxidant = copper ins.

b $Zn + 2HCl \rightarrow ZnCl_2 + H_2$ reductant = zin.

2 1 A substance in its elemental state = 0, regardless of how many atoms in the molecule/structure.

2 Oxidation number of a monatomic ion is th charge on the ion.

3 Fluorine has oxidation number −1

4 Hydrogen has oxidation number +1 (unless rule 2 contradics this).

5 Oxygen has oxidation number −2 (unless contradicted by rule 2, 3, or 4)

6 Chlorine, bromine and iodine have oxidation number −1 (unless contradicted b rule 5).

7 The sum of oxidation numbers in a compound or ion must equal the charge on the molecule (0) or ion (overall charge on the ion)

3 Oxidation numbers are calculated for each atom in the equation according to the rules given. If the oxidation number of the species increases from left to right side of the equation, it has oxidised. If the oxidation number of the species decreases from left to right side of the equation, it has reduced.

4 **a** copper ins: $Cu^{2+} + 2e^{-} \rightarrow Cu$

ii chlorin gas: $Cl_2 + 2e^{-} \rightarrow 2Cl^{-}$

b iron(II) ions: $Fe^{2+} \rightarrow Fe^{3+} + e^{-}$

ii iron meal: $Fe \rightarrow Fe^{2+} + 2e^{-}$

5 **a** +2

ii +3

iii +2

v +4

v +4

v +5

b −1

ii −2

iii −2

v −3

6 F = −1, thus B = +3

Cl = −1, thus Ti = +4

3(+1) + P + 4(−2) = 0, thus P = +5

+1 + Cl + 3(−1) = −1, thus Cl = +4

Al + 6(−1) = −3, thus Al = +3

+1 + N + 2(−2) = 0, thus N = +3

S + 3(−2) = −2, thus S = +4

+1 + −2 + I = 0, thus I = +1

Pb + 2(−2) = 0, thus Pb = +4

7 **i** Mn from +3 to +4 = oxidation

ii Br from −1 to 0 = oxidation

iii Fe from +2 to +3 = oxidation

v Cr from +3 to +3 = neither

v S from 0 to −2 = reduction

v O from −2 to −2 = neither

v Ni from +3 to +2 = reduction

v C from −4 to +2 = oxidation

8 a, d and e

9 **a/b** (**8a**) phosphorus is oxidised (+3 to +5), chlorine is the oxidising agent

(**8e**) sulfur is oxidised (+2 to +4), oxygen is the oxidising agent

10 **i** $H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$

$2I^- \rightarrow I_2 + 2e^-$

overall: $H_2O_2 + 2H^+ + 2I^- \rightarrow 2H_2O + I_2$

ii $I_2 + 2e^- \rightarrow 2I^-$

$H_2SO_3 + H_2O \rightarrow SO_4^{2-} + 4H^+ + 4e^-$

overall: $H_2SO_3 + H_2O + 2I_2 \rightarrow SO_4^{2-} + 4H^+ + 4I^-$

iii $Br_2 + 2e^- \rightarrow 2Br^-$

$HNO_2 + H_2O \rightarrow NO_3^- + 3H^+ + 2e^-$

overall: $HNO_2 + H_2O + Br_2 \rightarrow NO_3^- + 3H^+ + 2Br^-$

v $Fe^{2+} \rightarrow Fe^{3+} + e^-$

$Cl_2 + 2e^- \rightarrow 2Cl^-$

overall: $2Fe2^+ + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-$

b hydrogen peroxide, H_2O_2

iii bromne, Br_2

c **ii** sulfurous acid, H_2SO_3

v iron(ii) ions, Fe^{2+}

Chapter reiew quesions

1 **a** A metal hydroxide and hydrogen gas.

b lithium, sodium, potassium

2 **a** calcium, potssium, barium

b zinc, iron, aluminium

c god, siver, platinum

3 **a** zin, aluminum, calcium

b hydrogen gas

4 $Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$

$Cu(s) + 2Ag^+(aq) \rightarrow 2Ag(s) + Cu^{2+}(aq)$

$Zn(s) + Pb^{2+}(aq) \rightarrow Zn^{2+}(aq) + Pb(s)$

5 From most to leastreactive: group 1 metals, group 2 metals, transition metals

6 Reactivity increases goig down a group.

7 **a** As ionisation energy decreases, metal activity increases.

b As electronegativity increases, meta reactivity decreases.

8 Oxidation occurs when the oxidation number ofan element increases. Reduction occurs when the oxidation number of an elemet decreases.

9 The oxidation number of an element in a molecule or ion is the charge the atom of that element would carry if the molecule or ion were cmpletely ionic.

10 **a** $Zn + 2HCl \rightarrow ZnCl_2 + H_2$

$CuSO_4 + Zn \rightarrow ZnSO_4 + Cu$

b $\mathrm{Zn \rightarrow Zn^{2+} + 2e^{-},\ 2H^{+} + 2e^{-} \rightarrow H_2}$

$\mathrm{Cu^{2+} + 2e^{-} \rightarrow Cu,\ Zn \rightarrow Zn^{2+} + 2e^{-}}$

11 a $\mathrm{2Zn(s) + O_2(g) \rightarrow 2ZnO(s)}$

b $\mathrm{Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)}$

c $\mathrm{2Al(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2(g)}$

d $\mathrm{Fe(s) + CuSO_4(aq) \rightarrow Cu(s) + FeSO_4(aq)}$

12 a (**c**) $\mathrm{2Al(s) + 6H^{+}(aq) + 6Cl^{-}(aq) \rightarrow 2Al^{3+}(aq) + 6Cl^{-}(aq) + 3H_2(g)}$

$\mathrm{2Al(s) + 6H^{+}(aq) \rightarrow 2Al^{3+}(aq) + 3H_2(g)}$

(**d**) $\mathrm{Fe(s) + Cu^{2+}(aq) + SO_4^{2-}(aq) \rightarrow Cu(s) + Fe^{2+}(aq) + SO_4^{2-}(aq)}$

$\mathrm{Fe(s) + Cu^{2+}(aq) \rightarrow Cu(s) + Fe^{2+}(aq)}$

b (**c**) $\mathrm{Al \rightarrow Al^{3+} + 3e^{-}}$

$\mathrm{2H^{+} + 2e^{-} \rightarrow H_2}$

(**d**) $\mathrm{Fe \rightarrow Fe^{2+} + 2e^{-}}$

$\mathrm{Cu^{2+} + 2e^{-} \rightarrow Cu}$

13 a zinc, iron, magnesium (any metal above copper in the activity series)

b tin, lead, iron

14 a left to right

ii right to left

iii left to right

v left to right

b $\mathrm{Sn + 2Ag^{+} \rightarrow Sn^{2+} + 2Ag}$

ii $\mathrm{2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu}$

iii $\mathrm{Mg + Fe^{2+} \rightarrow Mg^{2+} + Fe}$

v $\mathrm{Pb + 2Ag^{+} \rightarrow Pb^{2+} + 2Ag}$

c $\mathrm{Sn \rightarrow Sn^{2+} + 2e^{-},\ 2Ag^{+} + 2e^{-} \rightarrow 2Ag}$

iii $\mathrm{Mg \rightarrow Mg^{2+} + 2e^{-},\ Fe^{2+} + 2e^{-} \rightarrow Fe}$

15 a (any metal above iron in the activity series) magnesium, calcium, zinc

ii (any metal between zinc and copper in the activity series) tin, lead, iron

b $\mathrm{Mg(s) + Fe^{2+}(aq) \rightarrow Mg^{2+}(aq) + Fe(s)}$

$\mathrm{Ca(s) + Fe^{2+}(aq) \rightarrow Ca^{2+}(aq) + Fe(s)}$

$\mathrm{Zn(s) + Fe^{2+}(aq) \rightarrow Zn^{2+}(aq) + Fe(s)}$

ii $\mathrm{Sn(s) + Cu^{2+}(aq) + NO_3^{-}(aq) \rightarrow Sn^{2+}(aq) + Cu(s) + NO_3^{-}(aq)}$

$\mathrm{Pb(s) + Cu^{2+}(aq) + NO_3^{-}(aq) \rightarrow Pb^{2+}(aq) + Cu(s) + NO_3^{-}(aq)}$

$\mathrm{Fe(s) + Cu^{2+}(aq) + NO_3^{-}(aq) \rightarrow Fe^{2+}(aq) + Cu(s) + NO_3^{-}(aq)}$

16 a iodine (reduced from 0 to −1)

ii sulfuric acid (the sulfur in it is reduced from +6 to +4)

b Fe^{2+} (oxidised from +2 to +3)

ii HCl (the chlorine in it is oxidised from −1 to 0)

c See above

17 a–d $CuS(s) + O_2(g) \rightarrow CuO(s) + SO_2(g)$

Yes, it is a redox raction.

sulfur – oxidised from −2 to +4

oxygen – reduced from 0 to −2

ii $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2$

Not a redox reaction.

iii $Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$

Yes, it is a redox raction.

calcium – oxidised from 0 to +2

hydrogen – reduced from +1 to 0

v $Na_2O(s) + HCl \rightarrow NaCl(aq) + H_2O(l)$

Not a redox reaction.

v $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

Yes, it is a redox raction.

nitrogen – oxidised from +2 to +4

oxygen – reduced from 0 to −2

18 Al_2Cl_6 $2(Al) + 6(-1) = 0$, $Al = +3$

H_3BO_3: $3(-1) + B + 3(-2) = 0$, $B = +3$

$HClO_4$: $+1 + Cl + 4(-2) = 0$, $Cl = +7$

H_2SiO_4: $2(+1) + Si + 4(-2) = 0$, $Si = +6$

$Na_2S_2O_3$: $2(+1) + 2(S) + 3(-2) = 0$, $S = +2$

$HAsO_2$: $+1 + As + 2(-2) = 0$, $As = +3$

H_3AsO_4: $3(+1) + As + 4(-2) = 0$, $As = +5$

19 a $CrSO_4$ sulfate ion is −2, thus $Cr = +2$

Cr_2O_3 $2(Cr) + 3(-2) = 0$, $Cr = +3$

CrO_2^- $Cr + -2 = -1$, $Cr = +3$

CrO_4^{2-} $Cr + 4(-2) = -2$, $Cr = +6$

b $MnSO_4$ sulfate ion is −2, thus Mn = +2

$Mn(OH)_3$ hydroxide ion is −1, thus Mn = +3

MnO_2 Mn + 2(−2) = 0, Mn = +4

K_2MnO_4: 2(+1) + Mn + 4(−2) = 0, Mn +6

$KMnO_4$: +1 + Mn + 4(−2) = 0, Mn = +7

c NH_3 N + 3(+1) = 0, N = −3

N_2H_4 2(N) + 4(+1) = 0, N = −2

N_2O 2(N) + −2 = 0, N = +1

N: N + −2 = 0, N = +2

N_2O_3 2(N) + 3(−2) = 0, N = +3

NO_2 N + 2(−2) = 0, N = +4

N_2O_4 2(N) + 4(−2) = 0, N = +4

N_2O_5 2(N) + 5(−2) = 0, N = +5

d HB: +1 + Br = 0, Br = −1

HOBr: +1 + −2 + Br = 0, Br = +1

Br_2 Br = 0

$HBrO_3$: +1 + Br + 3(−2) = 0, Br = +5

20 Oxidation numbers should increase for one element (oxidation or loss of electrons), and decrease for another element (reduction or gan of electrons).

21 a/b **i** carbon is oxidised (0 to +4)

sulfur is reduced (+6 to +4)

ii not a redox reaction

iii oxygen is oxidised (−1 in hydrogen peroxide to 0)

manganese is reduced (+7 to +2)

v not a redox reaction

v copper is oxidised (0 to +2)

nitrogen is reduced (+5 to +2)

v not a redox reaction

v carbon is oxidised (+3 to +4)

bromine is reduced (0 to −1)

c sulfur (as part of sulfuric acid)

v not a redox reaction

v nitrogen (as part of nitric acid)

d **ii** not a redox reaction

iii oxygen (as part of hydrogen peroxide)

v carbon (as part of $H_2C_2O_4$)

22 a/b This will depend on experiments condcted in class. Check with your teacher, but it should be based on Investgation 11.3 or a similar experiment.

23 a $Fe(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2(g)$

$FeS(s) + H_2SO_4(aq) \rightarrow FeSO_4(aq) + H_2S(g)$

b Equaton 1: iron is oxidised, hydrogen is reduced

Equaton 2: not a redox reaction

24 a Due to relative reactivity, the lithium should react with the sulfuric acid in preference to the water, but as it is dilute, some lithium would als react with the water.

b $Li(s) + H_2O(l) \rightarrow LiOH(aq) + H_2(g)$

$Li(s) + H_2SO_4(aq) \rightarrow Li_2SO_4(aq) + H_2(g)$

Both reactions produced hydrogen gas so there is no difference between the two reactions from this point o view. Any hydroxide ions produced by reaction with water would be neutralised by the hydrogen ions from the sulfuric acid,producing water. Thus, it does not matter whether the lithium reacts with the watr or acid. The final components of the solution/products will be the same overall.

25 $CuS(s) + O_2(g) \rightarrow CuO(s) + SO_2(g)$

Sulfur: oxidised from −2 to +4

Oxygn: reduced from 0 to −2

As seen in the above example, use of oxidation numbers shows very effectively whether a reaction is a redox reaction and which specific elements are oxidised (increase in oxidation number) or reduced (decrease in oxidation number).

26 $Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow Cr^{3+} + 7H_2O$

$Fe^{2+} \rightarrow Fe^{3+} + e^-$

Overall: $Cr_2O_7^{2-} + 14H^+ + 6Fe^{2+} \rightarrow Cr^{3+} + 7H_2O + 6Fe^{3+}$

As seen in the above example, use of the correct procedure for balancing half equations and then balancing electrons to add the two half equations together provides an easy and useful method of writing full redox reactions.

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 12: Galvanic cells and standard electrode potentials

Worked examle 1.1

a At the silverwire: $Ag^+(aq) + e^- \rightarrow Ag(s)$

At the platinm wire: $Sn^{2+} \rightarrow Sn^{4+} + 2e^-$

Overall reation: $2\,Ag^+(aq) + Sn^{2+} \rightarrow Sn^{4+} + Ag(s)$

b Electrons flow from the platinum wireto the silver wire.

Within the ell: negative ions flow out of the Ag^+ solution and into the Sn^{2+}/Sn^{4+} solution. Positive ions flow in the opposite drection.

c Anode = tin electrode. Cathode = silver electrode.

Worked examle 1.2

a At the aluminium electrode: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

At the platinm wire: $2H^+(aq) + 2e^- \rightarrow H_2(g)$

Overal: $6H^+(aq) + 2Al(s) \rightarrow 2Al^{3+}(aq) + 3H_2(g)$

b Electrons flow from the aluminium strip to the platinum wire.

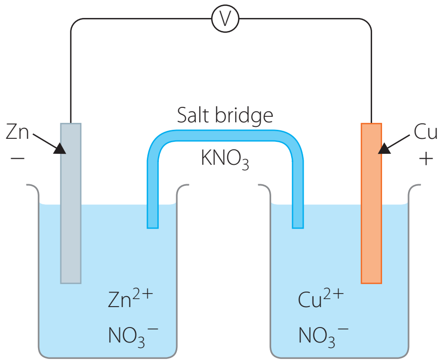
Within the ell: negative ions flow from the platinum wie to the aluminium strip. Positive ions flow in the opposite diection.

c Anode = aluminium electrode. Cathode = hydrogen eletrode.

Check your understandng 12.1/12.2/12.3

1 **a** In the zinc/copper galvanic cell in the diagram shown, the zinc is the anode and consists of a zinc electrode in a zinc ion solution. The copper is the cathode and consists of a copper electrode in a copper io solution.

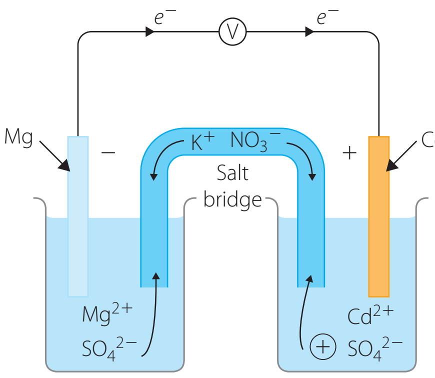
b The zinc is more reactive than the copper, so will oxidise ($Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$), producing electrons. The electrons travel through the external wire to the copper electrode where reduction occurs ($Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$). The flow of electrons through the external wre is electricity.

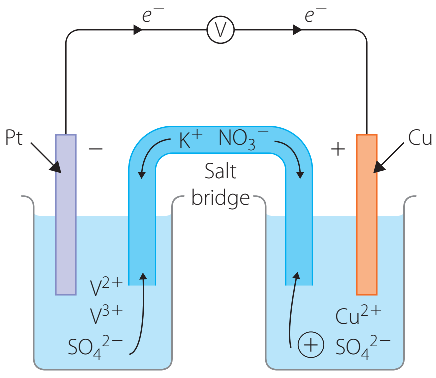


2 **a** Redox reactions generat electricity.

b Other reactions do not have elctron transfer. It is the transfer and capture of electrons in redox reactions that allow for electriity to be produced.

3 **a** Cel A: Cel B:





b Cel A: $Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-$, $Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

Overall: $Mg(s) + Cd^{2+}(aq) \rightarrow Mg^{2+}(aq) + Cd(s)$

Cel B: $V^{2+}(aq) \rightarrow V^{3+}(aq) + e^-$, $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

Overall: $V^{2+}(aq) + Cu^{2+}(aq) \rightarrow V^{3+}(aq) + Cu(s)$

c See diagrams

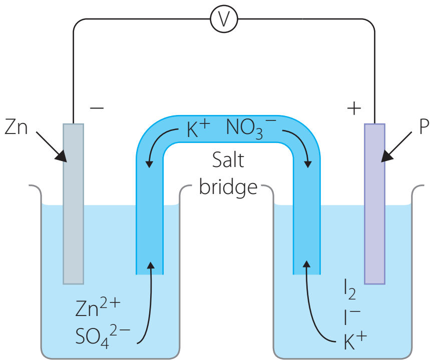
d Cel A: Anode = magnesium electrode. Cathode = cadmium elecrode.

Cel B: Anode = vanadium electrode. Cathode = copper electrode.

4 **a** **ii, iii, v**

b Reactions i and iv have no electron transfer and are not redo reactions, thus cannot be used.

5



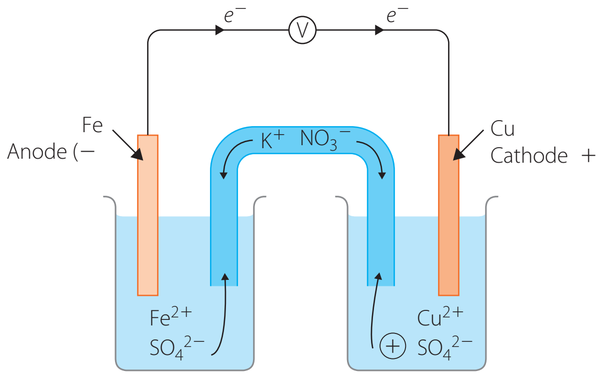
a Oxidation is occurring at the zinc electrode ($Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$) and electrons are beng produced. Reduction is occurring at the iodine/platinum electrode ($I_2(aq) + 2e^- \rightarrow 2I^-(aq)$)

b Overall: $Zn(s) + I_2(aq) \rightarrow Zn^{2+}(aq) + 2I^-(aq)$

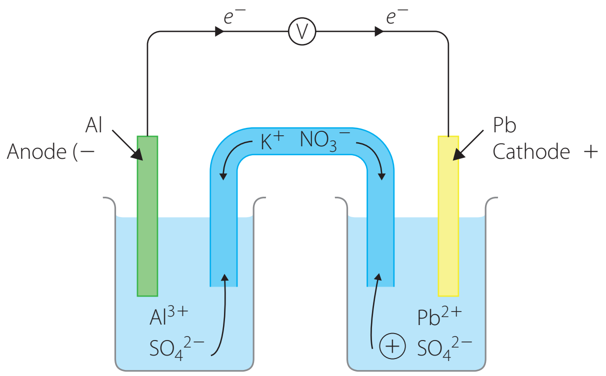
c See diagram

d Anode = zinc electrode.

6 **a**



ii



b Iron is more reactive so will oxidise andbe the negative anode. Copper will be the positive cathode.

ii Aluminium is more reactive so will oxidise andbe the negative anode.Lead will be the positive cathode.

c $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^-, Cu^{2+} + 2e^- \rightarrow Cu(s)$

Overall: $Cu^{2+} + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$

ii $Al(s) \rightarrow Al^{3+}(aq) + 3e^-, Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

Overall: $Pb^{2+}(aq) + Al(s) \rightarrow Al^{3+}(aq) + Pb(s)$

d **e** See diagrams

7 **a** Yes, as chloride ions are provided for the cathode reaction.

b No, as chloride ions are not provided for he cathode reaction.

c No, as chloride ions are not provided for he cathode reaction.

Worked examle 1.3

1 **a** $Sn^{4+} + 2e^- \rightarrow Sn^{2+}, E^\ominus \ = +015\text{ V}$

$Fe \rightarrow Fe^{2+} + 2e^-, E^\ominus{}_2 = -(-045)\text{V}$

$E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = +015 \ + 045 \ = +060\text{ V}$

b $Fe^{2+} + 2e^- \rightarrow Fe, E^\ominus \ = -045\text{ V}$

$Pb \rightarrow Pb^{2+} + 2e^-, E^\ominus{}_2 = -(-013)\text{V}$

$E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = -045 \ + 013 \ = -032\text{ V}$

Worked examle 1.4

1 $Br_2 + 2e^- \rightarrow 2Br^-, E^\ominus \ = +109\text{ V}$

$Mg \rightarrow Mg^{2+} + 2e^-, E^\ominus{}_2 = -(-237)\text{V}$

$E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = +109 \ + 237 \ = +346$ V, with the bromine/platinum electrode being positive with respect to the magnesium electrode.

2 $Ag^+ + e^- \rightarrow Ag, E^\ominus \ = +080\text{ V}$

$Cd \rightarrow Cd^{2+} + 2e^-, E^\ominus{}_2 = -(-040)\text{V}$

$E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = +080 \ + 040 \ = 120$ V, with the silver electrode being positive with respect to the cadmium elecrode.

Worked examle 1.5

$E^{\ominus}_{\text{total}} = E^{\ominus}_{\text{Cu}} - E^{\ominus}_{\text{T}}$

$194 = +034 - E^{\ominus}_{\text{T}}$

$E^{\ominus}_{\text{T}} = -160 \text{ V}$

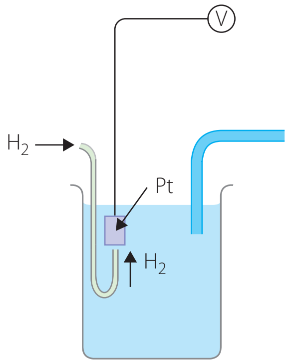
Check your understandng 12.4/12.5/12.6

1 **a** The potential of an electrode in its standard state (concentration of 1.000 mol L^{-} for solutions and 1000 kPa for gases) relative to the standahydrogen electrode.

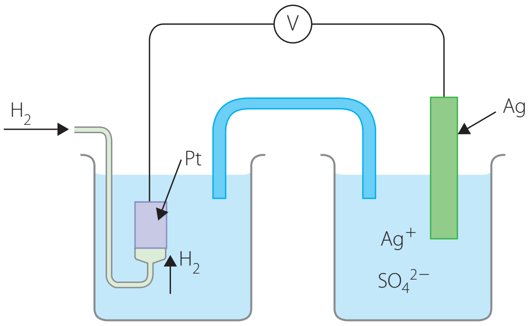
b The half cell (electrode) is connected to the stdard hydrogen electrode. The difference in voltage is the standard electrode potential of the electrod (a hydrogen is arbitrarily 0.00V).

2 A standard electrode potential is foundat a concentration of 1.000 mol L^{-} for solution and 100.0 kPa for gaes. Electrode potential is when standard onditions are not followed.

3 **a**



b



4 $E^{\ominus}_{\text{total}} = E^{\ominus}_{\text{oidtion}} + E^{\ominus}_{\text{reduction}}$

5 A negative value indicates a non-spontaneous reaction, or a reaction that will occur in the reverse direction.

6 **a** $E^{\ominus}_{\text{total}} = E^{\ominus}_{\text{Au}} - E^{\ominus}_{\text{H}}$

$199 = E^{\ominus}_{\text{Au}} - 000$

$E^{\ominus}_{\text{Au}} = 199 \text{ V}$

ii $Au^{3+} + 3e^{-} \rightarrow Au$

$H_2 \rightarrow 2H^{+} + 2e^{-}$

b $E^{\ominus}_{\text{total}} = E^{\ominus}_{\text{H}} - E^{\ominus}_{\text{Cd}}$

$040 = 000 - E^{\ominus}_{\text{Cd}}$

$E^{\ominus}_{\text{Cd}} = -040 \text{ V}$

ii $2H^{+} + 2e^{-} \rightarrow H_2$

$Cd \rightarrow Cd^{2+} + 2e^{-}$

c $E^{\ominus}_{\text{total}} = E^{\ominus}_{\text{C}} - E^{\ominus}_{\text{H}}$

$136 = E^{\ominus}_{\text{C}} - 000$

$E^{\ominus}_{\text{C}} = 1.36 \text{ V}$

ii $Cl_2 + 2e^- \rightarrow 2Cl^-$

$H_2 \rightarrow 2H^+ + 2e^-$

7 **a** magnesium in magnesim sulfate, zincin zinc sulfate iron in iron sulfate, copper in copper sulfate solution

b platinum with chlorine gas (Cl_2) and chloride ion (Cl^-) solution, platinum with hydrogen gas (H_2) and hydrogen ion (H^+) solution

c Assuming both are oxidaton reactions:

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

$2H^+(aq) \rightarrow H_2(g) + 2e^-$

8 **a** $E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = +034 \ + -(-026) \ = +060$ V

ii $E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = -(-040) \ + 000 \ = +040$ V

iii $E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = -166 \ + -(+109) \ = -257$ V

v $E^\ominus{}_{total} = E^\ominus \ + E^\ominus{}_2 = -(+077) \ + 054 \ = -023$ V

b **iii** $2Al^{3+}(aq) + 6Br^-(aq) \rightarrow 2Al(s) + 3Br_2(aq)$

v $Fe^{2+}(aq) + 2I_2(aq) \rightarrow Fe^{3+}(aq) + 4I^-(aq)$

c left to right

ii left to right

iii right to left

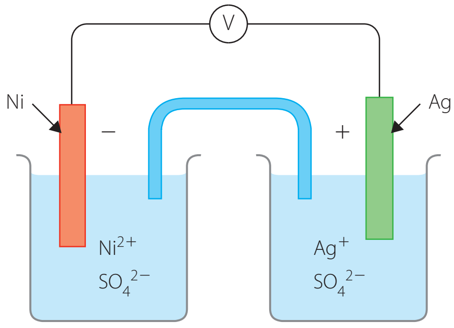
v right to left

9 **a** $E^\ominus{}_{total} = E^\ominus{}_N + E^\ominus{}_{Ag} = -(-026) \ + 080 \ = +060$ V, where silver is the positive electrode.

ii $E^\ominus{}_{total} = E^\ominus{}_{Mg} + E^\ominus{}_{Br} = -(-237) \ + 1.09 = +346$ V, where bromine is the positive electrode.

iii $E^\ominus{}_{total} = E^\ominus{}_{Fe} + E^\ominus{}_C \ = -(077) \ + 1.36 = +059$ V, where the chlorine is the postive electrode.

b



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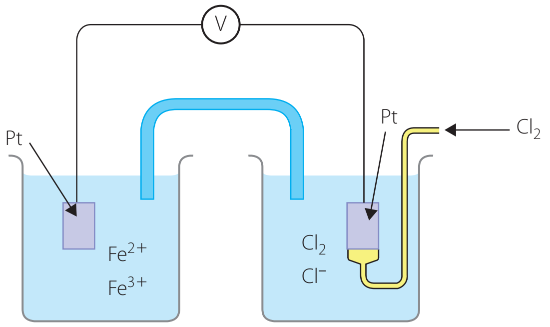

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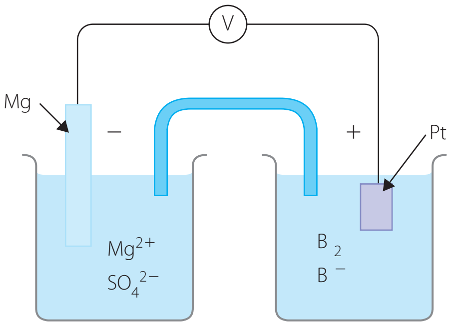


c $Ni(s) \rightarrow Ni^{2+}(aq) + 2e^-$

$Ag^+(aq) + e^- \rightarrow Ag(s)$

ii $Mg(s) \rightarrow Mg^{2+} + 2e^-$

$Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$

iii $Cl_2(g) + 2e^- \rightarrow 2Cl^-(aq)$

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-$

10 a $E^\ominus_{total} = E^\ominus_{Fe} + E^\ominus_{Pb} = -(-045) + -013 = +032$ V

Goes in the forward direction as written.

b $E^\ominus_{total} = E^\ominus + E^\ominus_{C} = +054 + -(+136) = -082$ V

Goes in the reverse direction.

c $E^\ominus_{total} = E^\ominus_{N} + E^\ominus_{Sn} = -026 + -(015) = -041$ V

Goes in the reverse direction.

d $E^\ominus_{total} = E^\ominus_{A} + E^\ominus_{Cd} = -(-166) + (-040) = +126$ V

Goes in the forward direction as written.

Worked example 1.6

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s), E^\ominus = +034$

$Hg(l) \rightarrow Hg^{2+}(aq) + 2e^-, E^\ominus = -(+085)$

$E^\ominus_{total} = +034 + -085 = -051$ V, thus as the voltage is negative, the reaction is non-spontaneous.

Check your understanding 12.7

1 Between zinc (–076) and aluminium (–166) The value for titanium is –160 V so will fit between these two.

2 a $Hg^{2+}(aq) + 2e^- \rightarrow Hg(l), E^\ominus = +085$

$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-, E^\ominus = -(-013)$

$E^\ominus_{total} = +085 + 0.13 = +098$ V, thus as the voltage is positive, the reaction will occur.

b $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s), E^\ominus = -014$

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-, E^\ominus = -(+034)$

$E^\ominus_{total} = -014 + -034 = -048$ V, thus as the voltage is negative, the reaction will not occur.

c $Ag^+(aq) + e^- \rightarrow Ag(s), E^\ominus = +080$

$Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-, E^\ominus = -(-040)$

$E^\ominus_{total} = +080 + 040 = +120$ V, thus as the voltage is positive, the reaction will occur.

d $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s), E^\ominus = (-076)$

$Pb(s) \rightarrow Pb^{2+}(aq) + 2e^-, E^\ominus = -(-013)$

$E^\ominus_{total} = -076 + 0.13 = -063$ V, thus as the voltage is negative, the reaction will not occur.

3 **a** chlorine **b** zinc **c** no **d** yes

4 **ab** $E^{\ominus}_{\text{total}} = +136 + -(+093) = +043$ V, positive value = spontaneous reaction.

ii $E^{\ominus}_{\text{total}} = -045 + -(015) = -060$ V, negative value = non–spontaneous reaction.

iii $E^{\ominus}_{\text{total}} = +054 + -(014) = +040$ V, positive value = spontaneous reaction.

v $E^{\ominus}_{\text{total}} = -(+085) + 000 = -085$ V, negative value = non–spontaneous reaction.

v $E^{\ominus}_{\text{total}} = +117 + -(-076) = +193$ V, positive value = spontaneous reaction.

5 **ab** Tin(II) chloride will reduce both (lower in the electrode potential list so is a stronger reductant than both).

ii HNO_2 will reduce Br_2 but not I_2 (lower than bromine so a stronger reductant, but higher than iodine so a weaker reductant).

iii Chromium(III) sulfate will reduce neither (higher than both bromine and iodine so is a weaker reductant).

Chapter review questions

1 Redox reactions that involve transfer of electrons.

2 Electrode can mean the piece of metal used in a galvanic cell, or can refer to the entire half cell including the metal and solution.

3 When a U-shaped tube is used, so the same electrolyte solution can be used for both metal electrodes.

4 **a** E = standard voltage of a redox reaction. ε = standard electrode potential of an electrode. E can refer to the voltage of the reaction in either direction. ε can only be used for reduction reactions.

b The first equation already takes into account that the oxidation reaction has been reversed to find E_{oid} The second equation is used when the values for reduction reactions are used for both half equations; the negative sign accounts for the reversal of the oxidation reaction.

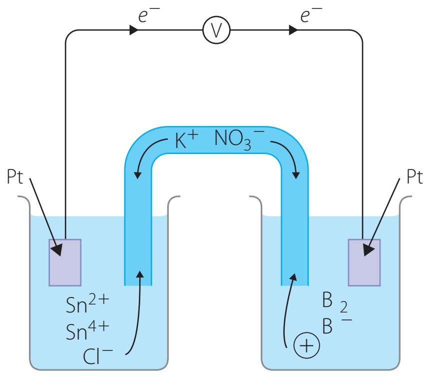
5 The reaction is non-spontaneous and will not occur on its own. It will occur spontaneously in the reverse direction.

6 Metals higher in the activity series have lower standard electrode potential values.

7 Stronger oxidants have higher standard electrode potential values and are found higher in the list of electrode potentials.

8 The reaction higher in the electrode potential list will be the reduction reaction and will occur as written. The reaction lower in the electrode potential list will be the oxidation reaction and will occur in the reverse manner as written.

9 **a**



b $Br_2(aq) + 2e^- \rightarrow 2Br^-(aq)$

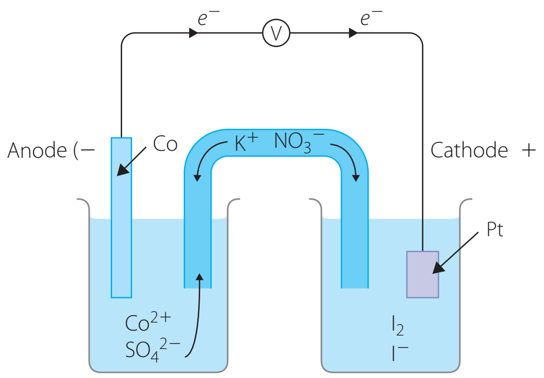
$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$

Overall: $Br_2(aq) + Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2Br^-(aq)$

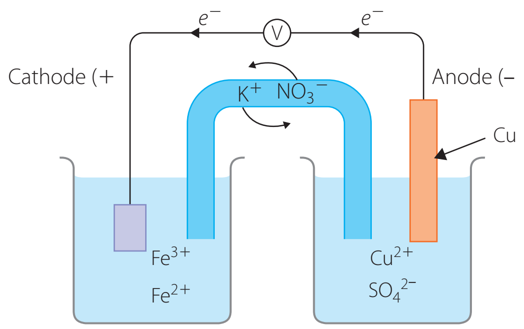
c See diagram

d Anode = tin electrode. Cathode = bromine eletrode.

10 a



b



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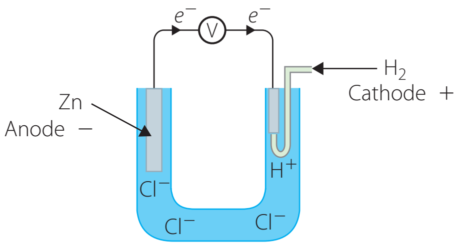


12 a between silver and platinum

b between aluminium and zinc

13 ab $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s), E^\ominus = (-013)$

$Mg(s) \rightarrow Mg^{2+}(aq) + 2e^-, E^\ominus = -(-237)$

$E^\ominus{}_{total} = -013 + 237 = 224$ V, value is positive so reactionis spontaneous.

ii $Ni^{2+}(aq) + 2e^- \rightarrow Ni(s), E^\ominus = (-026)$

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-, E^\ominus = -(+077)$

$E^\ominus{}_{total} = -026 - 077 = -103$ V, value is negative so nonspontaneous.

iii $H_2S(g) \rightarrow S(s) + 2H^+(aq) + 2e^-, E^\ominus = -(+014)$

$I_2(aq) + 2e^- \rightarrow 2I^-(aq), E^\ominus = +054$

$E^\ominus{}_{total} = +054 - 0.14 = 040$ V, value is positive so reactionis spontaneous.

v $Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq), E^\ominus = +015$

$Fe^{2+}(aq) \rightarrow Fe^{3+}(aq) + e^-, E^\ominus = -(+077)$

$E^\ominus{}_{total} = +015 - 077 = -062$ V, value is negative so nonspontaneous.

v $Ag_2O(s) + H_2O(l) + 2e^- \rightarrow 2Ag + 2OH^-, E^\ominus = +034$

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-, E^\ominus = -(-014)$

$E^\ominus_{total} = +034 \ + 0.14 = 048$ V, value is positive so reactionis spontaneous.

14 bromine, nitric acid, iron(III) sulfate, potassium permanganate, tin(IV) chloride

15 Iron metal will cause iodine to reduce, while the iron metal oxidises.

$Fe \rightarrow Fe^{2+} + 2e^-$

$I_2 + 2e^- \rightarrow 2I^-$

Overall: $Fe + I_2 \rightarrow Fe^{2+} + 2I$

ii Silver will cause dichromate ions to reduce, while the silver oxidises.

$Ag \rightarrow Ag^+ + e^-$

$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$

Overall: $Cr_2O_7^{2-} + 14H^+ + 6Ag \rightarrow 2Cr^{3+} + 7H_2O + 6Ag^+$

iii Aluminium metal will cause tin(IV) ions to reduce, whil the aluminium oxidises.

$Al \rightarrow Al^{3+} + 3e^-$

$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

Overall: $3Sn^{4+} + 2Al \rightarrow 3Sn^{2+} + 2Al^3$

16 a (anything above bromide and iodide) chlorine (Cl_2)

b (anything above iodide, but below bromide) silver ions (Ag^+)

c (anything below bromine) nickel metal (Ni)

d (anything below Fe^{3+}, but above Fe^{2+}) lead metal (Pb)

e (anything below Fe^{2+}) zinc metal (Zn)

f (anything above H_2S) copper ions (Cu^{2+})

17 a aluminium, iron, hydrogen, sulfur dioxide solution, iron(II) sulfate, manganese(II) sulfate

b Stronger reductants are at the bottom of the electrode potential table, reducing strength decreases as you go up thetable.

18 a (anything below hydrogen peroxide) chloride ions, copper metal, hydroxide ions

b (anything below tin(IV) ions) zinc, nickel, magnesium

c (anything between nitric acid and hydrogen ions) silver, copper

d (anything below titanium – really aluminium or lower) aluminium, magnesium

19 The table of standard electrode potentials is useful for identifying which reactants are more likely to reduce or oxidise, predicting the strength of a substance as an oxidant or reactant, predicting how two reactants are likely to react if added together, and for calculating the voltage generated by a redox reaction. Overall, this makes the table of standardpotentials extremely useful.

20 a Zinc is more reactive than iron so will oxidise in preference to iron. Aluminium is also more reactive than iron so will protect the iron in the same way as zinc.

b Copper is less reactive than iron, so iron will oxidise in preference to the copper, so it will not be protected.

21 a Cell voltage $= 1.6 + -(-036) = +196$ V, the lead dioxide electrode s positive.

b $6 \times 1.96 = 1176$ V. So yes, this is consistent witha 12 V battery.

c $PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) + Pb(s) \rightarrow 2PbSO_4(aq) + 2H_2O(l)$

d The hydrogen ions will be used up, so the pH of the solution will increase.

STUDENT BOOK ANSWERS

Module two: Introduction to quantitative chemistry

Chapter 13: Rates of reaction

Check your understanding 13.1/13.2

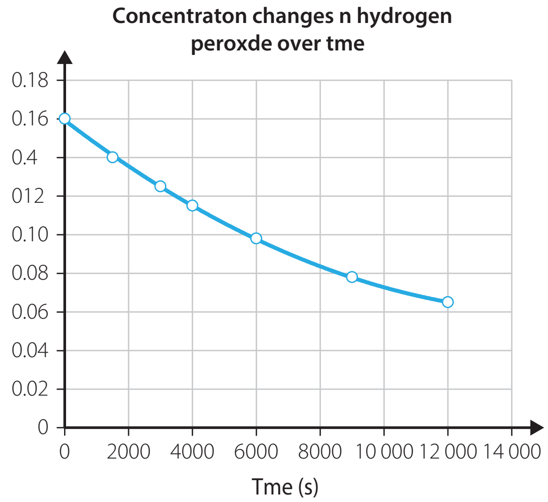
1 **a** A homogenous reaction has the reactants present as a homogenous mixture of gases or solutions. A heterogeneous reaction occurs at the interface between two phases; for example, between a solid or a gas, or between a solid and a solution

b Homogenous reaction: the reaction between nitric oxide and oxygen to form nitrogen dioxide – all gass.

Heterogeneous reaction: the reaction of zinc metal solid with a solution of hydrochloric acid to form zinc chloride and hydrogen gas.

2 Stirring ensures particles come into contact more often as the particles are dispersed, so more surface area is available for reaction.

3 **a**



b $$\text{average rate} = \frac{\Delta c}{\Delta t} = \frac{016 - 0125}{3000} = 1.7 \times 10^{-5}\ \text{mol s}^{-}\ \ (2\text{sf})$$

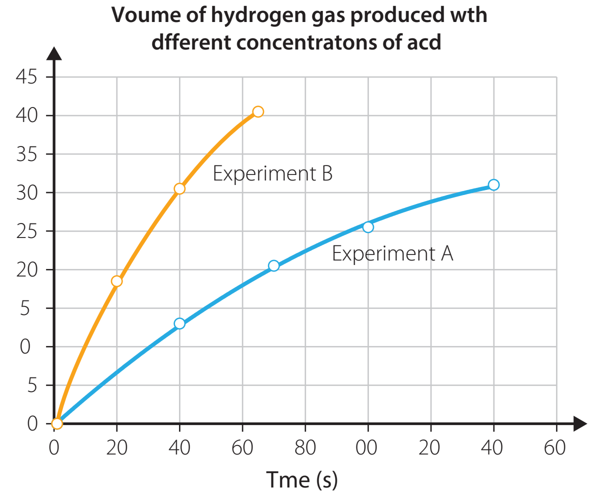
$$\text{average rate} = \frac{\Delta c}{\Delta t} = \frac{0098 - 0078}{3000} = 6.7 \times 10^{-6}\ \text{mol s}^{-}\ \ (2\text{sf})$$

$$\text{average rate} = \frac{\Delta c}{\Delta t} = \frac{0078 - 0065}{3000} = 4.3 \times 10^{-6}\ \text{mol s}^{-}\ \ (2\text{sf})$$

c Rate of reaction decreases as the reaction proceeds.

d As the reactants are used up they react more slowly, thus the rate of reaction decreases.

4 **a**



b Experiment B has a greater rate of reaction as it has the same change in concentration as experiment A, but in a shorter period of time, so the rate of reaction is faster.

c An increase in the concentration of hydrochloric acid gives a faster rate of reaction with magnesium.

d Experiment A: $\text{rate} = \dfrac{\Delta c}{\Delta t} = \dfrac{13-0}{40} = 033\ \text{mL s}^{-}$ (2sf)

Experiment B: $\text{rate} = \dfrac{\Delta c}{\Delta t} = \dfrac{18\ 5-0}{20} = 093\ \text{mL s}^{-}$ (2sf)

5 **a** Experiment A forms the same volume of nitrogen as experiment B in a much shorter period of time, thus experiment A has the faster initial rate of reaction.

b Reaction A will have the higher reactant concentration as there is more reactant available, thus a faster rate of reaction occurs.

6 a, d, e

7 **a** $(1 \times 1) \times 6 = 6\ \text{cm}^2$

b $(0.1 \times 0.1) \times 6 = 006\ \text{cm}^2$

ii $(10 \times 10 \times 10) = 1000$

iii $006 \times 1000 = 60\ \text{cm}^3$

v 10 1 smaller : larger

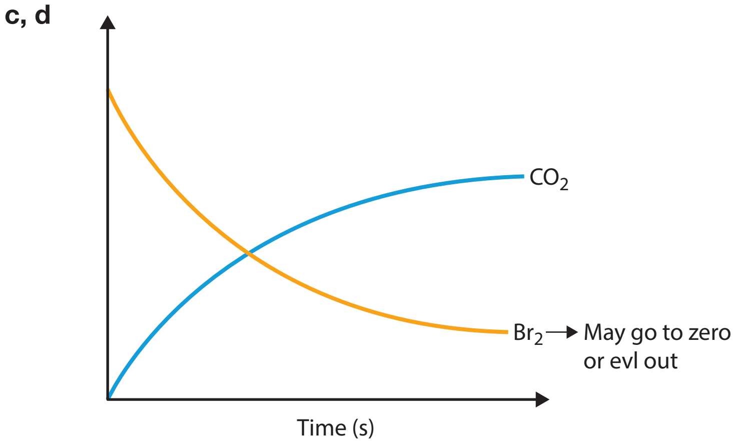
v Reaction occurs when particles are in contact. If more surface area, then more of the reactants are in contact, so the rate of reaction increases.

8 **a** $\text{Br}_2\ \text{rate} = \dfrac{\Delta c}{\Delta t} = \dfrac{0\ 0050 - 0\ 0010}{156} = 2.6 \times 10^{-5}\ \text{mol L}^{-}\ \text{s}^{-}$ (2sf)

$\text{Br}_2\ \text{rate} = \dfrac{\Delta c}{\Delta t} = \dfrac{0\ 0100 - 0\ 0010}{72} = 1.3 \times 10^{-4}\ \text{mol L}^{-}\ \text{s}^{-}$ (2sf)

$\text{Br}_2\ \text{rate} = \dfrac{\Delta c}{\Delta t} = \dfrac{0\ 0050 - 0\ 0010}{330} = 1.2 \times 10^{-5}\ \text{mol L}^{-}\ \text{s}^{-}$ (2sf)

b Both substances are involved – as concentration Br_2 increases but HCOOH the same, the reaction rate increases. As concentration HCOOH increases, but Br_2 the same, the reaction rate increases.



Check your understandng 13.3/13.4/13.5

1 See Figure13.6 page 308.

2 For a chemical reaction to occur, particles must collide (come into contact), have a minimum amount of kinetic energy (activation energy) and have the correct orientation. If the concentration of reactants is increased, then more collisions between reactants per unit volume per unit time, so the rate of reaction increases.

3 As temperature increases, all particles gain kinetic energy, increasing the proportion of particles that are above the activation energy for the reaction. Thus, more particles will have successful collisions per unit volume per unit time and reaction rate icreases markedly.

4 **a** Homogenous catalysts reduce the activation energy of a reaction by providing an alternative chemical pathway that requires less energy to occur.

b Heterogeneous catalysts provide a surface that reactant particles adsorb (stick) to, allowing for chemical bonds to break and reform.

5 Rhodium – catalyses the reaction between carbon monoxide and nitric oxide to form carbon dioxide and nitrogen gas in a car's catalytic converter.

Nitrogen dioxide – catalyses the reaction between sulfur dioxide and oxygen to form sulfur trioxide.

6 **a** Room temperature is aound 300 K. At this temperature, reaction A has the higher rate at room temperaure.

b Reaction B has the higher activation energy. The more rapid a rate of reaction increase, seen in the steeper slope of B, the higher the activation eergy of the reaction.

c See **a** and **b**

7 **a** One or both of the reactant particles adsorb to the surface of the vanadium pentoxide, weakening or breaking the bonds etween atoms. Other reactant particles collide with the adsorbed molecules, allowing the reaction to occur due to the weakened or broken bonds in the adsorbed molecules.

b Reactant particles adsorb to the surface of the iron solid, causing the bonds between the nitrogen atoms and hydrogen atoms to weaken or break. Neighbouring atoms can react, or there are collisions with other non-adsorbed reactant partiles to form ammonia.

8 There are various ways to design this experiment, but it would involve weighing the manganese dioxide prior to its addition to the hydrogen peroxide. At the end of the reaction, filter and dry the manganese dioxide, then reweigh. The mass should be the same before addition to the hydogen peroxide and upon recovery.

Chapter reiew quesions

1 The slope of the graph is $\frac{\text{concentration}}{\text{time}}$, which is the same as the formula for working out rate of reaction. Thus, the slope of the graph is representative of the rate of reaction. The steeper the slope, the more change in concentration in a shorter period of time, thus the rate of reaction is increased.

2 See Figure13.5 page 307.

3 Activation energy is the minimum amount of energy reactants must have to undergo a successful collision and formproducts.

4 Collision theory states that for a chemical reaction to occur, particles must collide (come into contact), have a minimum amount of kinetic energy (activation energy) and have the correct orientation. If the concentration of reactants is increased, then there are more collisions between reactants per unit volume per unit time, so the rate of reaction increases.

5 A homogenous catalyst decreases the activation energy of a reaction by providing an alternative chemical pathway that requires les energy to occur. Thus, a greater proportion of the reactants have the minimum kinetic energy required to have a successful collision and form products.

6 Heterogeneous catalysts provide a surface that reactant particles adsorb (stick) to, allowing for chemical bonds to break and reform. As particles adsorb to the surface, the bonds between atoms weaken or break, making it easier for them to react with other particles and form products.

7 An enzyme is a biologicalcatalyst. Enzymes are proteins that catalyse specific reactions in living organisms. Lipase is an enzyme found in the stomach that catalyses the decomposition of fats to glycerol and fatty cids.

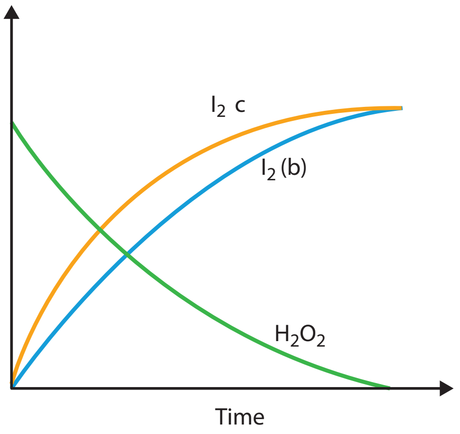
8 **a** average rate $= \frac{\Delta c}{\Delta t} = \frac{0\ 200 - 0\ 186}{20} = 7.0 \times 10^{-4}$ mol L^{-} min^{-} (2sf)

b average rate $= \frac{\Delta c}{\Delta t} = \frac{0\ 161 - 0\ 150}{20} = 5.5 \times 10^{-4}$ mol L^{-} min^{-} (2sf)

c After 1 hour the reactants have been consumed, so less reactant molecules are present per unit volume. With less reactant molecules there are less collisions per unit volume per unit time, so the rate of reaction is dereased.

9 **a** Curve A = 020 mol L^{-} Curve B = 040 mol L^{-} Curve B has a steeper slope so a faster rate of reaction (greater change in concentration per unit time). A faster rate of reaction occurs with higher concentration as more particles means more collisions pr unit volume, per unit time.

b



c See graph in (**b**)

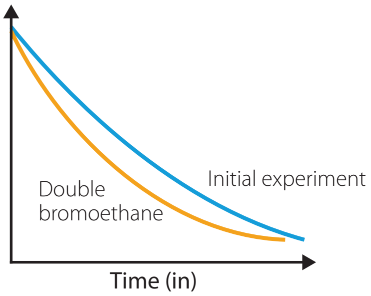
d (**b**) – As one mole iodine produced for each mole hydrogen peroxide consumed, the curves should be of roughly the same shape but inverted.

(**c**) – Higher temperature means a greater proportion of the particles above the activation energy, so more successful collisions per unit volume per unit time, thus the cuve will have a steeper slope. But, as the same number of particles were involved, the amount of final product should be the same.

10 a Experimet B. The slope of the graph is steeper than A, meaning a greater change in concentration per unit time, thus a faster rat of reaction.

b Experimet B. It has the faster rate of reaction, which means more collisions are occurring per unit volume, per unit time. The experiment at a higher temperature will have a greater proportion of the particles above the activation energy for the reaction, so will have more successful collisions and the faster reactin rate.

c The concentration of hydroxide ions will start at the same plae as the initial experiment B. With more particles of bromoethane per unit volume, there will be more collisions per unit volume per unit time, and thus a faster rate of reaction and a steeper slope.



11 A higher temperature in industry requires the use of fuel, which is expensive, and requires specialised machinery or devices that also cost money to et up and maintain. If the reaction can be carried out at room temperature, it is more cost efficientfor the industry. Thus, the room temperature reaction is the better tempeature.

12 Increases the surface area available for reaction, effectively increasing the number of collisions the particles undergo per unit volumeper unit time. Increased collision increase the rate of reaction.

13 Catalysts are not used up in a reaction, even though they may take part. The chlorine atom that is present at the start is reformed at the ed, and can be reused. Thus, the chlorine atom is a catalyst for this reaction.

14 a Reaction i is the only reaction that converts NO, whil both reactions convert CO. Assuming equal volumes of substances formed, the first reaction would need to be faster to ensure all NO converted in a suitabletime.

b Surface area of the catalysts – increasing surface area would increase reaction rate, vice versa with a decrese. Increasing or decreasing temperature inside the catalytic converter would increase or decrease (respectively) therate of reaction.

End-of-modue reiew

1 a Barium bromide and sulfuric acid would make the precipitate barium sulfate. Magnesium chloride and sodium hydroxide would make the precipitate magnesium hydroxide.

b $Mg^{2+}(aq) + 2OH^{-}(aq) \rightarrow Mg(OH)_2(s)$

c $Ba^{2+}(aq) + 2Br^{-}(aq) + 2H^{+}(aq) + SO_4^{2-}(aq) \rightarrow BaSO_4(s) + 2H^{+}(aq) + 2Br^{-}(aq)$

d For the barium sulfate precipitate equation, the spectator ions are the hydrogen ions and the bromide ions.

2 **a** A balanced equation has the same number of atoms of the same type on both sides of the equation. A balanced equation identifies the correct mole ratio f reactants and products.

b $2Al(s) + 3S(s) \rightarrow Al_2S_3(s)$

ii $Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

iii $2Pb(NO_3)_2(s) \rightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$

v $NaHSO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

v $ZnCO_3(s) + 2HNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + H_2O(l) + CO_2(g)$

3 **a** When heat is applied, all the particles in a reaction mixture gain more kinetic energy. A greater proportion of reactant particles will now have enough kinetic energy to be above the activation energy for the reaction. More successful collisions will occur per unit volume per unit time, thus increasing the reaction rate.

b Catalysts act in two ways. Homogenous catalysts lower the activation energy of a reaction by providing an alternative reaction pathway that requires less energy to get the same end result. Heterogeneous catalysts provide a surface that reactant particles adsorb to, weakening or breaking the bonds and allowing for other particles to react more easily, inceasing the rate of reaction.

4 HO l: $+1 + -2 + Cl = 0$, thus $Cl = +1$

$HClO_4$: $+1 + Cl + 4(-2) = 0$, thus $Cl = +7$

5 **a** Passing electricity through a substance to cause a chemical reaction like decomposition.

b Water will break down into hydrogen and oxygn gas by electrolysis.

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

6 **a** React with cold water to produce a solution and a gas.

$2Li(s) + 2H_2O(l) \rightarrow 2LiOH(aq) + H_2(g)$

ii React with acids to produce a solution and a gas.

$2Li(s) + 2H^{+}(aq) \rightarrow 2Li^{+}(aq) + H_2(g)$

b Does not react.

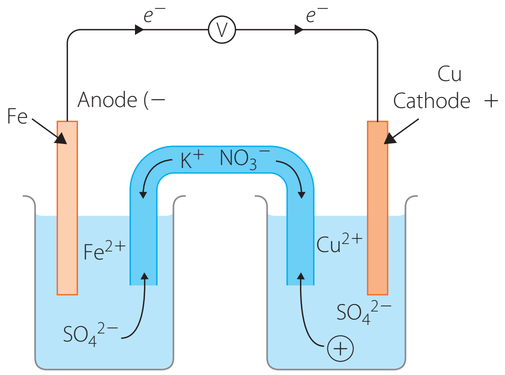
ii React with acids to produce a solution and a gas.

$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

c Does not react

ii Does not react

7 **a** Check with your teacher for the exact method you used, but it should be similar to Investigation 12.1 page 75. The following diagram shows the construction of n iron–copper galvanic cell.



b $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$, $Cu^{2+} + 2e^{-} \rightarrow Cu(s)$

Overall: $Cu^{2+} + Fe(s) \rightarrow Fe^{2+}(aq) + Cu(s)$

c See diagram

8 **a** aluminium hydroxide, copper nitrate, lithium hydroxide, zinc carbonate (hydroxides, carbonates, many nitrates)

b Soluble: barium chloride, copper nitrate, potassium sulfate, sodium chloride

Insoluble: aluminium hydroxide, lead oxide, lithium hydroxide, zinc carbonate

c zinc carbonate – would react to form carbon dioxide gas

ii aluminium hydroxide, lead oxide, lithium hydroxide

d barium chloride

ii potassium sulfate

iii none (only carbonates, oxides and hydroxides would react)

e **(a)** $2Al(OH)_3(s) \rightarrow Al_2O_3(s) + 3H_2O(l)$

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

$2LiOH(s) \rightarrow Li_2O(s) + H_2O(l)$

$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g)$

(c) $ZnCO_3(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + CO_2(g) + H_2O(l)$

(c) $Al(OH)_3(s) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O$

$PbO(s) + HCl(aq) \rightarrow PbCl_2(aq) + H_2O(l)$

$LiOH(s) + HCl(aq) \rightarrow LiCl(aq) + H_2O(l)$

(d) $BaCl_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(aq) + 2NaCl(aq)$

(d) $K_2SO_4(aq) + Ba(NO_3)_2(aq) \rightarrow BaSO_4(s) + 2KNO_3(aq)$

9 **a** Using $E = E_{ox} + E_{red}$

R/S cell: $+051 = E_R + 092$ Thus, E_R (oxidation) $= -041V$.

P/S cell: $+102 = E_P + 092$ Thus, E_P (oxidation) $= +011V$.

Q/R cell: $+090 = E_Q + 041$ Thus, E_Q (oxidation) $= +090V$.

b Values in (a) given in oxidatio potential. Values for reduction potential:

$S^{2+} + 2e^{-} \rightarrow S, E = +092V$

$R^{2+} + 2e^{-} \rightarrow R, E = +041V$

$P^{2+} + 2e^{-} \rightarrow P, E = -011V$

$Q^{2+} + 2e^{-} \rightarrow Q, E = -090V$

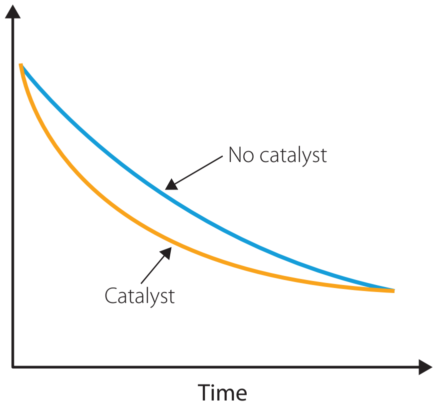
Thus, in order of decreasing reactivity, Q → P → R → S

c Values reversed to find reduction potential. A more negative reduction potential means a more reactive metal (highly reactive metals like K and Na have very high neative reduction potentials). Thus, Q is the most reactive and S is theleast reactive.

10 a $HCOOH(aq) + Br_2(g) \rightarrow CO_2(g) + 2H^{+}(aq) + 2Br^{-}(aq)$

b Experiment I – has a steeper slope showing a larger change in concentration (through intensity of colour) per unit time, thus indicating a faste rate of reaction. Higher concentrations give a faster rate as there are more collisions occur per unit volume per unit time as more particles are present per unit volume.

c A catalyst increases the reaction rate, thus the slope of the graph is steeper, but there are no extra particles, thus it ends at the same point as the initial experiment.



11 a $Zn(s) + Ag_2O(s) \rightarrow 2Ag(s) + ZnO(s)$, the silver oxide electrode is the psitive electrode.

b The potassium hydroxide acts as a salt bridge, allowing ion flow to balance the charge between the two electrodes.

c $E = +034\ + -(-126)\ = 160\text{ V}$

12 a/b Cu metal displaces Pd^{2+} out of solution so Cu is more reactive than Pd – thus a stronger reducing aget. Cu > Pd

Cd metal displaces Cu^{2+} out of solution so Cd is more reactive than Cu – thus a stronger reducing aget. Cd > Cu > Pd

V metal displaces Cd^{2+} out of solution so V is more reactive than Cd – thu a stronger reducing agent.

Overall: V > Cd > Cu > Pd

STUDENT BOOK ANSWERS

Module four: Drivers of reactions

Chapter 14: Energy changes in chemical reactions

Worked examale 1.1

$q = mc\Delta T = 253 \times 0.14 \times (55 - 18) = 131054 = 130$ J (2sf)

Check your understandng 14.1/14.2

1 $\text{J K}^{-}\ \text{g}^{-}$ (joules per kelvin per gram) or $\text{J K}^{-}\ \text{kg}^{-}$ (joules per kelvin per kilogram)

2 Temperature – physical quantity used to measure the degree of hotness or coldness of an object. Quantity of heat is the amount of heat an object contains. A red-hot pin and a (larger) red-hot horseshoe can be the same temperature, but will contain different amounts of heat.

3 Exothermic – any combustion reaction, reaction of a metal with acid.

Endothermic – decomposition of calcium carbonate.

4 $q = mc\Delta T = 1000 \times 4.18 \times (100–20) = 3.3 \times 10^{5}$ J (2sf)

5 **a** $q = mc\Delta T = 25 \times 039 \times (55–18) = 360$ J (2sf)

b $q = mc\Delta T = 25 \times 0.14 \times (55–18) = 130$ J (2sf)

c $q = mc\Delta T = 250 \times 4.18 \times 36 = 3.8 \times 10^{4}$ J (2sf)

d $q = mc\Delta T = 150 \times 045 \times (50–25) = 1.7 \times 10^{3}$ J (2sf)

e $q = mc\Delta T = 3.5 \times 244 \times 47 = 400$ J (2sf)

f $q = mc\Delta T = 1500 \times 2.17 \times 54 = 1.8 \times 10^{5}$ J (2sf)

6 $q = mc\Delta T$

$153 \times 10^{3} = 125 \times c \times (236 - 18.0)$

$c = 2.19\ \text{J K}^{-}\ \text{g}^{-}$ (3sf)

7 **a** $NH_3(aq) + HCl(aq) \rightarrow NH_4Cl(aq)$

ii exothermic

iii If the tube becomes warm, heat is being released from the chemical reaction.

b $K_2CO_3(aq) + MgSO_4(aq) \rightarrow K_2SO_4(aq) + MgCO_3(s)$

ii endothermic

iii If the tube becomes cold, heat is being absorbed by the chemical reaction.

Worked examle 1.2

1 $Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$$n = \frac{m}{MM} = \frac{1\ 1}{65\ 38} = 00168 \text{ mol}$$

$$\text{Heat released} = \frac{2\ 6}{0\ 0168} = 150 \text{ kJ mol}^{-} \ \ (2\text{sf})$$

2 $NiCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + NiCO_3(aq)$

$$n = \frac{m}{MM} = \frac{5\ 0}{58\ 69 + (2 \times 35\ 45)} = 00386 \text{ mol}$$

$$\text{Heat released} = 59 \times 00386 \ = 23 \text{ J } (2\text{sf})$$

Worked examle 1.3

a $Mg(s) + H_2SO_4(aq) \rightarrow MgSO_4(aq) + H_2(g)$

b $q = mc\Delta T = (850 + 1.05) \times 4.18 \times (23.1 - 174) \ = 203 \ \times 10^4 \text{ J}$

$$n(Mg) = \frac{m}{MM} = \frac{1\ 05}{24\ 31} = 00432 \text{ mol}$$

$$\Delta H = \frac{2\ 03 \ \times \ 10^4}{0\ 0432} = 4695 \ \times 10^5 \text{ J mol}^{-} \ = -470 \text{ kJ mol}^{-} \ \ (3\text{sf})$$

Check your understandng 14.3/14.4

1 When one mole of all the reacting substances is involved, as enthalpy change is measured in kJ mol^{-}

2 20 kJ mol^{-}

3 **a** $2Mg(s) + O_2(g) \rightarrow 2Mg(s)$

$$n(Mg) = \frac{m}{MM} = \frac{0\ 50}{24\ 31} = 002057 \text{ mol}$$

$$\Delta H = \frac{12\ 4}{0\ 02057} = -600 \text{ kJ mol}^{-} \ \ (2\text{sf})$$

b $Zn(s) + HCl(aq) \rightarrow ZnCl_2(aq) + H_2O(l)$

$$n = \frac{m}{MM} = \frac{1\ 1}{65\ 38} = 00168 \text{ mol}$$

$$\Delta H = \frac{2\ 6}{0\ 0168} = -150 \text{ kJ mol}^{-} \ \ (2\text{sf})$$

c $MgCl_2(aq) + Na_2CO_3(aq) \rightarrow 2NaCl(aq) + MgCO_3(aq)$

$$n = \frac{m}{MM} = \frac{18\ 0}{24\ 31 + (2 \times 35\ 45)} = 01891 \text{ mol}$$

$$\Delta H = \frac{9\ 2}{0\ 1891} = +487 \text{ kJ mol}^{-} \ \ (3\text{sf})$$

4 **a** i, iii, iv – all with negative ΔH valus.

b iii and iv would increase in temperature as they are exothermic; v would decrease in temperature as it is endothemic.

5 **a** $n = \frac{m}{MM} = \frac{5\ 0}{32\ 07 + (2 \times 16\ 00)} = 00780$ mol

Heat released $= 198 \times 00780 = 16$ kJ (2sf)

b $n = \frac{m}{MM} = \frac{5\ 0}{58\ 69 + (2 \times 35\ 45)} = 003858$ mol

Heat absorbed $= 59 \times 003858 = 23$ kJ (2sf)

c $n = c \times V = 020 \times 0025 = 00050$ mol

Heat released $= 106 \times 00050 = 053$ kJ (2sf)

6 **a** $\frac{198}{2} = 99, \Delta H = -99$ kJ mol^{-}

b $\Delta H = +198$ kJ mol^{-} (reverse the equation, reverse the sign of ΔH)

7 **a** exothermic

b $\Delta H =$ negative

ii **a** endothermic

b $\Delta H =$ positive

iii **a** endothermic

b $\Delta H =$ positive

v **a** exothermic

b $\Delta H =$ negative

8 **a** $NaOH(aq) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$

$q = mc\Delta T = 75 \times 4.18 \times (241 - 232) = 28215$ J

$n = c \times V = 020 \times 0025 = 00050$ mol

$\Delta H = \frac{282\ 15}{0\ 0050} = 564 \times 10^{4}$ J mol^{-} $= -56$ kJ mol^{-} (2sf)

b $Pb(NO_3)(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$

$q = mc\Delta T = 80 \times 4.18 \times (222 - 19.6) = 86944$ J

$n = c \times V = 020 \times 0050 = 00100$ mol

$\Delta H = \frac{869\ 44}{0\ 0100} = 8694 \times 10^{4}$ J mol^{-} $= -87$ kJ mol^{-} (2sf)

9 **a** $q = mc\Delta T = 02553 \times 42 \times 10^{3} \times (220 - 18.6) = 3646$ J $= 3646$ kJ

$n = \frac{m}{MM} = \frac{5\ 3}{40\ 08 + (2 \times 35\ 45)} = 00478$ mol

$\Delta H = \frac{3\ 646}{0\ 0478} = -76$ kJ mol^{-} (2sf)

b $q = mc\Delta T = 01056 \times 42 \times 10^{3} \times 23 = 1020\ \text{J} = 1020\ \text{kJ}$

$$n = \frac{m}{MM} = \frac{5\ 6}{(2\times 22\ 99) + (2\times 30\ 97) + (8\times 16\ 00) + (10\times 1\ 008)} = 002276\ \text{mol}$$

$$\Delta H = \frac{1\ 020}{0\ 02276} = +45\ \text{kJ mol}^{-}\ \ (2\text{sf})$$

c $q = mc\Delta T = 02083 \times 4200 \times (202 - 19.5) = 612\ \text{J} = 0612\ \text{kJ}$

$$n = \frac{m}{MM} = \frac{8\ 3}{137\ 3 + (2\times 35\ 45)} = 00399\ \text{mol}$$

$$\Delta H = \frac{0\ 612}{0\ 0399} = -15\ \text{kJ mol}^{-}$$

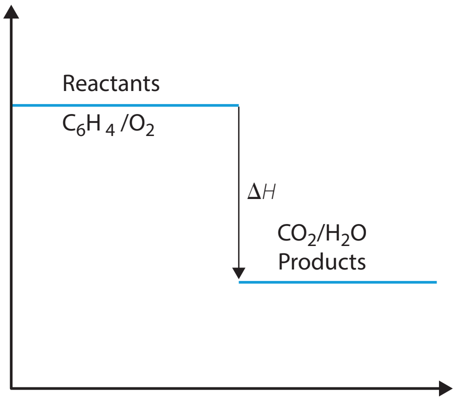
Check your understandng 14.5

1 **a** Use a spirit burner of fuel that is weighed before and after burning to get the mass of fuel burnt, and hence the number of moles of fuel. The flame from the spirit burner is used to heat a measured and known mass of water that has its initial and final temperature measured to calculate the rise in temperaure. The heat of combustion can then be calculated.

b Heat loss to the environment (air, beaker, tripod, other equipment), and incomplete combustion of the fuel producing less heat than forpdete combustion.

2 **a** $\Delta H = -4160\ \text{kJ mol}^{-}$

b



3 $q = mc\Delta T = 500 \times 4.18 \times (290 - 15.3) = 28633\ \text{J} = 28633\ \text{kJ}$

$$n = \frac{m}{MM} = \frac{0\ 58}{(4\times 12\ 01) + (10\times 1\ 008)} = 0009979\ \text{mol}$$

$$\Delta H = \frac{28\ 633}{0\ 009979} = -2900\ \text{kJ mol}^{-}$$

4 $q = mc\Delta T = 350 \times 4.18 \times 77 = 112651\ \text{J} = 11265\ \text{kJ}$

$1360 \times 05 = 680\ \text{kJ mol}^{-}$ (50% heat escapes)

$$n = \frac{112\ 65}{680} = 01657\ \text{mol}$$

$m = n \times MM = 01657 \times ((2 \times 12.01) + (6 \times 1.008) + 16.00)) = 7.6\ \text{g}$

5 **a** $q = mc\Delta T = 1750 \times 4.18 \times 202 \ = 147763\ \text{J} = 148\ \text{kJ}$

$$\frac{148}{9\ 50} = 15.6\ \text{kJ g}^{-}$$

b $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l)$

c $$n = \frac{m}{MM} = \frac{9\ 50}{(6 \times 12\ 01) + (12 \times 1\ 008) + (6 \times 16\ 00)} = 00527\ \text{mol}$$

$$\Delta H = \frac{148}{0\ 0527} = -2810\ \text{kJ mol}^{-}$$

6 **a** $MM = 2 \times 1008 \ = 2016\ \text{g mol}^{-}$

$$\frac{285}{2\ 016} = 141\ \text{kJ g}^{-}$$

ii $MM = (2 \times 12.01) + (6 \times 1.008) + 16.00 = 46068\ \text{g mol}^{-}$

$$\frac{1360}{46\ 068} = 2952\ \text{kJ g}^{-}$$

iii $MM = (2 \times 12.01) + (2 \times 1008) \ = 26036\ \text{g mol}^{-}$

$$\frac{1300}{26\ 036} = 4993\ \text{kJ g}^{-}$$

v $MM = (3 \times 12.01) + (8 \times 1008) \ = 44094\ \text{g mol}^{-}$

$$\frac{2220}{44\ 094} = 5035\ \text{kJ g}^{-}$$

v $MM = (12 \times 12.01) + (22 \times 1.008) + (11 \times 1600) \ = 342296\ \text{g mol}^{-}$

$$\frac{5650}{342\ 296} = 16.51\ \text{kJ g}^{-}$$

v $MM = (8 \times 12.01) + (18 \times 1008) \ = 114224\ \text{g mol}^{-}$

$$\frac{5460}{114\ 224} = 4780\ \text{kJ g}^{-}$$

b Hydrogen, propane, acetylene, octane, ethanol, sucrose

c Have a very high heat released per gram comparedto their enthalpy value.

d Lowers the heat produced per gram significantly due to the higher moecular mass of the oxygen atom.

7 $m(C_2H_5OH) = 15566 \ - 15475 \ = 0.91\ \text{g}$

$q = mc\Delta T = 250 \times 4.18 \times (373 \ - 19.5) = 18601\ \text{J} = 18601\ \text{kJ}$

$MM = (2 \times 12.01) + (6 \times 1.008) + 16.00 = 46068\ \text{g mol}^{-}$

$$\frac{18\ 601}{0\ 91} = 2044\ \text{kJ g}^{-}$$

$$n = \frac{m}{MM} = \frac{0\ 91}{(2 \times 12\ 01) + (6 \times 1\ 008) + 16\ 00} = 001975\ \text{mol}$$

$$\Delta H = \frac{18\ 601}{0\ 01975} = -942\ \text{kJ mol}^{-}$$

a Loss of heat to the surroundings, including the air,beaker and thermometer. Incomplete combustion of the fuel also causes less heat per gram to be producedaffecting the final result.

b That all heat produced from the combustion of the fuel goes directly to heat the water and is not lost to the surroudings. There is no loss of heat from the watr during the experiment.

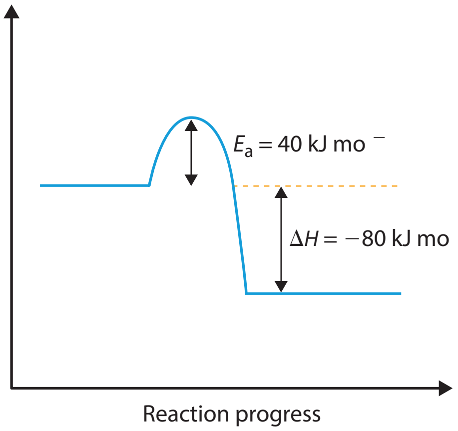
Check your understandng 14.6

1 **a** See Figure14.6b page 335

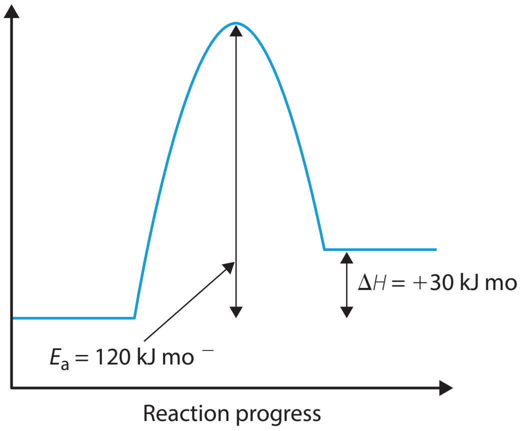
b See Figure14.6a page 335

2 Enthalpy change does not change whn a catalyst is used. The activation energy is lower when a catalyst is used than when a catalyt is not used.

3 **a**



b



4 Diagram 1: $\Delta H = -30$ kJ mol^{-} , $E_a = 40$ kJ mol^{-}

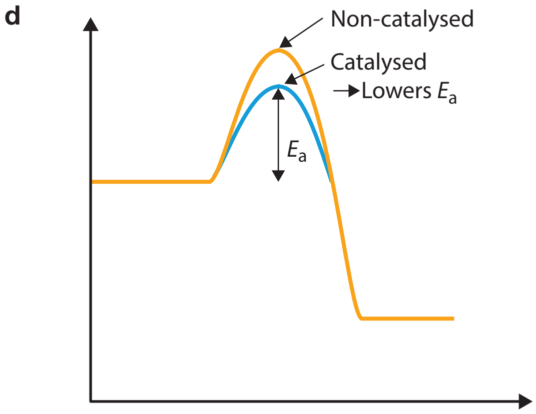
Diagram 2: $\Delta H = +15$ kJ mol^{-} , $E_a = 45$ kJ mol^{-}

5 **a** Endothermic = C Exothermic = A, B andD.

b Greatest = C (greatest difference between the reactants and th top of the energy peak). Smallest = D

c D has the smallest activation energy, so any increase in temperature (in terms of percentage) would have a greate impact. C has the greatest activation energy, so any increase in temperature would have less effect on the rate.

d



Chapter review questions

1 $q = mc\Delta T$, where q = heat absorbed (or released), m = mass of substance, c = specific heat capacity of substance, ΔT = temperature decrease (or increase)of the substance.

2 Exothermic – heat is produced during an exothermic reaction so the surroundings get hotter.

3 See Figure14.6b page 335

4 Measure out a known mass of water and place into a polystyrene cup. Measure the initial temperature of the water. Measure out a known mass of lithium metal and add to the water. Measure the temperature of the water on completion of the reaction. Use $q = mc\Delta T$ to calculate the amount of heat released or absorbed in joules. Convert this value to kilojoules. Calculate the number of moles of lithium used, and calculate the enthalpy in kJ mol^{-} by dividing the heat released/absorbed (in kJ) by the number of moles.

5 $\Delta H = -5460$ kJ mol^{-}

6 **a** $KNO_3(s) \rightarrow K^+(aq) + NO_3^-(aq)$, $\Delta H = +36$ kJ mol^{-}

b The enthalpy value is positive so the reaction is endothermic. Heat is absorbed from the surroundings by the reaction, so the beaker would get colder.

7 **a** Check with your teacher, but it should be based on the experiment you designed in Investigation 14.3. The basic outline is as follows: Use a spirit burner of fuel that is weighed before and after burning to get the mass of fuel burnt, and hence the number of moles of fuel. The flame from the spirit burner is used to heat a measured and known mass of water that has its initial and final temperature measured to calculate the rise in temperature. The heat of combustion can then be calculated.

b Loss of heat to the surroundings is the main error. This can be minimised by using a heat shield (aluminium foil on cardboard to reflect heat back into the equipment), and by using a lid to reduce the amount of heat lost from the water.

8 $q = mc\Delta T = 64 \times 039 \times (800 - 18) = 1952 \times 10^4$ J = 20 kJ

9 $q = mc\Delta T = 250 \times 4180 \times (45 - 16) = 3.031 \times 10^7$ J $= 3.031 \times 10^4$ kJ

10 **a** $\Delta H = 2 \times -286 = -572$ kJ mol^{-}

b $\Delta H = -(-286) = +286$ kJ mol^{-}

11 **a** $1000 \times 070 = 700$ g in 100 L

$$n = \frac{m}{MM} = \frac{700}{(8 \times 12\ 01) + (18 \times 1\ 008)} = 6128 \text{ mol}$$

Heat released = 5460 × 6128 = 33460 kJ

b $\frac{33460}{2200} = 1520$ mol propane required

$m = n \times MM = 1520 \times ((3 \times 12.01) + (8 \times 1008)) = 6702$ g

$\frac{670\ 2}{049} = 1368$ mL $= 1.4$ L

c octae : propane = 1.00 L : 1.4 L

Thus .4 × \$145 = \$203

12 a $2HNO_3(aq) + Ba(OH)_2(aq) \rightarrow Ba(NO_3)_2(aq) + 2H_2O(l)$

b $n = c \times V = 0430 \times 0050 = 00215$ mol

c $n(HNO_3) = c \times V = 0386 \times 0100 = 00386$ mol

If all HNO_3 reacts then $0.5 \times 00386 = 00193$ mol $Ba(OH)_2$ required, so the barium hydroxide is in excess and the nitric acid is the limiting reagent.

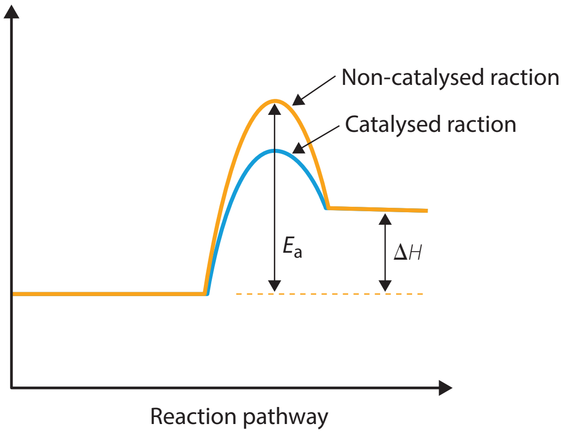
$q = mc\Delta T = 150 \times 4.18 \times (19.6 - 16.3) = 2069$ J $= 2069$ kJ

$\Delta H = \frac{2069}{00386} = -54$ kJ mol^{-}

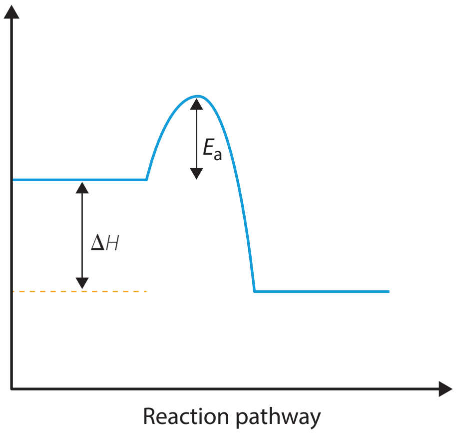
d Heat loss from the solution before the final temperature is recorded (which will then be lower than the actul). Heat loss to the surroundings (air, polystyrene cup).

In the calculation, there is an assumption made that all solutions have the same density as water 1.00 g = 100 mL, which is nottrue.

13 a, b



c



d $\Delta H = -64$ kJ mol^{-}

14 a $N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(l)$

$\frac{1\,0 \times 10^6}{580} = 1724\text{ mol}$

$m = n \times MM = 1724 \times ((2 \times 14.01) + (4 \times 1008)) = 5526 \;\times 10^4\text{ g} = 55\text{ kg}$

b $n(O_2) = n(N_2H_4) = 1724\text{ mol}$

$m(O_2) = n \times MM = 1724 \times (2 \times 1600) = 55168\text{ g} = 55\text{ kg}$

15 a $4CH(s) + 5O_2(g) \rightarrow 4CO_2(g) + 2H_2O(l)$

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

$2C_8H_{18}(s) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l)$

$C_2H_6O(s) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$

b Coal : carbon dioxide = 1 1, thus 455 kJ relased.

Methane : carbon dioxide = 1 1, thus 890 kJ relased.

Octane : carbon dioxide = 1 : 3, thus $\frac{5460}{3}$ = 1820 kJ released.

Ethaol : carbon dioxide = 1 : 2, thus $\frac{1360}{2}$ = 680 kJ released.

c coal, ethanol, methane, octane

d octane, methane, ethanol, coal

e The amount of fossil fuels (trucks, trains, excavators, machinery) that is required to extract/mine and process/transpor the fuels.

STUDENT BOOK ANSWERS

Module four: Drivers of reactions

Chapter 15: Enthalpy and Hess's law

Worked examle 1.1

$H_2(g) + \frac{1}{2} O_2(g) \rightarrow H_2O(l),\ \Delta H\ \ = -285$ kJ mol^{-}

$\frac{1}{2} O_2(g) + H_2O(l) \rightarrow H_2O_2(l),\ \Delta H_2 = +97$ kJ mol^{-}

Overall: $H_2(g) + O_2(g) \rightarrow H_2O_2(l),\ \Delta H = \Delta H\ \ + \Delta H_2 = -285 + 97 = -188$ kJ mol^{-}

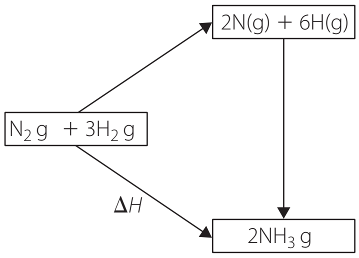
Check your understandng 15.1/15.2

1 Hess's law is used to calculate enthalpy changes for reactions that do not allow measurement of ΔH direcly. This is done by using the enthalpy values for other, more easily measured reactions.

2 **a** $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$

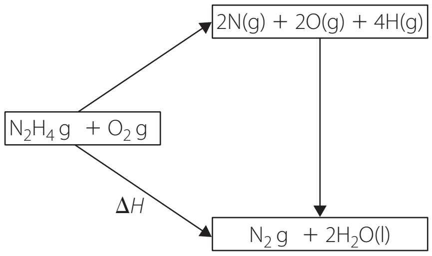
b The covalent bonds between nitrogen atoms, and the covalent bonds between hydrogen atoms must be broen. Bonds need to form between nitroen and hydrogen atoms.

c



3 **a** $N_2H_4(g) + O_2(g) \rightarrow N_2(g) + 2H_2O(l)$

b



4 **a** $\Delta H = 05\ \ \times -198 = -99$ kJ mol^{-}

b $\Delta H = -(-198) = +198$ kJ mol^{-}

5 Reactions add together, so hydroxide ions cancel on both sides.

$\Delta H = -56 + -55 = -111$ kJ mol^{-}

6 First reaction stays as written ($\Delta H = -178$ kJ mol^{-}), second reaction reverses ($\Delta H = +82$ kJ mol^{-}) When added, the calcium oxide cancels n both sides.

$\Delta H = -178 + 82 = -96$ kJ mol^{-}

7 First reaction stays as written ($\Delta H = +66$ kJ mol^{-}), second reaction reverses ($\Delta H = +57$ kJ mol^{-}) When added, the nitrogen dioxide cancelson both sides.

$\Delta H = +66 + 57 = +123$ kJ mol^{-}

Worked examle 1.2

1 $H_2S(g) + SO_2(g) \rightarrow S(s) + H_2O(g)$

$\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(H_2O{,}g) - \{\Delta H^{\ominus}{}_f(H_2Sg) + \Delta H^{\ominus}{}_f(SO_2{,}g)\}$

$\Delta H^{\ominus} = -242 - (-21 + -297) = +76$ kJ mol^{-}

2 $3CO(g) + Fe_2O_3(s) \rightarrow 2Fe(s) + 3CO_2(g)$

$\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(CO_2\,g) - \{\Delta H^{\ominus}{}_f(C{,}g) + \Delta H^{\ominus}{}_f(Fe_2O_3{,}s)\}$

$\Delta H^{\ominus} = 3(-393) - \{3(-111) + -823\} = -23$ kJ mol^{-}

Worked examle 1.3

$C_2H_6O(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l)$, $\Delta H = -1360$ kJ mol^{-}

$\Delta H^{\ominus} = \{\Delta H^{\ominus}{}_f(CO_2\,g) + \Delta H^{\ominus}{}_f(H_2O{,}l)\} - \{\Delta H^{\ominus}{}_f\,(C_2H_6O{,}l)\}$

$-1360 = \{2(-393) + 3(-285)\} - \{\Delta H^{\ominus}{}_f\,(C_2H_6O{,}l)\}$

$\Delta H^{\ominus}{}_f\,(C_2H_6O{,}l) = \{2(-393) + 3(-285)\} + 1360$

$\Delta H^{\ominus}{}_f\,(C_2H_6O{,}l) = -\,281$ kJ mol^{-}

Check your understandng 15.3/15.4/15.5

1 **a** The stable form of the substance (solid, lquid or gas) at 100.0 kPa and the specified temperature.

b The substance present at a pressre of 100.0 kPa.

c The substance present at a concntration of 1.00 mol L $^{-}$

2 The increase in enthalpy when one mole of a compound in its standard state is formed from its elements in their standard states. For example,formation of water:

$H_2(g) + \frac{1}{2}\,O_2(g) \rightarrow H_2O(l)$, $\Delta H = -285$ kJ mol^{-}

3 Write the balanced equation for the coustion of he substance. Use the formula:

$\Delta H^{\ominus} = \Sigma\Delta H^{\ominus}{}_f(\text{products}) - \Sigma\Delta H^{\ominus}{}_f(\text{reactants})$

to calculate the enthalpy of formation ofthe reactant in question.

4 liquid water, gaseous bromine, liquid chlorine, gaseous sulfur

5 **a** $2Al(s) + \frac{3}{2}\,O_2(g) \rightarrow Al_2O_3(s)$

b $P(s) + \frac{3}{2}\,Cl_2(g) \rightarrow PCl_3(g)$

c $2Na(s) + S(s) + 2O_2(g) \rightarrow Na_2SO_4(s)$

d $\frac{3}{2}\ \mathrm{H_2(g) + B(s)} + \frac{3}{2}\ \mathrm{O_2(g) \rightarrow H_3BO_3(s)}$

e $\mathrm{Br_2(l) \rightarrow Br_2(g)}$

f $\mathrm{H_2(g)} + \frac{1}{2}\mathrm{O_2(g) \rightarrow H_2O(l)}$

6 $-396 = -297 + \Delta H$

$\Delta H = -396 + 297 = -99\ \text{kJ mol}^{-}$

7 **a** $\Delta H^\circ = \Delta H^\circ{}_f(\mathrm{CaCO_3,s}) - \{\Delta H^\circ{}_f(\mathrm{CaO,s}) + \Delta H^\circ{}_f(\mathrm{CO_2,g})\}$

$\Delta H^\circ = (-1207) - (-636 + -393) = -178\ \text{kJ mol}^{-}$

b $\Delta H^\circ = \Delta H^\circ{}_f(\mathrm{Al_2O_3,s}) - \Delta H^\circ{}_f(\mathrm{Fe_2O_3\ s})$

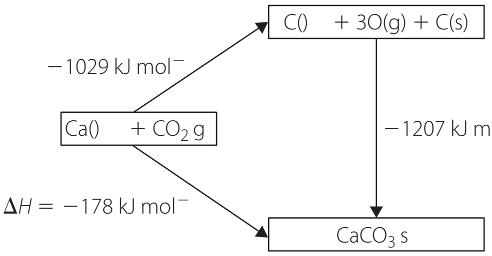
$\Delta H^\circ = -1670 - (-823) = -847\ \text{kJ mol}^{-}$

c $\Delta H^\circ = \{3\Delta H^\circ{}_f(\mathrm{SO_2\ g}) + 2\Delta H^\circ{}_f(\mathrm{H_2O,l})\} - \{2\Delta H^\circ{}_f(\mathrm{H_2SO_4,l})\}$

$\Delta H^\circ = \{3(-297) + 2(-285)\} - 2(-814) = +167\ \text{kJ mol}^{-}$

d $\Delta H^\circ = 2\Delta H^\circ{}_f(\mathrm{PBr_3,l}) = 2 \times -185 = -370\ \text{kJ mol}^{-}$

8 **a**



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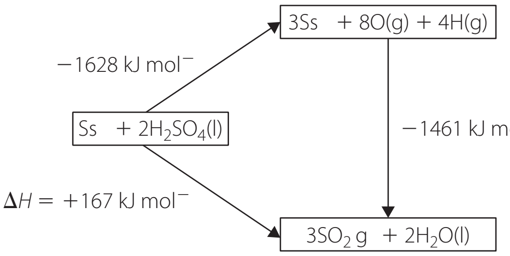


9 $\mathrm{C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)}$

$\Delta H^\circ = \{2\Delta H^\circ{}_f(\mathrm{CO_2\ g}) + 2\Delta H^\circ{}_f(\mathrm{H_2O,l})\} - \Delta H^\circ{}_f(\mathrm{C_2H_4\ g})$

$-1411 = \{2(-393) + 2(-285)\} - \Delta H^\circ{}_f(\mathrm{C_2H_4\ g})$

$\Delta H^\circ{}_f(\mathrm{C_2H_4}) = \{2(-393) + 2(-285)\} + 1411$

$\Delta H^\circ{}_f(\mathrm{C_2H_4}) = +55\ \text{kJ mol}^{-}$

10 $\mathrm{C_4H_{10}(g)} + \frac{13}{2}\ \mathrm{O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)}$

$\Delta H^\circ = \{4\Delta H^\circ{}_f(\mathrm{CO_2\ g}) + 5\Delta H^\circ{}_f(\mathrm{H_2O,l})\} - \Delta H^\circ{}_f(\mathrm{C_4H_{10}\ g})$

$-2880 = \{4(-393) + 5(-285)\} - \Delta H^\circ{}_f(\mathrm{C_4H_{10}\ g})$

$\Delta H^\circ{}_f(\mathrm{C_4H_{10}}) = \{4(-393) + 5(-285)\} + 2880$

$\Delta H^\circ{}_f(\mathrm{C_4H_{10}}) = -117\ \text{kJ mol}^{-}$

Worked example 1.4

ΔH° is the negative of the enthalpy of formation of methane = +75 kJ mol^{-}

$\Delta H^{\circ}{}_2$ is the heat of atomisation of carbon = 717 kJ mol^{-}

$\Delta H^{\circ}{}_3$ is the 4 × heat of atomisation of hydrogen = 4 × 218 = 872 kJ mol^{-}

$\Delta H^{\circ}{}_4$ is $4 \times B_{C\text{–}H}$

$4 \times B_{C\text{–}H} = \Delta H_4 = \Delta H^{\circ} + \Delta H^{\circ}{}_2 + \Delta H^{\circ}{}_3 = +75 + 717 + 872 = 1664$

$$B_{C\text{–}H} = \frac{1664}{4} = 416 \text{ kJ mol}^{-}$$

Worked example 1.5

Broken: 4 × C—H and 4 × Cl—Cl

Formed: 4 × C—Cl and 4 × H—Cl

$\Delta H^{\circ} = (4B_{C\text{–}H} + 4B_{Cl\text{–}l}) - (4B_{C\text{–}Cl} + 4B_{H\text{–}Cl})$

$\Delta H^{\circ} = (4 \times 413 + 4 \times 242) - (4 \times 339 + 4 \times 432) = -464$ kJ mol^{-}

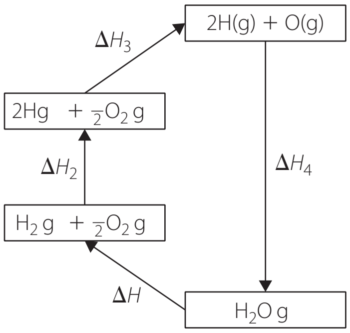
Check your understanding 15.6/15.7

1 ΔH° is the negative of the enthalpy of formation of gaseous water

$\Delta H^{\circ}{}_2$ is 2 × the heat of atomisation of hydrogen

$\Delta H^{\circ}{}_3$ is the 4 × heat of atomisation of oxygen

$\Delta H^{\circ}{}_4$ is $2 \times B_{O\text{–}H}$



2 The greater the bond energy of the weakest bond, the more stable the compound is and the less reactive it is.

3 2 × 473 = 946 kJ mol^{-}

4 **a** Broken: 1 × C—H bond, 1 × Br—Br bond

Formed: 1 × C—Br, 1 × H—Br

b Broken: 1 × C—Cl, 1 × O—H

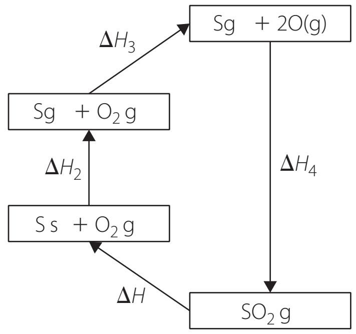
Formed: 1 × C—O, 1 × H—Cl

5 **a** $\Delta H^{\ominus}$ is the negative of the enthalpy of formation of sulfur dioxide

$\Delta H^{\ominus}_{2}$ is the heat of atomisation of sulfur

$\Delta H^{\ominus}_{3}$ is the $2 \times$ heat of atomisation of oxygen

$\Delta H^{\ominus}_{4}$ is $2 \times B_{\text{O-S}}$



b $\Delta H^{\ominus} = +297$ kJ mol^{-}, $\Delta H^{\ominus}_{2} = 279$ kJ mol^{-}, $\Delta H^{\ominus}_{3} = 2 \times 249 = 498$ kJ mol^{-}

$\Delta H^{\ominus}_{4} = 2B_{\text{O-S}} = +297 + 279 + 498 = 1074$

$B_{\text{O-S}} = \dfrac{1074}{2} = 537$ kJ mol^{-}

6 **a** $\Delta H^{\ominus} = +46$ kJ mol^{-}, $\Delta H^{\ominus}_{2} = 2 \times 473 = 946$ kJ mol^{-}, $\Delta H^{\ominus}_{3} = 6 \times 218 = 1308$ kJ mol^{-}

b $\Delta H^{\ominus}_{4} = 2300$ kJ mol^{-}

c $6B_{\text{N-H}} = \Delta H_{4} = 2300$, thus $\dfrac{2300}{6} = 383 = B_{\text{N-H}}$

7 **a** 1 × N—N, 4 × N—H

b Ammonia has 3 × N—H bonds

N—H = 391 kJ mol^{-} N—N = 163 kJ mol^{-}

As hydrazine has the weakest bond, it will be more reactive and less stable than ammonia.

8 Broen: 1 × N—N, 4 × N—H, 1 × O═O

Fored: 1 × NN, 4 × H—O

$\Delta H^{\ominus} = (B_{\text{N—N}} + 4B_{\text{N—H}} + B_{\text{O O}}) - (B_{\text{NN}} + 4B_{\text{H-O}})$

$\Delta H^{\ominus} = (163 + 4 \times 391 + 498) - (945 + 4 \times 463) = -572$ kJ mol^{-}

9 Bond energy for C—Cl is 339, while forC—F it is 485. Tetrachloromethane contains the weaker bonds of the two molecules so is less stable nd more reactive. Thus, it is more likely to decompose under UV light.

10 Broen: 4 × H—S, 3 × O—O

Fored: 4 × SO, 4 × O—H

$\Delta H^{\ominus} = (4B_{\text{S—H}} + 3B_{\text{OO}}) - (4B_{\text{SO}} + 4B_{\text{H—O}}) = (4 \times 360 + 3 \times 302) - (4 \times 540 + 4 \times 463)$

$\Delta H^{\ominus} = -1666$ kJ mol^{-}

Check your understandng 15.8

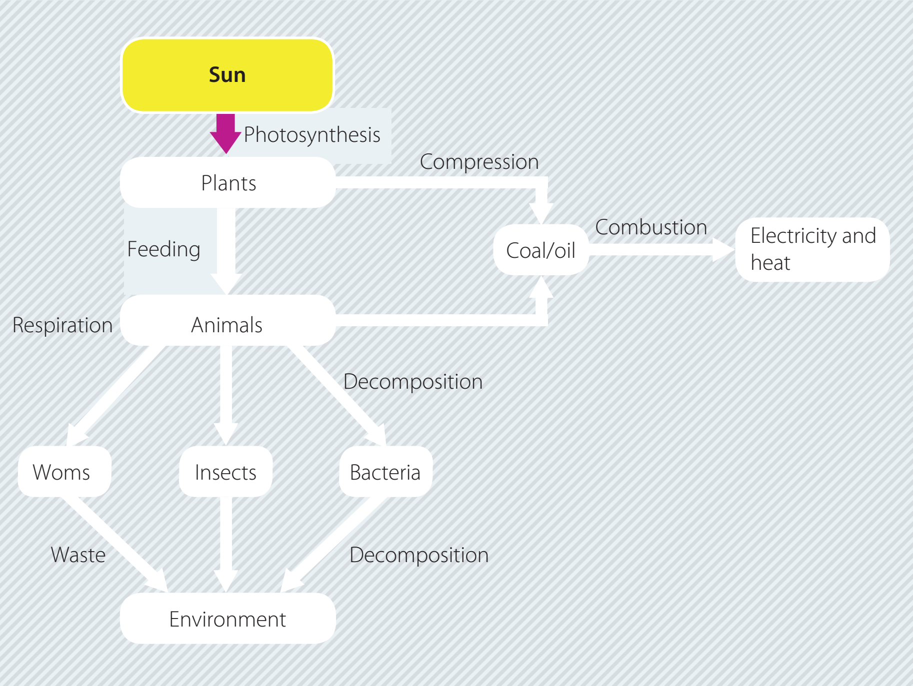
1 **a** $12CO_2(g) + 11H_2O(l) \rightarrow C_{12}H_{22}O_{11}(s) + 12O_2(g)$, $\Delta H\ \ = +5640$ kJ mol

$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$, $\Delta H_2 = +55$ kJ mol^{-}

$\Delta H = \Delta H\ \ + \Delta H_2 = +5640 + 5.5 = +56455$ kJ mol^{-}

b Energy comes fromthe sun.

2 **a, b**



c The energy in fossil fuels comes indirectly from the sun through photosynthesis into plants, then by feeding into animals. Animals and plant remains become fossil fuels, thus transferring the sun's energy into the fuels.

3 Energy stored in one mole of gluose is 2814 kJ.

Orange liht: $\dfrac{2814}{8400} \times 100 = 335\%$

Blue light: $\dfrac{2814}{12000} \times 100 = 235\%$

Hence the orange light is more efficient as a greater proportion of its eergy ends up as stored glucose.

Chapter reiew quesions

1 The law of conservation of energy states that energy cannot be created or destroyed, but only changed from one form t another. Hess's law is a version of this law that states the enthalpy change from reactants to products is the same no matter which chemical pathway is used, or how many steps are involved in eachpathway.

2 Stability is a measure of how difficult it is to decompose the compound in the absence of other substances, while reactivity is how easily the compound undergoes chemical reaction when mixed with other subsances. Stability is about keepng its structure intact. Reactivity describes what happens when substances undergo chemical changes to orm new substances.

3 **a** 0 kJ mol^{-}

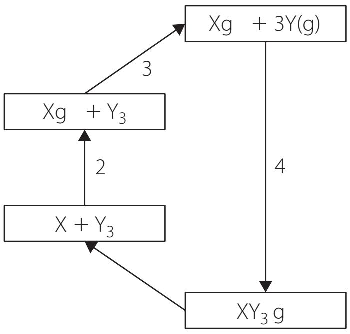
b Enthalpy of formation definition is about formation of a substance from its elements in their standard states, thus the elements cannot have an ethalpy of formation.

4 See Figure15.7 page 352

5 Broen: 2 × C—H, 2 × Br—Br

Fored: 2 × C—Br, 2 × H—Br

6 The reverse of the enthalpy of formation: $XY_3 \rightarrow X + 3Y$ (1 in the diagram) added to the heat of atomisation of X (2 in the diagram), and the heat of atomisation of 3Y (3 in the diagram) will be equal to 3 times the bond energy of X—Y (4 in the diagram).



7 $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_{12}(s) + 6O_2(g),\ \Delta H\ = X$

$C_6H_{12}O_6(s) \rightarrow C_6H_{12}O_6(aq),\ \Delta H_2 = Y$ kJ mol^{-}

$\Delta H = \Delta H\ + \Delta H_2 = X + Y$ kJ mol^{-}

8 Reverse equation 1: $\Delta H = +133$ kJ mol^{-}

Half equation 2: $\Delta H = -223$ kJ mol^{-}

$LiOH(aq) + H_2(g) \rightarrow LiH(s) + H_2O(l)$

$Li(s) + H_2O(l) \rightarrow LiOH(aq) + \frac{1}{2} H_2(g)$

When added, the LiOH and the H_2O and half of the H_2 in the first equation cancel, leaving the desired equation:

$\Delta H = +133 + -223 = -90$ kJ mol^{-}

9 $CaH_2(s) + 2H_2O(l) \rightarrow 2H_2(g) + Ca(OH)_2(s),\ \Delta H = -244$ kJ mol^{-}

$\Delta H^{\ominus} = \Delta H^{\ominus}_f(Ca(OH)_2,s) - \{\Delta H^{\ominus}_f(CaH_2,s) + 2\Delta H^{\ominus}_f(H_2O,l)\}$

$-244 = -987 - \{\Delta H^{\ominus}_f(CaH_2,s) + 2(-285)\}$

$\Delta H^{\ominus}_f(CaH_2,s) = -987 + 244 + 570$

$\Delta H^{\ominus}_f(CaH_2,s) = -193$ kJ mol^{-}

10 $CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g),\ \Delta H = -282$ kJ mol^{-}

$2H_2(g) + O_2(g) \rightarrow 2H_2O(l),\ \Delta H = -570$ kJ mol^{-}

$CO_2(g) + 2H_2O(l) \rightarrow CH_3OH(l) + 2O_2(g),\ \Delta H = +725$ kJ mol^{-}

Adding all three together cancels oxygen, carbon dioxide and water.

$\Delta H = -282 + -570 + 725 = +73$ kJ mol^{-}

11 a $\Delta H^{\ominus} = 2\Delta H^{\ominus}_{f}(HI,g) - \Delta H^{\ominus}_{f}(I_2\,g)$

$\Delta H^{\ominus} = (2 \times +26) - (+62) = -10$ kJ mol

b $\Delta H^{\ominus} = \{\Delta H^{\ominus}_{f}(I_3^{-}\,aq) + 2\Delta H^{\ominus}_{f}(Cl^{-},aq)\} - 3\Delta H^{\ominus}_{f}(I^{-}\,g)$

$\Delta H^{\ominus} = \{-52 + (2 \times -167)\} - (3 \times -55) = -221$ kJ mol^{-}

c $\Delta H^{\ominus} = 2\Delta H^{\ominus}_{f}(H_3PO_4,l) - \{\Delta H^{\ominus}_{f}(P_2O_5,s) + 3\Delta H^{\ominus}_{f}(H_2O,l)\}$

$\Delta H^{\ominus} = (2 \times -1492) - \{-1280 + (3 \times -188)\} = -2984 - (-1844) = -1140$ kJ mol^{-}

12 $-21 + 279 + 436 = 2$ (S—H bond energy)

(S—H bond energy) = +347 kJ mol^{-}

13 a O—H = 463, H—S = 360 Water has the highest bond energies so it is more stable than hydrogen sulfide.

b C—H = 413, C—Br = 285 Bromomethane has the weakest bond (the C—Br bond) so is the least stable. Methane is the more stable molecule.

c The bond between oxygen atoms in oxygen (498) is stronger than in ozone (302), so the oxygen molecule is more stable than ozone.

14 a C—H = 413, C—Cl = 339 Thus, the C—Cl bond will break first when the compounds are heated as it has the lowest bond energy.

b C—H = 413, C—F = 485 Thus, the C—H bond will break first when the compounds are heated as it has the lowest bond energy

15 CCl_4 has bonds C—Cl with a bond energy of 339. CF_4 has C—F bonds with bond energy of 485. CCl_4 contains the weaker bonds of the two molecules so is less stable and more reactive. Thus it is more likely to decompose under UV light. Thus bond energies are extremely useful in discussing the stability and reactivity as both properties are directly linked to quantitative values.

16 a $5C(s) + 5H_2(g) \rightarrow C_5H_{10}(g)$

$5C(s) \rightarrow 5C(g) = 5 \times 717 = 3585$ kJ mol^{-}

Bonds broken: 5 × H—H bonds = $5 \times 436 = 2180$ kJ mol^{-}

Bonds formed:

5 × C—C bonds = $5 \times 348 = 1740$ kJ mol^{-}

10 × C—H bonds = $10 \times 413 = 4130$ kJ mol^{-}

$\Delta H^{\ominus} = (3585 + 2180) - (1740 + 4130) = -105$ kJ mol^{-}

b $C_5H_{10}(g) + \frac{15}{2} O_2(g) \rightarrow 5CO_2(g) + 5H_2O(l)$

$\Delta H^{\ominus} = \{5\Delta H^{\ominus}_{f}(CO_2\,g) + 5\Delta H^{\ominus}_{f}(H_2O,l)\} - \Delta H^{\ominus}_{f}(C_5H_{10}\,g)$

$-3320 = (5 \times -393) + (5 \times -285) - \Delta H^{\ominus}_{f}(C_5H_{10}\,g)$

$\Delta H^{\ominus}_{f}(C_5H_{10}\,g) = -70$ kJ mol^{-}

c Bond energies are not exactly the same in all compounds. Values for bond energies are averages taken over a wide range of compounds containing that bond. Hence, the answer for (b) would be more accurate.

17 Br2(l) → Br2(g)

Bonds boken: 1 × Br—Br = 193 kJ mol^{-}

$\Delta H^{\ominus}{}_{\text{a}}$ (Br) = 112 kJ mol^{-}

Bonds frmed: 1 × Br—Br = 193 kJ mol^{-}

$\Delta H^{\ominus}{}_{\text{vap}} = 193 - 193 + (2 \times 112) = 224$ kJ mol^{-}

STUDENT BOOK ANSWERS

Module four: Drivers of reactions

Chapter 16: Entropy and Gibbs free energy

Worked examle 1.1

1 a 4 moles of gas produce 2 moes of gas. Effectively 2 moles of gas have been lost, thus there is less randomness in the products, the entropy drive is to the reverse reaction, and the entropy for the reaction is ngative.

b 5 moles of gas are produced from a solid, thus increasing the entropy in the forward direction, thus the entropy for the reaction is positive.

2 a It is an exothermic reaction so the energy drive is i the forward direction. As the entropy drive is in the reverse direction, the energy drive must be larger than the entropy drive.

b It is an endothermic reaction so the energy drive is i the reverse direction. As the entropy drive is in the forward direction, it must be larger han the energy drive.

Worked examle 1.2

$$\Delta S^{\ominus} = \{S^{\ominus}(\mathrm{CO}) + 3S^{\ominus}(\mathrm{H_2})\} - \{S^{\ominus}(\mathrm{CH_4}) + S^{\ominus}(\mathrm{H_2O})\}$$
$$\Delta S^{\ominus} = (198 + (3 \times 131)) - (186 + 189)$$
$$\Delta S^{\ominus} = 216\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$$

Check your understandng 16.1/16.2

1 There is more randomness in a solutio than in a solid. Increased randomness results in an increase in entropy.

2 a Negative – solution/gas reactants have more randomnes than the solid product.

b Energy drive is forward (exothermic), entropy drive is reverse.

c Energy drive is stronger.

d Reaction is as written, so the forward drive (eergy) must be stronger.

ii a Positive – two moles of gas product has more disorder than theone mole of gas reactant.

b Energy drive is reverse (endothermic), entropy drive is forward.

c Entropy drive isstronger.

d Reaction is as written, so forward drive (entropy) must be stronger.

iii a Positive – 5 moles of gas product has more disorder than 3 moles gas/1 nle of solution reactants.

b Energy drive is forward (exothermic), entropy drive is forward.

c Cannot decide.

d Both drives are forward, so no way to deterine which is larger.

v **a** Negative – 3 moles of reactant gas has more disorder than moles of product gas.

b Energy drive is forward (exothermic), epyr drive is reverse.

c Energy drive is stronger.

d Reaction is as written, so forward drive (enrgy) must be stronger.

v **a** Positive – gas products have more disorder than he solid reactants.

b Energy drive is reverse (endothermic), epyr drive is forward.

c Entropy drive isstronger.

d Reaction is as written, so forward drive (entropy) must be stronger.

3 Energy drive is everse. The beaker getting cold indicates an endothermic reaction. Entropy drive is forward as there is more disorder in the solution pro uct than the solid reactant. The entropy drive must be stronger as the reaction occurs as written, so th forward drive is stronger.

4 **a** negative

ii negative

iii positive

b $\Delta S^{\ominus} = 2S^{\ominus}(\mathrm{NH_3}) - \{S^{\ominus}(\mathrm{N_2}) + 3S^{\ominus}(\mathrm{H_2})\}$

$\Delta S^{\ominus} = (2 \times 1923) - (192 + 3 \times 131)$

$\Delta S^{\ominus} = -200\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

ii $\Delta S^{\ominus} = 2S^{\ominus}(\mathrm{CO_2}) - \{2S^{\ominus}(\mathrm{CO}) + S^{\ominus}(\mathrm{O_2})\}$

$\Delta S^{\ominus} = (2 \times 214) - (2 \times 198 + 205)$

$\Delta S^{\ominus} = -173\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

iii $\Delta S^{\ominus} = \{2S^{\ominus}(\mathrm{H_2O}) + S^{\ominus}(\mathrm{O_2})\} - 2S^{\ominus}(\mathrm{H_2O_2})$

$\Delta S^{\ominus} = (2 \times 189 + 205) - (2 \times 1096)$

$\Delta S^{\ominus} = +3638\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

c All three predictions were accurate.

5 **a** $\Delta S^{\ominus} = 2S^{\ominus}(\mathrm{HBr}) - \{S^{\ominus}(\mathrm{H_2}) + S^{\ominus}(\mathrm{Br_2})\}$

$\Delta S^{\ominus} = (2 \times 1986) - (131 + 2454)$

$\Delta S^{\ominus} = +208\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

b $\Delta S^{\ominus} = \{3S^{\ominus}(\mathrm{SO_2}) + 2S^{\ominus}(\mathrm{H_2O})\} - \{S^{\ominus}(\mathrm{S}) + 2S^{\ominus}(\mathrm{H_2SO_4})\}$

$\Delta S^{\ominus} = (3 \times 248 + 2 \times 189) - (31.8 + 2 \times 1569)$

$\Delta S^{\ominus} = +7764\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

c $\Delta S^{\ominus} = S^{\ominus}(\mathrm{PCl_5}) - \{S^{\ominus}(\mathrm{PCl_3}) + S^{\ominus}(\mathrm{Cl_2})\}$

$\Delta S^{\ominus} = 3642 - (217.15 + 2230)$

$\Delta S^{\ominus} = +37005\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

d $\Delta S^{\ominus} = \{2S^{\ominus}(\text{Fe}) + S^{\ominus}(\text{Al}_2\text{O}_3)\} - \{S^{\ominus}(\text{Fe}_2\text{O}_3) + 2S^{\ominus}(\text{Al})\}$

$\Delta S^{\ominus} = (2 \times 273\ \ + 509)\ - (874\ \ + 2 \times 2833)$

$\Delta S^{\ominus} = -386\ \text{J}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}$

e $\Delta S^{\ominus} = \{2S^{\ominus}(\text{Cl}^{-}) + S^{\ominus}(\text{I}_2)\} - \{S^{\ominus}(\text{Cl}_2) + 2S^{\ominus}(\text{Fe}_2\text{O}_3)\}$

$\Delta S^{\ominus} = (2 \times 565\ \ + 164)\ - (2230\ \ + 2 \times 874)$

$\Delta S^{\ominus} = -268\ \text{J}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}$

Worked examle 1.3

$\Delta H^{\ominus} = 2\ \Delta H^{\ominus}{}_{\text{f}}(\text{SO}_3\,\text{g}) - 2(\ \Delta H^{\ominus}{}_{\text{f}}(\text{SO}_2\,\text{g}) = (2 \times -396) - (2 \times -297) = -198\ \text{kJ mol}^{-}$
$\Delta S^{\ominus} = 2S^{\ominus}(\text{SO}_3) - \{2S^{\ominus}(\text{SO}_2) + S^{\ominus}(\text{O}_2)\} = (2 \times 257) - (2 \times 248 + 205) = -187\ \text{J}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}\ =$
$0187\ \text{kJ}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}$
$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = -198 - (298 \times -0187)\ = -142\ \text{kJ mol}^{-}$ As the value is negative, the reaction proceeds as ritten.

Check your understandng 16.3/16.4

1 Used to indicate standard conditions ar used: pressure of 100.0 kPa, concentration of 1.00 mol L $^{-}$, and at the specified temperature(often 298 K).

2 $\Delta H^{\ominus}$ kJ mol^{-}

$\Delta G^{\ominus}$ kJ mol^{-}

$\Delta S^{\ominus}$ J^{-} K^{-} mol^{-}

3 A spontaneous reaction goes to virtual completin in the direction written. It has a very large negative Gibbs free energy change value ($\Delta G^{\ominus}$)

4 **a** Gibbs free energy change ($\Delta G^{\ominus}$) is negative and the reactioncpeds as written.

b Gibbs free energy change ($\Delta G^{\ominus}$) is negative and vry large.

5 **a, b**

$\Delta H^{\ominus} = \{\ \Delta H^{\ominus}{}_{\text{f}}(\text{HCl,g}) + \Delta H^{\ominus}{}_{\text{f}}(\text{NH}_3\text{,g})\} - \Delta H^{\ominus}{}_{\text{f}}(\text{NH}_4\text{C,s})$

$\Delta H^{\ominus} = (-92 + -46) - (-314) = +176\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(\text{HCl}) + S^{\ominus}(\text{NH}_3\} - S^{\ominus}(\text{NH}_4\text{Cl}) = 1868\ \ + 1923\ \ - 946\ \ = +2845\ \text{J}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = +176 - (298 \times 02845)\ = +9122\ \text{kJ mol}^{-}$

Reaction does not go as written as Gibbs free enegy change is positive.

ii $\Delta H^{\ominus} = 3\ \Delta H^{\ominus}{}_{\text{f}}(\text{CO}_2\,\text{g}) - \{\ \Delta H^{\ominus}{}_{\text{f}}(\text{Fe}_2\text{O}_3\text{,s}) + 3\ \Delta H^{\ominus}{}_{\text{f}}(\text{C,g})\}$

$\Delta H^{\ominus} = (3 \times -393) - (-823 + 3 \times -111) = -23\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{3S^{\ominus}(\text{CO}_2) + 2S^{\ominus}(\text{Fe})\} - \{3S^{\ominus}(\text{CO}) + S^{\ominus}(\text{Fe}_2\text{O}_3)\} = (3 \times 214 + 2 \times 273)\ - (3 \times 198 + 874)\ = +152\ \text{J}^{-}\ \ \text{K}^{-}\ \ \text{mol}^{-}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = -23 - (298 \times 00152)\ = -275\ \text{kJ mol}^{-}$

Reaction goes as written as Gibbs free enerychange is negative.

iii $\Delta H^{\ominus} = 2\,\Delta H^{\ominus}{}_f(\mathrm{N,g}) \quad - 2\,\Delta H^{\ominus}{}_f(\mathrm{N,g})$

$\Delta H^{\ominus} = (2 \times 90) - (2 \times 33) = +114\ \mathrm{kJ\ mol^{-}}$

$\Delta S^{\ominus} = \{2S^{\ominus}(\mathrm{NO}) + S^{\ominus}(\mathrm{O})_2\} - 2S^{\ominus}(\mathrm{NO_2}) = (2 \times 211 + 205) - (2 \times 240) = +147\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 114 - (298 \times 0147) = +110\ \mathrm{kJ\ mol^{-}}$

Reaction does not go as written as Gibbs free energy change is positive.

v $\Delta H^{\ominus} = \{2\,\Delta H^{\ominus}{}_f(\mathrm{H_2O,g}) + 2\,\Delta H^{\ominus}{}_f(\mathrm{SO_2\,g}) + \Delta H^{\ominus}{}_f(\mathrm{CO_2,g})\} - 2\,\Delta H^{\ominus}{}_f(\mathrm{H_2SO_4\,l})$

$\Delta H^{\ominus} = (2 \times -242 + 2 \times -297 + -393) - (2 \times -814) = +157\ \mathrm{kJ\ mol^{-}}$

$\Delta S^{\ominus} = \{2S^{\ominus}(\mathrm{H_2O}) + 2S^{\ominus}(\mathrm{SO_2}) + S^{\ominus}(\mathrm{CO_2})\} - \{S^{\ominus}(\mathrm{C}) + 2S^{\ominus}(\mathrm{H_2SO_4})\} = (2 \times 189) + (2 \times 248) + 214 - (6 + 2 \times 1569) = +768\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 157 - (298 \times 0768) = -72\ \mathrm{kJ\ mol^{-}}$

Reaction goes as written as Gibbs free energy change is negative.

6 Type 3 reactions will go as written if heat is continually supplied. As they require a constant input of energy to proceed, they are classified as non-spontaneous.

7 **a** $\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(\mathrm{MgCO_3,s}) - \{\Delta H^{\ominus}{}_f(\mathrm{Mg^{2+}\,aq}) + \Delta H^{\ominus}{}_f(\mathrm{CO_3{}^{2-}\,aq})\}$

$\Delta H^{\ominus} = (-1096) - (-467 + -677) = +48\ \mathrm{kJ\ mol^{-}}$

$\Delta S^{\ominus} = S^{\ominus}(\mathrm{MgCO_3}) - \{S^{\ominus}(\mathrm{Mg^{2+}}) + S^{\ominus}(\mathrm{CO_3{}^{2-}})\} = 657 - (-1381 + -569) = +261\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 48 - (298 \times 0261) = -30\ \mathrm{kJ\ mol^{-}}$

ii $\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(\mathrm{Na_2CO_3,s}) - \{\Delta H^{\ominus}{}_f(\mathrm{Na^{+}\,aq}) + \Delta H^{\ominus}{}_f(\mathrm{CO_3{}^{2-}\,aq})\}$

$\Delta H^{\ominus} = (-1131) - (-240 + -677) = -214\ \mathrm{kJ\ mol^{-}}$

$\Delta S^{\ominus} = S^{\ominus}(\mathrm{Na_2CO_3}) - \{S^{\ominus}(\mathrm{Na^{+}}) + S^{\ominus}(\mathrm{CO_3{}^{2-}})\} = 135 - (590 + -569) = +133\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = -214 - (298 \times 0.133) = -254\ \mathrm{kJ\ mol^{-}}$

b Both have negative Gibbs free energy change values, so both will proceed as written.

ii Hence, both of them are not very soluble in water as the precipitate will form spontaneously.

Worked example 1.4

$\Delta H^{\ominus} = \{\Delta H^{\ominus}{}_f(\mathrm{CO_2\,g}) + \Delta H^{\ominus}{}_f(\mathrm{CaO,s})\} - \Delta H^{\ominus}{}_f(\mathrm{CaCO_3\,s})$
$\Delta H^{\ominus} = (-393 + -636) - (-1207) = +178\ \mathrm{kJ\ mol^{-}}$
$\Delta S^{\ominus} = \{S^{\ominus}(\mathrm{CO_2}) + S^{\ominus}(\mathrm{CaO})\} - S^{\ominus}(\mathrm{CaCO_3}) = (214 + 398) - 93 = +161\ \mathrm{J^{-}\ K^{-}\ mol^{-}}$
500K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 178 - (500 \times 0.161) = +895\ \mathrm{kJ\ mol^{-}}$
1500K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 178 - (1500 \times 0.161) = -64\ \mathrm{kJ\ mol^{-}}$
The reaction will proceed as written at 1500 K (value is negative), but not at 500 K (value is positive).

Check your understanding 16.5

1 Calculate the enthalpy of formation and entropy change for 298K, then use $T = 900$ in the Gibbs free energy equation $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$

2 If the change in entropy is positive, the value for Gibbs free energy change decreases as temperature increases, hence it is more likely to be negativeat higher temperatures. Thus, reaction A is more likely to become spontaneous at a higher temperature. (If the entropy change is negative — reaction B — then Gibbs free energy change increases as temperature increases, making it less likely the reaction will be spontaneous.)

3 **a** $\Delta H^{\ominus} = \{\Delta H^{\ominus}{}_f(H_2O,g) + (\Delta H^{\ominus}{}_f(CuO,s)\} - \Delta H^{\ominus}{}_f(Cu(OH)_2\,s)$

$\Delta H^{\ominus} = (-242 + -157) - (-450) = +51$ kJ mol^{-}

$\Delta S^{\ominus} = \{S^{\ominus}(H_2O) + S^{\ominus}(CuO)\} - S^{\ominus}(Cu(OH)_2 = 189 + 43 - 108 = +124$ J^{-} K^{-} mol^{-}

300K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 51 - (300 \times 0124) = +138$ kJ mol^{-}

700K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 51 - (700 \times 0124) = -358$ kJ mol^{-}

b The reaction will occur as written for the reaction at 700 K only (Gibbs free energy change value is negative).

4 $\Delta H^{\ominus} = \{\Delta H^{\ominus}{}_f(H_2O,g) + \Delta H^{\ominus}{}_f(Na_2O,s)\} - 2\,\Delta H^{\ominus}{}_f(NaOH,s)$

$\Delta H^{\ominus} = -242 + -414 - (2 \times -425) = +194$ kJ mol^{-}

$\Delta S^{\ominus} = \{S^{\ominus}(H_2O) + S^{\ominus}(Na_2O)\} - 2S^{\ominus}(NaOH) = 189 + 75 - (2 \times 64) = +136$ J^{-} K^{-} mol^{-}

500K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = +194 - (500 \times 0136) = +126$ kJ mol^{-}

1000K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = +194 - (1000 \times 0136) = +58$ kJ mol^{-}

It will not decompose at either of those temperatures as the Gibbs freeenergy change values are positive.

Chapter review questions

1 Gases will mix and never un-mix into their different gases. Different liquids will mix if added, but will never un-mix of their own accord.

2 Entropy is a measure of the randomness of a substance. Enthalpy measures the energy in a chemical reaction.

3 $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus}$

4 Means the reaction occurs under standard conditions: pressureof 100.0 kPa, concentration of 1.00 mol L^{-}, and at the specified temperature (often 298 K).

5 If the Gibbs free energy value ($\Delta G^{\ominus}$) is large and negative, then a reactionwill be spontaneous.

6 That the enthalpy and entropy of a system are not substantially different with temperature change.

7 Enthalpy of formation is defined as the enthalpy change when a substance forms from its elements in their standard state. Thus, elements in the standard state cannot have an enthalpy of formation. All substances, including elements in their standard state, have a level of disorder or chaos that is measured by its entropy.

8 Some reactions will go in the direction as written, but only when heat is continuously applied, or it is ignited. Thus, they are not truly spontaneous as they will not occur on their own.

9 a $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = -20 - (300 \times 0080) = -44$ kJ mol^{-} $\rightarrow$ spontaneous

ii $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = 20 - (300 \times 0080) = -4$ kJ mol^{-} $\rightarrow$ spontaneous (just)

iii $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = -20 - (300 \times -0080) = 4$ kJ mol^{-} $\rightarrow$ non-spontaneous

v $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = 20 - (300 \times -0080) = 44$ kJ mol^{-} $\rightarrow$ non-spontaneous

b Reactions ii and iii are very close to being at 0, hence with small chaes, could easily change.

c If the change in entropy is positive, the value for Gibbs free energy change decreases as temperature increases, hence it is more likely to be negativeat higher temperatures. Thus, neither of the two non-spontaneous reactions will work as neither hve a positive entropy value.

10 a $NH_4Cl(s) \rightarrow NH_3(g) + HCl(g)$

$\Delta H^\ominus = \{ \Delta H^\ominus{}_f(HCl,g) + \Delta H^\ominus{}_f(NH_3,g)\} - \Delta H^\ominus{}_f(NH_4C,s = (-92 + -46) - (-314) = +176$ kJ mol^{-}

$\Delta S^\ominus = \{S^\ominus(HCl) + S^\ominus(NH_3)\} - S^\ominus(NH_4Cl) = 1868 + 1923 - 946 = 2845$ J^{-} K^{-} mol^{-}

500K $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = 176 + (500 \times 02845) = +318$ kJ mol^{-}

1000K $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = 176 + (1000 \times 02845) = +461$ kJ mol^{-}

b Both values are positive, thus neither reaction is spontaneous.

c Yes

d The reaction when reversed would have a negative enthalpy and negative entropy value, so the Gibbs free energy change value would be negative as well, thus the reaction would be spontaneous.

11 a $\Delta H^\ominus = \Delta H^\ominus{}_f(N_2O_4\,g) - 2\,\Delta H^\ominus{}_f(NO_2\,g) = 9 - (2 \times 33) = -57$ kJ mol^{-}

$\Delta S^\ominus = S^\ominus(N_2O_4) - 2S^\ominus(NO_2) = 3042 - (2 \times 240) = -176$ J^{-} K^{-} mol^{-}

b 298K $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = -57 - (298 \times -0176) = +127$ kJ mol^{-}

700K $\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus = -57 - (700 \times -0176) = +662$ kJ mol^{-}

c As both values are positive, neither reactin will be spontaneous.

d The reaction when reversed would have a negative enthalpy and negative entropy value, so the Gibbs free energy change value would be negative as well, thus the reaction would be spontaneous.

12 a

Reaction	I	II	III	IV
$\Delta H^\ominus$ (kJ mol^{-})	–50	–50	50	50
$\Delta S^\ominus$ (J K^{-} mol^{-})	100	–100	100	–100
Direction of energy drive	R	R	F	F
Direction of entropy drive	F	R	F	R
$\Delta G^\ominus{}_{300}$	–80	–20	20	80
$\Delta G^\ominus{}_{600}$	–110	10	–10	110
Direction of reaction at 300K	F	F	R	R
Direction of reaction at 600K	F	R	R	R

b II and III

c Statement B is true.

13 a False – the entropy is more important in determiing the spontaneity.

b True

c False

14 a $\Delta H^{\ominus} = \{\Delta H^{\ominus}{}_f(\text{Ag}^+\text{aq}) + \Delta H^{\ominus}{}_f(\text{Fe}^{2+},\text{aq})\} - \Delta H^{\ominus}{}_f(\text{Fe}^{3+}\text{aq}) = (106 + -89) - (-49) = +66\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(\text{Ag}^+) + S^{\ominus}(\text{Fe}^{2+})\} - \{S^{\ominus}(\text{Ag}) + S^{\ominus}(\text{Fe}^{3+})\} = (727 + -1376) - (4255 + -3159) = +2085\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

298K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 66 - (298 \times 02085) = +387\ \text{kJ mol}^{-}$

340K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 66 - (340 \times 02085) = -489\ \text{kJ mol}^{-}$

b At 298K it will go in the reverse direction as t e value is positive. At 340 K the value is negative so it will go in the forward directon, as written.

c At 298K the value is cloe to 0.0 kJ mol $^{-}$, so changes in ion concentration would change the values involved and could change the reactio to being spontaneous.

15 At 1200K the reaction requires sustaied heat to occur. As such, while it does occur as written, it is not a spontaneou reaction.

16 a $\text{MgO(s)} + \text{C(s)} \rightarrow \text{Mg(s)} + \text{CO(g)}$

$\text{PbO(s)} + \text{C(s)} \rightarrow \text{Pb(s)} + \text{CO(g)}$

b MgO

$\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(\text{C,g}) - \Delta H^{\ominus}{}_f(\text{MgO,s}) = -111 - (-602) = +491\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(\text{CO}) + S^{\ominus}(\text{Mg})\} - \{S^{\ominus}(\text{C}) + S^{\ominus}(\text{MgO})\} = (198 + 33) - (6 + 27) = +198\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

300K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 491 - (300 \times 0198) = +4316\ \text{kJ mol}^{-}$

1000K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 491 - (1000 \times 0198) = +293\ \text{kJ mol}^{-}$

1500K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 491 - (1500 \times 0198) = +194\ \text{kJ mol}^{-}$

2000K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 491 - (2000 \times 0198) = +95\ \text{kJ mol}^{-}$

PbO

$\Delta H^{\ominus} = \Delta H^{\ominus}{}_f(\text{C,g}) - \Delta H^{\ominus}{}_f(\text{PbO,s}) = -111 - (-217) = +106\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(\text{CO}) + S^{\ominus}(\text{Pb})\} - \{S^{\ominus}(\text{C}) + S^{\ominus}(\text{PbO})\} = (198 + 65) - (6 + 69) = +188\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

300K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 106 - (300 \times 0188) = +496\ \text{kJ mol}^{-}$

1000K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 106 - (1000 \times 0188) = -82\ \text{kJ mol}^{-}$

c It is not practical at all to extract magnesium as it would take extremely high temperatures, but it would be practical for this to occur with lead as only 1000K is reqired to extract the lead.

17 a $\Delta H^{\ominus} = \{3\,\Delta H^{\ominus}{}_f(\text{SO}_2\,\text{g}) + 2\,\Delta H^{\ominus}{}_f(\text{H}_2\text{O,g})\} - 2(\Delta H^{\ominus}{}_f(\text{H}_2\text{SO}_4\,\text{l})$

$\Delta H^{\ominus} = (3 \times -297 + 2 \times -242) - 2 \times (-814) = +253\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{2S^{\ominus}(\text{H}_2\text{O}) + 3S^{\ominus}(\text{SO}_2)\} - \{S^{\ominus}(\text{S}) + 2S^{\ominus}(\text{H}_2\text{SO}_4)\} = (2 \times 1887 + 3 \times 248) - (31.8 + 2 \times 1569) = +7758\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

300K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 253 - (300 \times 07758) = -29006\ \text{kJ mol}^{-}$

400K $\Delta G = \Delta H^{\ominus} - T\Delta S^{\ominus} = 253 - (400 \times 07758) = -5732\ \text{kJ mol}^{-}$

b 300K would be a better temperature as the value is greater, thus the reaction will tend more to completion. Also, at a lower temperature, less heat is needed to run the reaction which makes the reaction easier to undertake.

18 a $\Delta H^{\ominus} = 2\,\Delta H^{\ominus}{}_f(CO_2\,g) - \{2\,\Delta H^{\ominus}{}_f(N{,}g) \quad + 2(\,\Delta H^{\ominus}{}_f(C{,}g)\}$

$\Delta H^{\ominus} = (2 \times -393) - (2 \times 90 + 2 \times -111) = -744$ kJ mol^{-}

$\Delta S^{\ominus} = \{S^{\ominus}(N_2) + 2S^{\ominus}(CO_2)\} - \{2S^{\ominus}(NO) + 2S^{\ominus}(CO)\} = (192 + 2 \times 214) - (2 \times 211 + 2 \times 198) = -198$ J^{-} K^{-} mol^{-}

300K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = -744 - (300 \times -0198) = -685$ kJ mol^{-}

b 1000K $\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = -744 - (1000 \times -0198) = -546$ kJ mol^{-} Thus, the reaction is still spontaneous.

c To lower the activation energy of the reaction and increase the reaction rate.

End-of-module review

1 $NaCl(aq) + H_2O(l) \rightarrow NaOH(aq) + HCl(aq)$, $\Delta H = +55$ kJ mol^{-}

$NaOH(s) + HCl(aq) \rightarrow NaCl(aq) + H_2O(l)$, $\Delta H = -98$ kJ mol^{-}

Adding the reactions together cancels the sodium chloride hydrochloric acid and water.

$\Delta H = +55 + -98 = -43$ kJ mol^{-}

2 Check with your teacher for the exact details of the experiment you performed and the chemicals you used, but it should be based on Investigation 14.1 and will involve measuring initial and final temperatures of chemical reactions.

3 $\Delta H^{\ominus} = \{4\,\Delta H^{\ominus}{}_f(N{,}g) \quad + 6\,\Delta H^{\ominus}{}_f(H_2O{,}g)\} - 4\,\Delta H^{\ominus}{}_f(NH_3\,g)$

$\Delta H^{\ominus} = (4 \times 90 + 6 \times -242) - (4 \times -46) = -908$ kJ mol^{-}

4 a $C_4H_{10}(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

b $n = \dfrac{m}{MM} = \dfrac{3\ 0}{(4 \times 12\,01) + (10 \times 1\,008)} = 005162$ mol

$q = \Delta H \times n = 287 \ \times 10^3 \times 005162 \ = 148.1$ kJ $= 1.481 \times 10^5$ J

$q = mc\Delta T$

$1481 \ \times 10^5 = 1000 \times 4.18 \times (T - 200)$

$T - 200 \ = 3544$

$T = 554$°C

5 Broken: 1 $\times$ C—H, 1 $\times$ Br—Br

Formed: 1 $\times$ C—Br, 1 $\times$ H—Br

6 a $Li(s) + H_2O(l) \rightarrow LiOH + \frac{1}{2}H_2(g)$, $\Delta H = -223$ kJ mol^{-}

$LiOH(s) + H_2(g) \rightarrow H_2O(l) + LiH(s)$, $\Delta H = +133$ kJ mol^{-}

Adding together and cancelling $\frac{1}{2}$ hydrogen, water and lithium hydroxide leaves the desired reaction.

Diagram should be modelled on Figure 15.3 (page 349) showing the two alternate pathways described above.

b $\Delta H = -233 + 133 = -100\ \text{kJ mol}^{-}$

c Enthalpy of formation

7 **a** $\Delta H^{\ominus} = \Delta H^{\ominus}_{f}(C,g) \quad - \Delta H^{\ominus}_{f}(H_2O,g) = (-111) - (-242) = +131\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(CO) + S^{\ominus}(H_2)\} - \{S^{\ominus}(C) + S^{\ominus}(H_2O)\} = (198 + 131) - (6 + 189) = +134\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

b $500K\ \ \Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 131 - (500 \times 0134) \ = +64\ \text{kJ mol}^{-}$

ii $1200K\ \ \Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 131 - (1200 \times 0134) \ = -298\ \text{kJ mol}^{-}$

c The reaction will only occur when the Gibbs free energy value is negative, thus will occur at 1200K but not at 50K.

8 **a** $q = mc\Delta T = 575 \times 4.18 \times (31.2 - 18.6) = 30284\ \text{J} = 30284\ \text{kJ}$

$$n = \frac{m}{MM} = \frac{1\ 65}{(2 \times 12\ 01) + (6 \times 1\ 008) + 16\ 00} = 00358\ \text{mol}$$

$$\Delta H = \frac{30\ 284}{0\ 0358} = -846\ \text{kJ mol}^{-}$$

b Heat loss to the surroundings (air, equipment). Incomplete combustion of the fuel produces less fuel per gram than complete combustion.

9 **a** $NiO(s) + C(s) \rightarrow Ni(s) + CO(g)$

$\Delta H^{\ominus} = \Delta H^{\ominus}_{f}(C,g) \quad - \Delta H^{\ominus}_{f}(Ni,s) \quad = -111 - (-240) = +129\ \text{kJ mol}^{-}$

$\Delta S^{\ominus} = \{S^{\ominus}(CO) + S^{\ominus}(Ni)\} - \{S^{\ominus}(C) + S^{\ominus}(NiO)\} = (198 + 30) - (6 + 38) = +272\ \text{J}^{-}\ \text{K}^{-}\ \text{mol}^{-}$

$\Delta G^{\ominus} = \Delta H^{\ominus} - T\Delta S^{\ominus} = 129 - (300 \times 0272) \ = +474\ \text{kJ mol}^{-}$

As the value is positive, the reactionis non-spontaneous.

b $0 = 129 - (T \times 0272)$

$0272\ T = 129$

$T = 474K$

Any temperature above 474K would be needed.

10 $\frac{1}{2}N_2(g) + O_2(g) \rightarrow NO_2(g),\ \Delta H = +33\ \text{kJ mol}^{-}$

$\Delta H \ = \text{negative formation of } NO_2 = -\ 33\ \text{kJ mol}^{-}$

$\Delta H_2 = \text{heat of atomisation of nitrogen} = +473\ \text{kJ mol}^{-}$

$\Delta H_3 = 2 \times \text{heat of atomisation of oxygen} = 2 \times 249 = +498\ \text{kJ mol}^{-}$

$\Delta H_4 = 2 \times B_{NO}$

$\Delta H_4 = \Delta H \ + \Delta H_2 + H_3 = -33 + 473 + 498 = 938\ \text{kJ mol}^{-}$

$$B_{NO} = \frac{938}{2} = 469\ \text{kJ mol}^{-}$$