Buttershaw Business and Enterprise College

Chemistry Paper 1 Key Recall Facts

Atomic Structure and the Periodic Table, Bonding, Structure and Properties of Matter, Quantitative Chemistry, Chemical Changes and Energy Changes

"If I have seen further it is by standing on the shoulders of Giants," Sir Isaac Newton, 1675.*

Name……

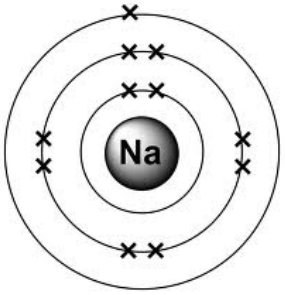
Group……

Teacher……

*Sir Isaac Newton developed the Universal Law of Gravitation, that states that gravity affects everything in the Universe, and the three Laws of Motion.

Atomic Structure and the Periodic Table

<table>
<tr><td>1. Name the 3 subatomic particles in the atom and state their location</td><td>• Protons found in the nucleus (centre of atom)
• Neutrons found in the nucleus (centre of atom)
• Electrons found on shells/energy levels around the nucleus</td></tr>
<tr><td>2. What is the relative charge, relative mass and symbol for:
Proton
Neutron
Electron</td><td><table><tr><th>Particle</th><th>Relative charge</th><th>Relative mass</th><th>Symbol</th></tr><tr><td>proton</td><td>+1</td><td>1</td><td>p</td></tr><tr><td>neutron</td><td>0</td><td>1</td><td>n</td></tr><tr><td>electron</td><td>-1</td><td>1/1836
(5.45×10^{-4})</td><td>e^-</td></tr></table></td></tr>
<tr><td>3. What are the rules for drawing electronic configuration?</td><td>1. Find out the number of electrons from periodic table (number of electrons is the same as the proton, or atomic number)
2. Max of 2 electrons in first shell
3. Max of 8 electrons in all other shells

Eg, Sodium:
  </td></tr>
<tr><td>4. What is the radius of an atom?</td><td>0.1nm (10^{-10}m)</td></tr>
<tr><td>5. Size of the nucleus of an atom?</td><td>(10^{-14}m)</td></tr>
<tr><td>6. Why do atoms have no overall charge?</td><td>Equal number of positive protons and negative electrons.</td></tr>
<tr><td>7. How does an atom form an ion?</td><td>The atom has gained or lost electrons to form a full, stable outer shell.
If the atom loses electrons, it will become a positive ion and if it gains electrons, it will become a negative ion.</td></tr>
</table>

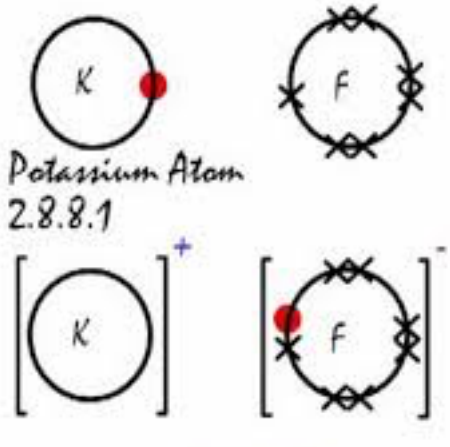
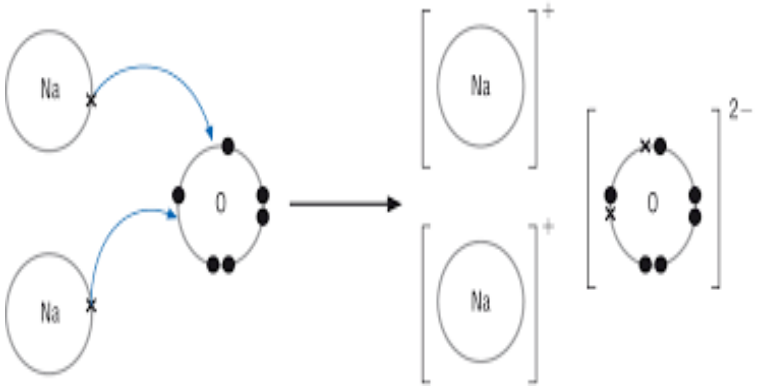
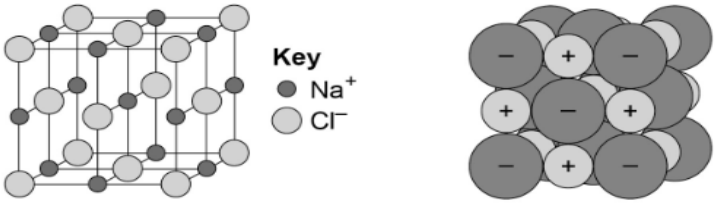
8. What is an element?	An element only contains one type of atom.
9. What is a compound?	A compound contains 2 or more elements chemically combined together.
10. What is a mixture?	2 or more substances (elements or compounds) not chemically combined together.
11. How do we separate mixtures?	• Filtration (separates an insoluble solid from a mixture with a liquid) • Evaporation/crystallisation (soluble solid from a solution) • Distillation (2 solvents) • chromatography (coloured compounds/dyes).
12. Name the Scientists in order of the discoveries made for the history of the atom. What did they discover?	• Dalton 1803 – Theory that all substances made of atoms and atoms are indivisible (spherical model) • JJ Thompson 1897 – Plum pudding model after discovering the electron. • Rutherford 1907 – Alpha scattering experiment that disproved the plum pudding model. • Neils Bohr 1913 – Idea of electrons in energy levels around nucleus. Bohr model. • Chadwick 1932 – Discovered the neutron
13. How did the alpha scattering experiment disprove the plum pudding model?	• Most of the alpha particles went straight through – this meant that most of the atom was empty space. • Some alpha particles deflected by a big angle – this meant large mass/positive charge concentrated in the nucleus. • Only a very few alpha particles deflected by a big angle – this meant nucleus is very small.
14. Compare the plum pudding and the Bohr model	In the plum pudding model, the protons are not subatomic particles but in a ball of positive charge. WHEREAS Nuclear model the protons are in the nucleus. Plum pudding electrons are embedded in the ball of positive charge. WHEREAS Nuclear model the electrons are in shells or energy levels.
15. What is mass number?	Number of protons + number of neutrons

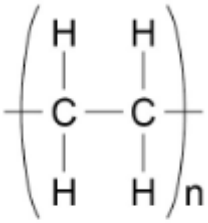
16. What is atomic number?	Number of protons
17. How would you work out the number of protons or electrons of an atom?	Look at the atomic number (this is similar for all atoms of a particular element)
18. How would you work out the number of neutrons of an atom?	Mass number – atomic number
19. What is an isotope?	An isotope is an atom of the same element with the same number of protons but different number of neutrons.
20. What is the relative atomic mass of an element?	The relative atomic mass of an element is an average value that takes account of the abundance of the isotopes of the element.
21. What is a group on the periodic table?	Column containing elements with the same number of electrons in the outer shell and similar properties. The group number tells you the number of electrons in the outer shell.
22. What is a period on the periodic table?	A row with the period number being the number of electron shells.
23. Why is it called a periodic table?	Similar properties occur at regular (periodic) intervals.
24. How are the elements in the modern periodic table arranged?	By atomic number.
25. How were the elements in the Early Periodic tables arranged?	By atomic weight (protons were not discovered at the time)
26. Why was Newlands' Periodic table not accepted?	He ended up putting elements in groups that did not have similar properties e.g., iron oxygen and sulfur.
27. Describe 3 things that Mendeleev did to allow his periodic table to be accepted	1) He left gaps for undiscovered elements. 2) He predicted properties of these elements and when they were discovered, his predictions were correct. 3) He switched some elements around (I and Te) to make sure they were in groups with elements that had similar properties.
28. Give the reason why the order of atomic weight was not always correct	Existence of isotopes.

29. Where are metals and non-metals located on the periodic table?	Metals are found on the left-hand side and the centre. Non-metals found on the right-hand side.
30. Why do metals form positive ions and non-metals negative ions?	Metals have less than 4 electrons in the outer shell, so less energy needed to lose electrons, so form positive ions. Non-metals have more than 4 electrons in the outer shell, so less energy to gain electrons, so form negative ions.
31. Give properties of metals	• High melting point • Good conductors of heat and electricity • Malleable • Generally, have high density. • Sonorous (make ringing sound when hit) • Ductile (can be stretched into wires)
32. Give properties of non-metals	• Generally low melting points • Poor conductors of heat and electricity • Brittle • Generally, have low density
33. What is the other name for group 0 elements?	Noble gases
34. Why are group 0 elements unreactive?	They have a full outer shell, so do not need to gain, or lose electrons.
35. Why does the boiling point increase as you go down Group 0 or group 7?	• Down the group, atom gets bigger. • Stronger intermolecular forces. • More energy needed to break forces.
36. What is the other name for group 1 elements?	Alkali metals
37. Why are they called alkali metals?	React with water to form a metal hydroxide, which is an alkaline solution.
38. Describe the properties of Group 1 Alkali metals.	• Reactive • Silvery solids • Form white compounds (e.g., sodium chloride) • Shiny when cut. • Can be cut with a knife – soft. • Solid metals at RT – melting point decreases down group. • Tarnish (turn dull) when react with O_2, therefore stored in oil. • Form +1 ions – lose one electron to form full outer shell (more stable).

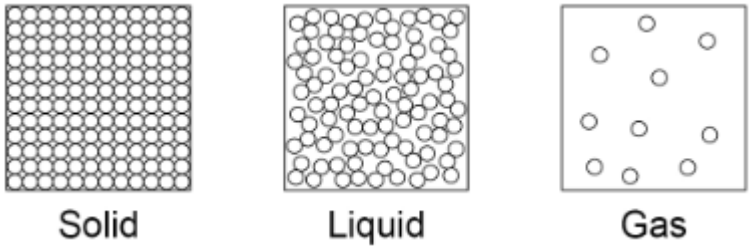
39. Why does reactivity increase down Group 1 (the alkali metals)?	• Going down group, there are more shells. • Distance between nucleus and outer electron is bigger. • Weaker attraction between nucleus and outer electron. • Less energy needed to lose electron. • Reactivity increases.
40. What products are made when a metal reacts with water?	Metal + water → Metal hydroxide + hydrogen Hydrogen gas is explosive.
41. What is the other name for group 7 elements?	Halogens
42. Why are chlorine, bromine, and iodine in the same group?	They have similar properties and have the same number of electrons in the outer shell.
43. Describe properties of the halogens (Group 7).	• Non-metals • Diatomic (2 atoms that are covalently bonded together) e.g., F_2, Cl_2, Br_2, I_2, At_2 • Form -1 ions – gain one electron to have full outer shell (stable) • Forms acidic hydrogen halides when reacted with hydrogen
44. Why does reactivity decrease down Group 7?	• Going down group, there are more shells. • Distance between nucleus and incoming electron is bigger. • Weaker attraction between nucleus and incoming electron. • More energy needed to gain electron. • Reactivity decreases.
45. Describe the properties of chlorine, bromine, and iodine.	Chlorine – pale yellow gas at RT (room temperature) Bromine – deep red liquid at RT but red-brown gas Iodine – grey solid at RT but purple gas

Bonding, Structure and Properties of Matter

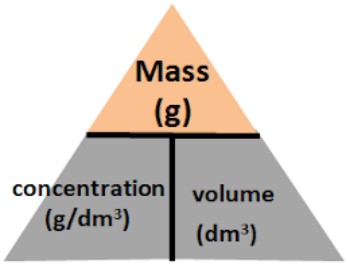
1. What is ionic bonding?	• Occurs between positive metal ion and negative non-metal ions. • Involves the transfer of electrons from the metal to the non-metal.
2. What is covalent bonding?	• Occurs between non-metal atoms. • Involves the sharing pairs of electrons held in place by strong attractions to the nucleus of atoms.
3. Why do atoms transfer or share electrons to form chemical bonds?	To gain a full outer shell of electrons, which is more stable.
4. Draw and explain the ionic dot and cross diagram for potassium fluoride.	 One potassium atom loses one electron to get a full outer shell and so becomes a positive (+1) ion. One fluorine atom gains that electron for a full outer shell and becomes a negative (-1) ion.
5. Draw and explain the ionic dot and cross for sodium oxide	 Two sodium atoms lose one electron each to form a +1 ion with a full outer shell. One oxygen atom gains both electrons to form a -2 ion with a full outer shell.
6. Describe the structure of an ionic compound.	A giant lattice with strong electrostatic forces of attraction between positive metal ions and negative non-metal ions. 

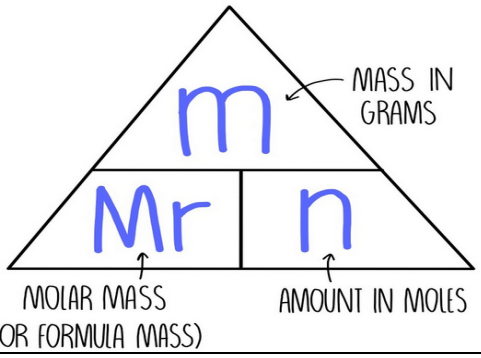
7. Why do ionic compounds have high melting points?	• Giant lattice • Strong electrostatic forces of attraction between oppositely charged ions. • Lots of energy needed to break forces.
8. Why do ionic compounds not conduct electricity when solid?	Ions are fixed and cannot move.
9. Why do ionic compounds conduct electricity when molten or dissolved in water?	Ionic bonds have broken, so ions are now free to move and carry charge through the liquid/solution.
10. Why do simple covalent molecules have low melting and boiling points?	• Small molecules • Weak intermolecular forces • Little energy needed to break forces
11. Why do simple covalent molecules not conduct electricity?	Molecules do not have an electrical charge (no delocalised electrons or ions)
12. Compare the strength of intermolecular forces and covalent bonds.	Covalent bonds are very strong but intermolecular forces are weak.
13. What are polymers?	Polymers are large molecules of monomers (small molecules) joined together in a polymerisation reaction.
14. Why do polymers have high melting points?	• The atoms are linked together by strong covalent bonds. • Long molecule, so strong intermolecular forces • Lots of energy needed to break forces.
15. Draw a diagram to show the representation of a polymer (polyethene)	 poly(ethene)
16. Give 3 examples of giant covalent structures.	Diamond, graphite and silica (silicon dioxide)
17. Which element makes up graphite and diamond?	Carbon
18. What are allotropes?	Made of the same element, but has a different structure/arrangment of atoms.
19. Why do giant covalent substances have high melting points?	• Giant lattice • Lots of strong covalent bonds • Lots of energy needed to break bonds

20. Explain why diamond is hard	Each carbon atom in diamond is connected to 4 other carbon atoms by interconnecting covalent bonds. Therefore, it has no layers which can slide over each other.
21. Explain why graphite is soft	Each carbon atom in graphite is connected to 3 other carbon atoms in layers of hexagonal rings. These layers can slide over each other as there are no covalent bonds between the layer, just weak intermolecular forces.
22. Explain why graphite conducts electricity but not diamond	Graphite has delocalised electrons that can move and carry charge through the whole structure. Diamond has no delocalised electrons.
23. What is graphene? Give uses of graphene.	Graphene is a single layer of graphite, which is used in electronics and composites.
24. Give 3 properties of graphene	• High melting point (same as diamond and graphite) • Conducts electricity (same as graphite) • Transparent (one layer thick) • Flexible (strong covalent bonds)
25. Describe the structure of fullerenes.	Allotrope of carbon with a hollow shape. The shape is based on hexagonal rings but can also contains rings of 5 or 7 carbon atoms.
26. What was the first fullerene discovered?	Buckminsterfullerene (C_{60}) which had a spherical structure
27. State one property of fullerenes and suggest 2 uses of fullerenes.	• Fullerenes have a large surface area. • They are useful as catalysts. • They can be used as lubricants.
28. Describe the structure of carbon nanotubes.	• Cylindrical fullerenes called 'buckytubes' • They have a very high length:diameter ratio • Tubes of graphene like sheets
29. What are the properties of carbon nanotubes?	• High tensile strength • Conducts electricity (and heat)
30. Give 3 uses of carbon nanotubes	• Nanotechnology • Electronics • Materials (tennis rackets)
31. What is metallic bonding?	Where positive metals ions are closely packed together with delocalised electrons flowing around them.
32. Why do metals conduct electricity?	Metals have delocalised electrons that can move and carry charge through the whole structure.

33. Why do metals conduct heat?	Delocalised electrons are free to move and transfer thermal energy.
34. Why do metals have high melting points?	• Giant lattices • Strong electrostatic forces between positive metal ions and elocalised electrons • Lots of energy needed to break forces
35. Why are pure metals malleable (soft)?	All atoms have the same size. Atoms are in layers. Atoms can slide over each other.
36. Why are alloys harder than pure metals?	Alloys contain *atoms* of different sizes. These different sizes distort the layered structure of the atoms in the alloy. Atoms cannot slide over each other.
37. Give the 4 state symbols	(s) solid (l) liquid (g) gas (aq) aqueous (dissolved in water)
38. Draw the particle model for solids, liquids, and gases	
39. Name the changes of state	Melting – solid → liquid Freezing – liquid → solid Boiling – liquid → gas Condensation – gas → liquid Sublimation – solid → gas
40. Give limitations of the particle model of matter	• Not all particles are spheres – some can be distorted by charged particles nearby. • Atoms are not solid – mostly empty space. • Does not show forces between the particles.

Quantitative Chemistry

1. What is the M*r* of water (H_2O)? A*r* – H = 1, O = 16	H – 2 x 1 = 2 O – 1 x 16 = 16 M*r* = 2 + 16 = 18
2. What is the M*r* of $KMnO_4$? A*r* – K = 38, O = 16, Mn = 55	K – 1 x 39 = 39 Mn – 1 x 55 = 55 O – 4 x 16 = 64 M*r* = 39 + 55 + 64 = 158
3. What is the M*r* of $Al_2(CO_3)_3$? A*r* – K = 38, O = 16, Mn = 55	Al – 2 x 27 = 54 C – 3 x 12 = 36 O – 9 x 16 = 144 M*r* = 54 + 36 + 144 = 234
4. Calculate the percentage by mass of iron in iron oxide (Fe_2O_3) A_r – Fe = 56, O = 16	Fe – 2 x 56 = 112 O – 3 x 16 = 48 M*r* = 112 + 48 = 160 % by mass of iron = (112/160) x 100 = 70%
5. What is the theory of the conservation of mass?	The law of conservation of mass states that no atoms are lost or made during a chemical reaction, so the mass of the products equals the mass of the reactants.
6. If there is a gaseous reactant, why does the mass appear to have changed?	Mass of reactant is not measured. Gas comes in from the atmosphere. Increases the mass.
7. If there is a gaseous product, why does the mass appear to have changed?	Products escapes into the atmosphere. Decreases the mass.
8. How do you calculate a mean of a set of values?	To work out the MEAN average: 1. Add all of your values for the results of the experiment together (excluding anomalous results if there are any) 2. Divide by the number of results you have.
9. How do you calculate the range of a set of values?	The range is the difference between the highest and lowest values.
10. How do you calculate the uncertainty of results from your experiment?	Range ÷ 2 Uncertainty = Mean ± (Range/2)
11. Write the concentration equation triangle	

12. **HT only** Define the term concentration	The mass of solute dissolved in a given voume of solvent. The more solute dissolved, the greater the concentration.
13. **HT only** What is a mole?	Unit of chemical measurement where 1 mole of a substance contains Avogadro's constant of particles.
14. **HT only** Write the mole equation triangle	
15. **HT only** What is Avogadro's constant?	Avogadro's constant is 6.02 x 10^{23} atoms/molecules/ions. 1 mole of copper contains 6.02 x10^{23} copper **atoms.** 1 mole of water contains 6.02 x10^{23} water **molecules.**
16. **HT only** How do we calculate number of particles from the number of moles of a substance?	Number of particles = number of moles x Avogadro's Constant
17. **HT only** Define limiting reactant.	The reactant for which there are less moles.
18. **HT only** Define excess reactant.	When the amount/moles of reactant is more than needed.

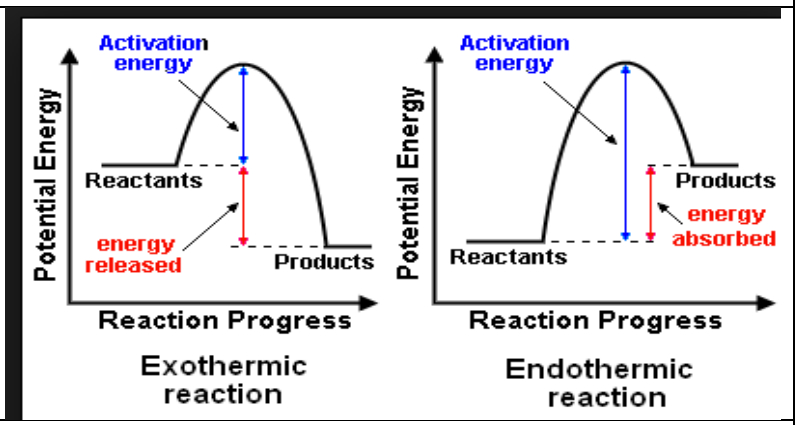
Chemical Changes

1. Write the equations for the following metals reacting with oxygen; sodium, magnesium	Sodium + oxygen → sodium oxide Magnesium + oxygen → magnesium oxide
2. Write the equations for the following metals reacting with water; sodium, magnesium	Sodium + water → sodium hydroxide + hydrogen Magnesium + water → magnesium hydroxide + hydrogen
3. What are oxidation and reduction, in terms of oxygen?	Oxidation is the gain of oxygen. Reduction is the loss of oxygen from a compound.
4. What charge do metal ions have?	Positive, because they lose electrons in chemical reactions form a full outer shell.
5. Why are group 1 metals the most reactive?	They only have 1 electron on the outer shell, so only need to lose 1 electron, so needs less energy to turn into a positive ion.
6. State whether a reaction occurs between the following metals and water or dilute acids: potassium, sodium, lithium, calcium, magnesium, zinc, iron and copper. If a reaction occurs, state the products of that reaction.	Potassium, sodium, lithium, calcium and magnesium all react with water to produce metal hydroxides and hydrogen. Zinc, iron and copper do not react with water. Potassium, sodium and lithium react **violently** with dilute acids to produce metal salts and hydrogen. Calcium and magnesium react with dilute acids to produce metal salts and hydrogen. Zinc and iron react **slowly** with dilute acids to produce metal salts and hydrogen. Copper does not react with dilute acids as it is less reactive than hydrogen
7. What is displacement?	Displacement is where a more reactive metal takes the place of a less reactive metal in a compound, for example: aluminium + iron oxide → aluminium oxide + iron.
8. Which metals can be naturally found in the Earth as pure metals?	Unreactive metals, such as gold, silver, and platinum.
9. What is a metal ore?	A naturally occurring that contains the metal in the form of a compound e.g., iron oxide
10. Which metals can be extracted from their oxides by reduction with carbon and why?	Zinc, iron, and copper can be extracted using reduction with carbon because carbon is more reactive than these metals.

11. **HT only** What is oxidation and reduction in terms of electrons?	Oxidation is loss of electrons. Reduction is gain of electrons. OILRIG
12. Write an equation for the following reactions. a) Magnesium and nitric acid b) Zinc and hydrochloric acid c) Iron and sulfuric acid	a) Magnesium + nitric acid → magnesium nitrate + hydrogen b) Zinc + hydrochloric acid → zinc chloride + hydrogen c) Iron + sulfuric acid → iron sulfate + hydrogen
13. Which ions are formed in aqueous solutions of acids and alkalis?	Acids: H^+ Alkalis: OH^-
14. How can the pH of a solution be measured?	Using universal indicator or a pH meter/probe
15. Why is a pH probe (or pH meter) more accurate than universal indicator?	It gives the precise pH value rather than having to compare colours, which could be subjective.
16. What colour ranges would indicate an acidic, neutral, or alkaline solution using universal indicator?	Red-yellow = acidic Green = neutral Blue-purple = alkaline
17. What pH are acidic, neutral, and alkaline solutions?	pH below 7 = acidic pH 7 = neutral pH above 7 = alkaline
18. What is the difference between a base and an alkali?	Bases are any substances with a pH above 7, whilst alkalis are bases that dissolved in water (soluble metal hydroxides)
19. State the ionic equation for neutralisation.	$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$ where (aq) is aqueous and (l) is liquid
20. Write an equation for the following reactions. a) Calcium oxide and nitric acid b) Zinc hydroxide and hydrochloric acid c) Iron carbonate and sulfuric acid	a) Calcium oxide + nitric acid → calcium nitrate + water b) Zinc hydroxide + hydrochloric acid → zinc chloride + water c) Iron carbonate + sulfuric acid → iron sulfate + carbon dioxide + water
21. **HT only** What is the difference between weak and strong acids?	Strong acids fully ionise in aqueous solutions and have more H+ ions. Weak acids only partially ionise in aqueous solution to release very few H+ ions
22. **HT only** Give examples of strong and weak acids	• Strong: hydrochloric acid, sulfuric acid, nitric acid • Weak: ethanoic acid, citric acid and carbonic acid.
23. **HT only** How is the pH of an acid related to its strength?	• The stronger an acid, the lower its pH.

24. **HT only** What is the difference between dilute and concentrated?	Dilute is when there a very few (acid) particles dissolved in a given volume of water. Concentrated is when there are lots of (acid) particles dissolved in a given volume of water.
25. **HT only** What is the relationship between pH and hydrogen ion concentration?	As the pH decreases by one unit, the hydrogen ion concentration of the solution increases by a factor of 10.
26. What is electrolysis?	The splitting of molten or dissolved ionic compounds using electricity.
27. Why do ionic compounds need to be aqueous or molten to be electrolysed?	When molten or aqueous, the ions in ionic compounds can move and therefore can conduct electricity.
28. Why are electrodes made from graphite?	Electrodes are made from graphite as it is cheap, has a very high melting point and can conduct electricity.
29. Why does electrolysis use large amounts of energy (and is therefore more expensive) compared reduction with carbon?	Large amounts of energy are used in the extraction process to melt the compounds and to produce the electrical current.
30. How can aluminium be extracted from its ore (bauxite)?	Aluminium is manufactured by the electrolysis of a molten mixture of aluminium oxide and cryolite
31. Why is a mixture of cryolite and aluminium oxide used in the manufacture of aluminium?	A mixture is used to lower the melting point of aluminium oxide. Therefore, less energy is required to produce the molten mixture required for electrolysis.
32. Why can't aluminium be extracted from its ore using reduction with carbon?	Aluminium is more reactive than carbon.
33. Why does the electrode need to be replaced regularly in the electrolysis of metal oxides, e.g., Aluminium oxide?	In the electrolysis of aluminium oxide, the positive electrode wears away as oxygen formed reacts with carbon to form carbon dioxide.
34. Describe and explain what forms at the positive and negative electrode of a molten ionic compound such as calcium chloride.	Calcium ions are the positive metal ions and are attracted to the negative electrode where is gains electrons (reduction) to form calcium atoms. $Ca^{2+} + 2e^- \rightarrow Ca$ Chloride ions are the negative non-metal ions and are attracted to the positive electrode where it loses electrons (oxidation) to form chlorine molecules. $2Cl^- \rightarrow Cl_2 + 2e^-$

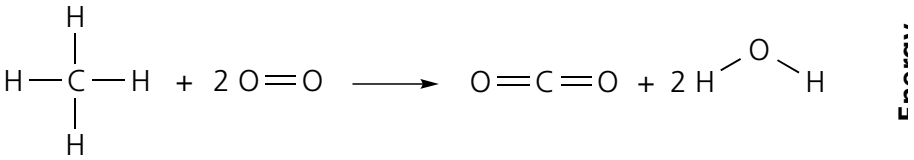
Energy Changes

1. Why does the total amount of energy in a reaction never change?	Energy is NEVER created or destroyed. Therefore, the total amount of energy at the end of a reaction must have been the same as the start.
2. What is an exothermic reaction in terms of temperature and energy change?	Temperature increases and energy is released into the surroundings. Products have less energy than reactants.
3. **HT only** What is an exothermic reaction in terms of the energy of the reactants and products?	An exothermic reaction occurs when the energy released by making bonds in the products is greater than energy needed to break the bonds in the reactants.
4. What is an endothermic reaction in terms of temperature and energy change?	Temperature decreases and heat energy is taken in from the surroundings. Products have more energy than surroundings
5. **HT only** What is an endothermic reaction in terms of the energy of the reactants?	An endothermic reaction occurs when the energy released by making bonds in the products is less than energy needed to break the bonds in the reactants.
6. Draw and label exothermic and endothermic energy diagrams.	
7. What is the collision theory?	Collision theory states that a reaction only occurs if particles collide with enough energy (activation energy).
8. What is activation energy?	Activation energy is the minimum energy needed for particles to react when they collide.
9. Give 3 examples of exothermic reactions.	Neutralisation, combustion, Self-heating cans
10. Give 2 examples of an endothermic reaction.	Thermal decomposition reactions, reaction of citric acid with sodium hydrogencarbonate

HT only - Calculating overall energy changes.
Worked example.

Find the energy change in the following reaction using the bond energies given.

Bond energies: C—H 412 kJ, O═O 496 kJ, C═O 743 kJ,
O—H 463 kJ



Explain why the reaction is exothermic or endothermic using bond energies.

Answer

Bonds broken:	Bonds made:
4C—H = 4(412)	2C═O = 2(743)
2O═O = 2(496)	4O—H = 4(463)
Total = 2640 kJ	Total = 3338 kJ

Energy change = energy needed to break bonds
– energy released making bonds

= 2640 – 3338

= –698 kJ

This reaction is exothermic because more energy is released making bonds than is needed to break bonds.