

Chemical reactions

Mark scheme and worked solutions for 30 questions: 15 multiple choice, 15 structured.

1 Core Physical and chemical changes [1]

Answer: A

Worked solution

A chemical change always produces one or more new substances with different properties, and is generally difficult to reverse. Melting, dissolving and evaporating do not change the chemical identity of the substance involved — wax is still wax whether solid or liquid, sugar is still sugar whether dissolved or not, and ethanol is still ethanol whether liquid or vapour. Rusting, however, converts iron and oxygen (with water) into a new compound, iron oxide, so it is chemical.

2 Core Rate of reaction [1]

Answer: A

Worked solution

One of the standard rate factors on this syllabus is temperature: raising the temperature of the reactants always speeds up a reaction, all else being equal. This is true whatever the reaction, so as the marble chips and acid are heated, the same amount of product forms in less time.

3 Core Rate of reaction [1]

Answer: B

Worked solution

A catalyst provides an easier route for the reaction to follow, so more particles react successfully per second, but the catalyst itself takes part in intermediate steps and is regenerated by the end of the reaction — its mass and chemical identity are unchanged.

4**Core**

Rate of reaction

[1]

Answer: B

Worked solution

To follow "volume of gas produced over time" you need equipment that actually collects and measures gas volume — a gas syringe (or, alternatively, a measuring cylinder inverted over water) does this directly. The balance method is a valid way to follow this reaction's rate, but it measures mass lost, not gas volume, so it does not fit what the question asks for.

5**Supplement**

Rate of reaction

[1]

Answer: C

Worked solution

Only collisions with energy greater than or equal to the activation energy lead to a reaction. By lowering the activation energy required, a catalyst means a larger proportion of the existing collisions now have enough energy to succeed, without needing to change the particles' concentration, energy or the system's pressure at all.

Mark scheme

(a)

[2]

- volume/concentration of acid (1)
- mass/size/number of marble chips, OR temperature (1)

ALLOW any two of: concentration of acid, volume of acid, mass of marble chips, size of marble chips, temperature

REJECT "time" (this is the measured variable, not a controlled one)

(b)

[2]

- $0.56 \div 60$ (1)
- $= 0.0093$ (g/s) (allow 0.009–0.01 g/s) (1)

ALLOW correct working shown even if final rounding differs slightly, e.g. 9.3×10^{-3} g/s

(c)

[2]

- rate is fastest at the start and gradually decreases / slows down over time (1)
- by 150 s (and 180 s) the mass has stopped changing, showing the reaction has finished / one reactant (the acid) has been used up (1)

ALLOW "the reaction has stopped" for the second point

REJECT any statement that the reaction speeds up over time

Worked solution

(a) A fair test keeps everything the same except the one variable being investigated. Anything else that could affect the rate — the concentration and volume of the acid, the mass and size of the marble chips, and the temperature — must be kept the same, so any two of these are acceptable.

(b) Mean rate = mass lost $\div$ time taken. Between 0 s and 60 s, 0.56 g of mass was lost, so the mean rate = $0.56 \div 60 = 0.0093$ g/s (to 2 s.f.).

(c) Reading down the table, the mass lost per 30 s interval shrinks steadily (0.32, 0.24, 0.16, 0.10, 0.06, 0.00 g), showing the reaction is fastest at the start (when acid concentration is highest) and slows as acid is used up. Between 150 s and 180 s the mass stays at 0.88 g, so no more gas is being produced — the reaction has finished, most likely because all the acid has reacted (the marble chips were in excess).

Mark scheme

(a)

[3]

- increasing concentration increases the number of acid particles per unit volume (1)
- so collisions between acid and magnesium particles happen more frequently (1)
- more frequent collisions per second means more successful collisions (with energy $\geq$ activation energy) per second, so the rate increases (1)

ALLOW reference to "more particles in the same volume" for the first point

(b)

[3]

- increasing temperature increases the average kinetic energy of the particles, so they move faster (1)
- particles collide more frequently AND with more energy per collision (1)
- a greater proportion of collisions now have energy $\geq$ activation energy, so a greater proportion of collisions are successful, increasing the rate (1)

ALLOW "more particles have energy greater than or equal to the activation energy"

Worked solution

(a) Collision theory says that reactions only happen when particles collide with enough energy. Increasing the concentration of the acid packs more acid particles into the same volume of solution. With more particles present, the magnesium surface is struck by acid particles more often each second, so more collisions — and therefore more successful collisions — occur per second, increasing the rate.

(b) Raising the temperature gives all the particles more kinetic energy on average, so they move around faster. This has two effects: particles collide more often (because they are moving faster), and, more importantly, a much larger fraction of those collisions now have at least the activation energy needed to react. Both effects increase the rate, but the increase in successful collisions from higher energy is the dominant one.

Mark scheme

(a) [1]

- the rate of reaction / how the rate changes over time OR the volume of gas produced by a given time (1)

ALLOW "the amount of product formed over time"

(b) [2]

- advantage: Method 1 is simple and does not require a bung/airtight seal, so gas can escape past a cotton wool plug without any leaks affecting the reading (1)
- disadvantage: the mass changes are small compared with the total mass of the flask and contents, so a sensitive (2 or 3 d.p.) balance is needed to detect them accurately / acid spray can escape from the flask and be counted as mass lost (1)

ALLOW any other valid, specific comparison, e.g. Method 2 gives a larger, easier-to-read volume reading but requires an airtight, leak-free seal and a gas syringe has a limited capacity (typically 100 cm³)

ALLOW Method 1 is suitable for this reaction because carbon dioxide is a relatively dense gas ($M_r = 44$), so the mass loss is large enough to measure, as an advantage

(c) [2]

- use a more sensitive/higher-precision balance (that reads to more decimal places) (1)
- with a justification, e.g. so smaller changes in mass can be detected/measured more precisely (1)

ALLOW "take readings more frequently" or "repeat and average the results" with a suitable justification

ALLOW plug the neck of the flask with cotton wool (1) so acid spray cannot escape while the gas can, so only the gas loss is measured (1)

Worked solution

(a) Both a gas syringe and a mass-loss balance let you record how much product (gas) has formed at set time intervals, so both can be used to work out the rate of reaction at any point, or the mean rate over the whole reaction.

(b) Each method has trade-offs. The mass-loss method is easy to set up (just a flask, cotton wool and a balance) and doesn't need a perfectly airtight seal. It works for this reaction because carbon dioxide is a relatively dense gas ($M_r = 44$), so enough mass is lost to measure. However, the mass lost is still small compared with the mass of the flask and its contents, so you need a balance sensitive enough (reading to 0.01 g or 0.001 g) to detect small changes reliably, and any acid spray that escapes is wrongly counted as gas lost. The gas syringe method measures the volume of gas directly, which changes by a large, easy-to-read amount, but the flask and syringe connection must be completely airtight or gas will escape unmeasured, and a syringe can only hold a limited volume (usually 100 cm³).

(c) To improve Method 1, use a balance with a higher precision (for example, one that reads to 0.001 g rather than 0.01 g), so that the small mass changes between readings can be measured more reliably,

reducing the effect of the balance's own reading error on the results.

9 Core Reversible reactions and equilibrium

[1]

Answer: A

Worked solution

Reversible reactions, where products can react to re-form the original reactants under suitable conditions, are shown using the double-headed arrow $\rightleftharpoons$, as in $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$. A single arrow $\rightarrow$ is reserved for a reaction that only goes to completion in one direction.

10 Core Reversible reactions and equilibrium

[1]

Answer: C

Worked solution

Anhydrous copper(II) sulfate is white. When water is added, water molecules bond into the crystal structure to form hydrated copper(II) sulfate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, which is blue. This clear white-to-blue colour change is why anhydrous copper(II) sulfate is used as a simple chemical test for the presence of water.

11 Supplement Reversible reactions and equilibrium

[1]

Answer: C

Worked solution

Check each option by counting atoms: $\text{N}_2 + 3\text{H}_2 \rightleftharpoons 2\text{NH}_3$ gives 2 N atoms and 6 H atoms on each side, so it is balanced. Option A (1:1:1) and option B (2:3:2) do not balance both nitrogen and hydrogen at once. Option D has correct numbers but the wrong arrow — the Haber process is a reversible reaction reaching equilibrium, not a reaction that runs only one way, so $\rightleftharpoons$ must be used.

12 Supplement Reversible reactions and equilibrium

[1]

Answer: B

Worked solution

It helps to keep the two industrial processes separate: Haber process = 450 °C, 20 000 kPa, iron catalyst, producing ammonia; Contact process = 450 °C, 200 kPa, vanadium(V) oxide catalyst, producing sulfur trioxide. Option A swaps in the Contact process's pressure and catalyst, a common mix-up, so B is the only fully correct match for the Haber process.

Answer: D

Worked solution

The Contact process needs a supply of sulfur dioxide gas, which is obtained industrially either by burning naturally occurring sulfur in air, or by roasting (heating strongly in air) sulfide ores. The other three processes listed are real industrial processes, but each produces different products (hydrocarbons and other fractions; nitrogen, oxygen and noble gases; and chlorine, hydrogen and sodium hydroxide, respectively), none of which is sulfur dioxide.

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Mark scheme

(a) [2]

- the rate of the forward reaction equals the rate of the reverse reaction (1)
- the concentrations of reactants and products are no longer changing (1)

(b) [2]

- position of equilibrium shifts towards the products / more NH_3 is formed (1)
- because there are fewer gas molecules on the product side (2 mol) than the reactant side (4 mol), so increasing pressure favours the side with fewer molecules (1)

ALLOW "shifts to reduce the pressure/number of gas molecules"

(c) [2]

- position of equilibrium shifts towards the reactants / less NH_3 is formed (1)
- because the forward reaction is exothermic, so increasing temperature favours the endothermic (reverse) reaction, which absorbs the extra heat (1)

ALLOW "shifts in the endothermic direction to oppose the change"

(d) [1]

- no effect on the position of equilibrium (1)

REJECT any answer suggesting the catalyst shifts the equilibrium; **ALLOW** a note that it reaches equilibrium faster, but the position (yield) is unchanged

Worked solution

(a) At equilibrium in a closed system, the forward and reverse reactions are still both happening, but at equal rates, so there is no further overall change in the amounts of reactants and products present — their concentrations stay constant.

(b) There are 4 moles of gas on the left ($1 \text{ N}_2 + 3 \text{ H}_2$) but only 2 moles of gas on the right (2 NH_3). Increasing the pressure favours whichever side has fewer gas molecules, since that reduces the total pressure again, so the equilibrium shifts towards the ammonia (product) side, increasing the yield of NH_3 .

(c) The forward (ammonia-forming) reaction releases heat (exothermic). By Le Chatelier's principle, raising the temperature favours the reaction that absorbs heat — the reverse, endothermic reaction — so the equilibrium shifts back towards N_2 and H_2 , reducing the yield of ammonia. (This is why industrially a compromise, moderate temperature is chosen rather than a very high one.)

(d) A catalyst speeds up both the forward and reverse reactions equally, so equilibrium is reached faster, but the position of equilibrium — and therefore the equilibrium yield of ammonia — is completely unaffected.

Mark scheme

(a) [2]

- $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ (1)
- correct state symbols as shown (1)

ALLOW correctly balanced equation with $\rightleftharpoons$; **REJECT**: a one-way arrow $\rightarrow$, or missing/incorrect state symbols losing the second mark only

(b) [2]

- hydrogen: from methane / natural gas (1)
- nitrogen: from the air (1)

(c) [2]

- 450 °C (1)
- 20 000 kPa (or 200 atm) and iron catalyst (1)

ALLOW either order; **ALLOW** small rounding of the pressure value if units are correct

Worked solution

(a) The Haber process combines nitrogen and hydrogen gas in a reversible reaction to form ammonia gas: $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$. Remember the double arrow, since the reaction does not go to completion.

(b) The nitrogen used is extracted from the air (which is about 78% nitrogen), while the hydrogen is obtained from methane (natural gas), typically via reaction with steam.

(c) Industrially, a compromise temperature of 450 °C, a high pressure of 20 000 kPa (about 200 atmospheres), and an iron catalyst (to speed up the reaction without needing to alter temperature or pressure further) are used.

Mark scheme

(a) [2]

- $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$ (1)
- correct state symbols as shown (1)

ALLOW correctly balanced equation with $\rightleftharpoons$

(b) [2]

- sulfur dioxide: from burning sulfur or roasting sulfide ores (1)
- oxygen: from the air (1)

(c) [2]

- 450 °C (1)
- 200 kPa (or 2 atm) and vanadium(V) oxide catalyst (1)

Worked solution

(a) In the key stage of the Contact process, sulfur dioxide reacts reversibly with oxygen to form sulfur trioxide: $2\text{SO}_2(\text{g}) + \text{O}_2(\text{g}) \rightleftharpoons 2\text{SO}_3(\text{g})$.

(b) The sulfur dioxide comes from burning sulfur, or from roasting sulfide ores in air, while the oxygen needed is simply taken from the air.

(c) This step is carried out at about 450 °C and 200 kPa (about 2 atmospheres), using a vanadium(V) oxide catalyst to increase the rate of reaction.

Mark scheme

(a) [3]

- a lower temperature would give a higher equilibrium yield of ammonia, since the forward reaction is exothermic (1)
- but a lower temperature would make the rate of reaction much slower / take far too long to reach equilibrium (1)
- so 450 °C is a compromise that gives a reasonably fast rate while still keeping an acceptable (economically worthwhile) yield (1)

(b) [3]

- a higher pressure would shift the equilibrium further towards ammonia, giving a higher yield, since there are fewer gas molecules on the product side (1)
- but very high pressures need much stronger (thicker-walled) reaction vessels and pipework, which are far more expensive to build and maintain, and increase the risk of a dangerous leak or explosion (1)
- so 20 000 kPa is a compromise that gives a good yield without the excessive cost and safety risk of much higher pressures (1)

Worked solution

(a) A lower temperature would favour the exothermic forward reaction, increasing the equilibrium yield of ammonia. However, lowering the temperature also slows down the rate of reaction (fewer particles have enough energy to react), so it would take an uneconomically long time to produce useful amounts of ammonia. 450 °C is chosen as a compromise: high enough for a reasonably fast rate, while still giving an acceptable yield.

(b) Increasing pressure would push the equilibrium towards the side with fewer gas molecules (ammonia), increasing yield. But building and running plant that can safely withstand pressures much higher than 20 000 kPa is very expensive (thicker steel vessels and pipes) and increases the danger of a serious leak or explosion. 20 000 kPa is chosen as a balance between a good enough yield and manageable cost/safety.

Mark scheme

(a) [2]

- the reaction can go in both the forward and reverse directions (1)
- under suitable conditions, the products can react to re-form the reactants (1)

(b) [2]

- it turns pink (1)
- because adding water shifts the equilibrium towards the hydrated (pink) form / the reverse reaction (as written) is favoured, forming hydrated cobalt(II) chloride (1)

ALLOW reference to Le Chatelier's principle, "opposing the change by using up the added water"

(c) [2]

- it turns blue, and water vapour/steam is given off (1)
- because heating removes water (as a gas) and/or supplies energy, shifting the equilibrium towards the anhydrous (blue) form / the forward reaction (as written) is favoured, forming anhydrous cobalt(II) chloride (1)

ALLOW "heating drives off the water of crystallisation"

Worked solution

(a) A reversible reaction is one where, under the right conditions, the forward reaction can be reversed: the products can react together again to re-form the original reactants, instead of the reaction only ever going one way to completion.

(b) Adding water to the blue anhydrous cobalt(II) chloride pushes the equilibrium towards the side that uses up water — the hydrated, pink form. This is the reverse reaction as the equation is written, so the sample turns pink (and, because hydration is exothermic, the sample may also get warm).

(c) Heating the pink hydrated solid strongly drives off the water as steam (an endothermic process), shifting the equilibrium towards the anhydrous, blue form. This is the forward reaction as written, so the solid turns blue.

Answer: D

Worked solution

The oxidation numbers in a neutral compound must sum to zero. Oxygen is almost always -2 , so the three oxygens contribute -6 in total. To balance this, the two iron atoms must together contribute $+6$, so each iron atom has an oxidation number of $+3$ — hence the name iron(III) oxide.

20**Core**

Redox

[1]

Answer: A

Worked solution

At this level, oxidation is defined as gaining oxygen and reduction as losing oxygen. CuO loses its oxygen atom to become Cu, so CuO is reduced. At the same time, H₂ gains that oxygen to become H₂O, so H₂ is oxidised. Since both happen simultaneously, this is a redox reaction.

21**Core**

Redox

[1]

Answer: D

Worked solution

Track the oxidation numbers: magnesium goes from 0 (as an uncombined element) to +2 (as Mg²⁺), an increase, so it is oxidised. Copper goes from +2 (as Cu²⁺) to 0 (as Cu metal), a decrease, so it is reduced. The sulfate ion, SO₄²⁻, appears unchanged on both sides of the equation, so it takes no part in the redox process.

22**Supplement**

Redox

[1]

Answer: C

Worked solution

The oxidation numbers in an ion must sum to the ion's overall charge. In NO₃⁻, the three oxygens contribute 3 × (-2) = -6. For the total to equal the ion's charge of -1, nitrogen must be +5, since (+5) + (-6) = -1.

23**Supplement**

Redox

[1]

Answer: B

Worked solution

This is a classic test for a reducing agent. Acidified potassium manganate(VII) solution is deep purple. When it reacts with a reducing agent such as iodide ions, the purple manganate(VII) ions are reduced to pale pink (often described as almost colourless) manganese(II) ions, while the iodide ions are oxidised to iodine, which turns the solution brown. The purple colour fading, with a brown colour appearing, is the visible sign that a redox reaction has taken place.

Mark scheme

(a) [2]

- Fe_2O_3 (1)
- it loses oxygen, changing from Fe_2O_3 to Fe (1)

(b) [1]

- CO, because it gains oxygen, changing from CO to CO_2 (1)

ALLOW mark awarded for identifying CO with any valid reason based on oxygen

(c) [1]

- because oxidation (of CO) and reduction (of Fe_2O_3) are happening at the same time (1)

ALLOW "because it involves simultaneous oxidation and reduction"

Worked solution

(a) Fe_2O_3 loses its oxygen atoms, ending up as pure iron, Fe — a loss of oxygen means Fe_2O_3 is reduced.

(b) CO gains an oxygen atom to become CO_2 — a gain of oxygen means CO is oxidised.

(c) Because both of these changes (the loss of oxygen from Fe_2O_3 and the gain of oxygen by CO) happen in the same reaction at the same time, this fits the definition of a redox reaction: one where oxidation and reduction occur simultaneously.

Mark scheme

(a) [1]

- iron(II) oxide (1)

(b) [1]

- iron(III) oxide (1)

(c) [1]

- copper(I) oxide (1)

(d) [1]

- copper(II) oxide (1)

Worked solution

The Roman numeral in a metal's name shows its oxidation number in that compound. In FeO, oxygen is -2 and the compound is neutral, so iron must be $+2$: iron(II) oxide. In Fe₂O₃, the two irons together balance three oxygens (-6), so each iron is $+3$: iron(III) oxide. In Cu₂O, the two coppers together balance one oxygen (-2), so each copper is $+1$: copper(I) oxide. In CuO, one copper balances one oxygen, so copper is $+2$: copper(II) oxide.

Mark scheme

(a) [1]

- loss of electrons (1)

(b) [2]

- $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$ (2)

ALLOW $\text{Zn(s)} - 2\text{e}^- \rightarrow \text{Zn}^{2+}(\text{aq})$ (2); correct species but unbalanced electrons/charge (1)

(c) [2]

- reduced (1)
- because it gains electrons, changing from Cu^{2+} to Cu (1)

REJECT "oxidised" for the verdict, even with correct working shown elsewhere

Worked solution

(a) In terms of electrons, oxidation is defined as the loss of electrons by a species.

(b) Zinc atoms each lose two electrons to form Zn^{2+} ions: $\text{Zn(s)} \rightarrow \text{Zn}^{2+}(\text{aq}) + 2\text{e}^-$. This is balanced for both atoms and charge (0 on the left; +2 and -2, i.e. 0 overall, on the right).

(c) Copper(II) ions gain the electrons released by zinc, becoming copper atoms: $\text{Cu}^{2+}(\text{aq}) + 2\text{e}^- \rightarrow \text{Cu(s)}$. Since Cu^{2+} gains electrons, it is reduced (while the zinc, which loses electrons, is oxidised).

Mark scheme

(a) [1]

- -3 (1)

(b) [2]

- $+6$ (2)

ALLOW 1 mark for correct method ($2\text{Cr} + 7 \times (-2) = -2$) shown with an incorrect final value

(c) [2]

- $+4$ (2)

ALLOW 1 mark for correct method ($\text{S} + 3 \times (-2) = -2$) shown with an incorrect final value

(d) [1]

- $+7$ (1)

Worked solution

(a) In NH_4^+ , hydrogen is $+1$ ($\times 4 = +4$). For the ion's total charge of $+1$, nitrogen must be $+1 - 4 = -3$.

(b) In $\text{Cr}_2\text{O}_7^{2-}$, oxygen is -2 ($\times 7 = -14$). For the ion's total charge of -2 , the two chromiums together must total $-2 - (-14) = +12$, so each chromium is $+12 \div 2 = +6$.

(c) In SO_3^{2-} , oxygen is -2 ($\times 3 = -6$). For the ion's total charge of -2 , sulfur must be $-2 - (-6) = +4$.

(d) In MnO_4^- , oxygen is -2 ($\times 4 = -8$). For the ion's total charge of -1 , manganese must be $-1 - (-8) = +7$.

Mark scheme

(a) [1]

- purple to (almost) colourless / pale pink (1)

ALLOW purple to brown / purple fades and a brown colour develops (the brown is the iodine formed, and is what is seen with excess iodide)

ALLOW "purple fades" without stating a specific final colour, provided a fade/loss of colour is stated

(b) [2]

- oxidising agent: a substance that oxidises another substance and is itself reduced (1)

ALLOW a substance that accepts / gains electrons (from another substance)

- the oxidising agent here is the manganate(VII) ion / potassium manganate(VII) (1)

(c) [2]

- reducing agent: a substance that reduces another substance and is itself oxidised (1)

ALLOW a substance that donates / loses electrons (to another substance)

- the reducing agent here is the iodide ion / potassium iodide (1)

(d) [2]

- reduced (1)

- because the oxidation number of manganese decreases from +7 (in MnO_4^-) to +2 (in Mn^{2+}) (1)

ALLOW ECF if an incorrect oxidation number of Mn was given elsewhere but the direction of change (decrease) is correctly stated

Worked solution

(a) The purple colour of the acidified manganate(VII) ions fades as they are converted to pale pink/almost colourless manganese(II) ions. At the same time the iodide ions are oxidised to iodine, so the solution turns brown; either "purple to colourless" or "purple to brown" is accepted.

(b) An oxidising agent oxidises another substance while being reduced itself. Here, the manganate(VII) ion takes electrons from the iodide ions (oxidising them to iodine) while itself being reduced to Mn^{2+} , so potassium manganate(VII) is the oxidising agent.

(c) A reducing agent reduces another substance while being oxidised itself. Here, the iodide ion gives up electrons to the manganate(VII) ion (reducing it) while itself being oxidised to iodine, so potassium iodide is the reducing agent.

(d) Manganese's oxidation number falls from +7 in MnO_4^- to +2 in Mn^{2+} , a decrease, which is the definition of reduction — so the manganate(VII) ion is reduced (consistent with it being the oxidising agent in part (b)).

Mark scheme

(a)

[2]

- oxidising agent: a substance that oxidises another substance and is itself reduced (1)
- reducing agent: a substance that reduces another substance and is itself oxidised (1)

ALLOW oxidising agent accepts / gains electrons; reducing agent donates / loses electrons

(b)

[4]

- oxidising agent: Fe_2O_3 (1), because it is reduced / it loses oxygen, going from Fe_2O_3 to Fe (1)
- reducing agent: CO (1), because it is oxidised / it gains oxygen, going from CO to CO_2 (1)

ALLOW reasons expressed in terms of "causing" oxidation/reduction of the other substance, provided the direction is correct

Worked solution

(a) An oxidising agent oxidises another substance and is itself reduced in the process; a reducing agent reduces another substance and is itself oxidised in the process.

(b) Iron(III) oxide loses its oxygen (is reduced) as it is converted to iron, and in doing so it takes that oxygen from carbon monoxide (oxidising the CO to CO_2) — so Fe_2O_3 is the oxidising agent. Carbon monoxide gains oxygen (is oxidised) as it is converted to carbon dioxide, and in doing so it removes oxygen from the iron(III) oxide (reducing it to iron) — so CO is the reducing agent.

Mark scheme

(a) [2]

- before: +2 (1)
- after: +3 (1)

(b) [2]

- before: 0 (1)
- after: -1 (1)

(c) [2]

- Fe^{2+} is oxidised, because its oxidation number increases (+2 $\rightarrow$ +3) (1)
- Cl_2 is reduced, because its oxidation number decreases (0 $\rightarrow$ -1) (1)

ALLOW ECF from incorrect oxidation numbers in (a)/(b), provided the direction of change used is consistent with those values

(d) [2]

- oxidation: $\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{e}^-$ (1)
- reduction: $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$ (1)

ALLOW multiplying the oxidation equation by 2 to balance electrons overall, provided both half-equations are individually correctly balanced for atoms and charge

Worked solution

(a)/(b) Iron starts as Fe^{2+} (oxidation number +2, since it is a simple ion) and ends as Fe^{3+} (+3). Chlorine starts as Cl_2 , an uncombined element, so its oxidation number is 0; it ends as Cl^- ions, each with oxidation number -1.

(c) Since iron's oxidation number increases (a loss of electrons), Fe^{2+} is oxidised. Since chlorine's oxidation number decreases (a gain of electrons), Cl_2 is reduced. This reaction is a redox reaction because oxidation and reduction happen simultaneously.

(d) The oxidation half-equation shows iron(II) losing one electron: $\text{Fe}^{2+}(\text{aq}) \rightarrow \text{Fe}^{3+}(\text{aq}) + \text{e}^-$. The reduction half-equation shows each chlorine atom in Cl_2 gaining one electron (two electrons in total, since Cl_2 contains two atoms): $\text{Cl}_2(\text{g}) + 2\text{e}^- \rightarrow 2\text{Cl}^-(\text{aq})$. Combining twice the oxidation equation with the reduction equation cancels the electrons and reproduces the overall equation given in the question.